Physical Sciences Grade 10 WebBook

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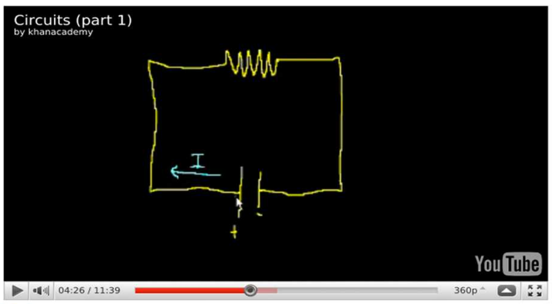

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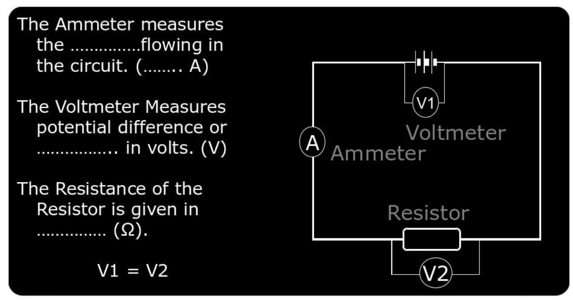

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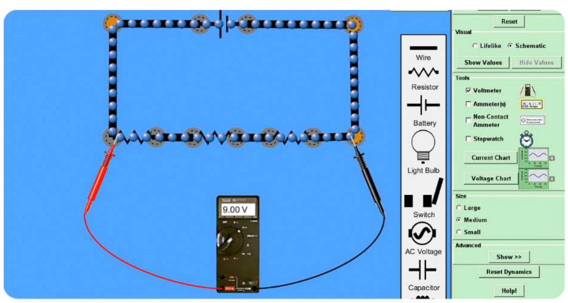

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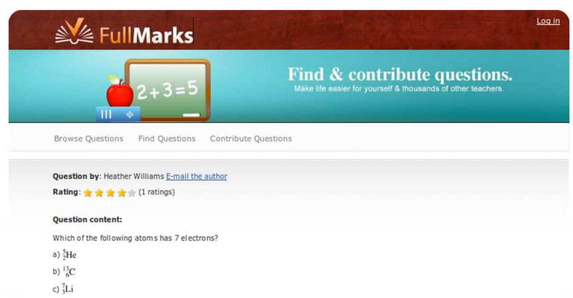

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Periodic Table of the Elements

No EN
Element
AMU

1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VII	9 VII	10 VII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 0
1 2.2 **H** 1,01																	2 **He** 4,00
3 0.98 **Li** 6,94	4 1.57 **Be** 9,01											5 2.04 **B** 10,8	6 2.55 **C** 12,0	7 3.04 **N** 14,0	8 3.44 **O** 16,0	9 3.98 **F** 19,0	10 **Ne** 20,2
11 0.93 **Na** 23,0	12 1.31 **Mg** 24,3											13 1.61 **Al** 27,0	14 1.90 **Si** 28,1	15 2.19 **P** 31,0	16 2.58 **S** 32,1	17 3.16 **Cl** 35,45	18 **Ar** 39,9
19 0.93 **K** 39,1	20 1.00 **Ca** 40,1	21 1.36 **Sc** 45,0	22 1.54 **Ti** 47,9	23 1.63 **V** 50,9	24 1.66 **Cr** 52,0	25 1.55 **Mn** 54,9	26 1.33 **Fe** 55,8	27 1.88 **Co** 58,9	28 1.91 **Ni** 58,7	29 1.90 **Cu** 63,5	30 1.65 **Zn** 65,4	31 1.81 **Ga** 69,7	32 2.01 **Ge** 72,6	33 2.18 **As** 74,9	34 2.55 **Se** 79,0	35 2.96 **Br** 79,9	36 **Kr** 83,8
37 0.82 **Rb** 85,5	38 0.95 **Sr** 87,6	39 1.22 **Y** 88,9	40 1.33 **Zr** 91,2	41 1.60 **Nb** 92,9	42 2.16 **Mo** 95,9	43 1.90 **Tc** (98)	44 2.20 **Ru** 101,1	45 2.28 **Rh** 102,9	46 2.20 **Pd** 106,4	47 1.93 **Ag** 107,9	48 1.69 **Cd** 112,4	49 1.78 **In** 114,8	50 1.96 **Sn** 118,7	51 2.05 **Sb** 121,8	52 2.10 **Te** 127,6	53 2.66 **I** 126,9	54 **Xe** 131,3
55 0.79 **Cs** 132,9	56 0.89 **Ba** 137,3	57-71 **La-Lu** Lanthanides	72 1.30 **Hf** 178,5	73 1.50 **Ta** 180,9	74 2.36 **W** 183,8	75 1.90 **Re** 186,2	76 2.20 **Os** 190,2	77 2.20 **Ir** 192,2	78 2.28 **Pt** 195,1	79 2.54 **Au** 197,0	80 2.00 **Hg** 200,6	81 1.62 **Tl** 204,4	82 2.33 **Pb** 207,2	83 2.02 **Bi** 209,0	84 2.00 **Po** (209)	85 2.20 **At** (210)	86 **Rn** (222)
87 0.7 **Fr** (223)	88 0.9 **Ra** 226,0	89-103 **Ac-Lr** Actinides	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (263)	107 **Bh** (262)	108 **Hs** (265)	109 **Mt** (266)	110 **Ds** (269)	111 **Rg** (272)	112 **Cn** (277)	113 **Uut** (284)	114 **Uuq** (289)	115 **Uup** (288)	116 **Uuh** (293)	117 **Uus** (282)	118 **Uuo** (282)
			57 1.10 **La** 138,9	58 1.12 **Ce** 140,1	59 1.13 **Pr** 140,9	60 1.14 **Nd** 144,2	61 **Pm** (145)	62 1.17 **Sm** 150,4	63 **Eu** 152,0	64 1.20 **Gd** 157,3	65 **Tb** 158,9	66 1.22 **Dy** 162,5	67 1.23 **Ho** 164,9	68 1.24 **Er** 167,3	69 1.25 **Tm** 168,9	70 **Yb** 173,0	71 1.27 **Lu** 175,0
			89 1.10 **Ac** 227,0	90 1.30 **Th** 232,0	91 1.50 **Pa** 231,0	92 1.38 **U** 238,0	93 1.36 **Np** 237,0	94 1.28 **Pu** (244)	95 1.30 **Am** (243)	96 1.30 **Cm** (247)	97 1.30 **Bk** (247)	98 1.30 **Cf** (251)	99 1.30 **Es** (252)	100 1.30 **Fm** (257)	101 1.30 **Md** (258)	102 1.30 **No** (258)	103 **Lr** (260)

- Transition Metal
- Metal
- Metalloid
- Non-metal
- Noble Gas
- Lanthanide
- Actinide

Contents

Chapter 1

Units

1.1 Introduction

(section shortcode: U10000)

Imagine you had to make curtains and needed to buy fabric. The shop assistant would need to know how much fabric you needed. Telling her you need fabric 2 wide and 6 long would be insufficient — you have to specify the **unit** (i.e. 2 *metres* wide and 6 *metres* long). Without the unit the information is incomplete and the shop assistant would have to guess. If you were making curtains for a doll's house the dimensions might be 2 centimetres wide and 6 centimetres long!

It is not just lengths that have units, all physical quantities have units (e.g. time, temperature, distance, etc.).

Definition: Physical Quantity
A physical quantity is anything that you can measure. For example, length, temperature, distance and time are physical quantities.

1.2 Unit Systems

(section shortcode: U10001)

1.2.1 SI Units

We will be using the SI units in this course. SI units are the internationally agreed upon units. Historically these units are based on the metric system which was developed in France at the time of the French Revolution.

Definition: SI Units
The name *SI units* comes from the French *Système International d'Unités*, which means *international system of units*.

There are seven base SI units. These are listed in Table 1. All physical quantities have units which can be built from these seven base units. So, it is possible to create a different set of units by defining a different set of base units.

These seven units are called base units because none of them can be expressed as combinations of the other six. This is identical to bricks and concrete being the base units of a building. You can build different things using different combinations of bricks and concrete. The 26 letters of the alphabet are the base units for a language like English. Many different words can be formed by using these letters.

Base quantity	**Name**	**Symbol**
length	metre	m
mass	kilogram	kg
time	second	s
electric current	ampere	A
temperature	kelvin	K
amount of substance	mole	mol
luminous intensity	candela	cd

Table 1: SI Base Units

1.2.2 The Other Systems of Units

The SI Units are not the only units available, but they are most widely used. In Science there are three other sets of units that can also be used. These are mentioned here for interest only.

1.2.3 c.g.s. Units

In the c.g.s. system, the metre is replaced by the centimetre and the kilogram is replaced by the gram. This is a simple change but it means that all units derived from these two are changed. For example, the units of force and work are different. These units are used most often in astrophysics and atomic physics.

1.2.4 Imperial Units

Imperial units arose when kings and queens decided the measures that were to be used in the land. All the imperial base units, except for the measure of time, are different to those of SI units. This is the unit system you are most likely to encounter if SI units are not used. Examples of imperial units are pounds, miles, gallons and yards. These units are used by the Americans and British. As you can imagine, having different units in use from place to place makes scientific communication very difficult. This was the motivation for adopting a set of internationally agreed upon units.

1.2.5 Natural Units

This is the most sophisticated choice of units. Here the most fundamental discovered quantities (such as the speed of light) are set equal to 1. The argument for this choice is that all other quantities should be built from these fundamental units. This system of units is used in high energy physics and quantum mechanics.

1.3 Writing Units as Words or Symbols

(section shortcode: U10002)

Unit names are always written with a lowercase first letter, for example, we write metre and litre. The symbols or abbreviations of units are also written with lowercase initials, for example m for metre and ℓ for litre. The exception to this rule is if the unit is named after a person, then the symbol is a capital letter. For example, the kelvin was named after Lord Kelvin and its symbol is K. If the abbreviation of the unit that is named after a person has two letters, the second letter is lowercase, for example Hz for hertz.

1.3.1 Naming of Units

For the following symbols of units that you will come across later in this book, write whether you think the unit is named after a person or not.

1. J (joule)
2. ℓ (litre)
3. N (newton)
4. mol (mole)
5. C (coulomb)
6. lm (lumen)
7. m (metre)
8. bar (bar)

Find the answers with the shortcodes:

(1.) lOX

1.4 Combinations of SI Base Units

(section shortcode: U10003)

To make working with units easier, some combinations of the base units are given special names, but it is always correct to reduce everything to the base units. Table 2 lists some examples of combinations of SI base units that are assigned special names. Do not be concerned if the formulae look unfamiliar at this stage - we will deal with each in detail in the chapters ahead (as well as many others)!

It is very important that you are able to recognise the units correctly. For instance, the **n**ewton (N) is another name for the **kilogram metre per second squared** (kg·m·s^{-2}), while the **k**ilogram metre squared per second squared (kg·m^2·s^{-2}) is called the **j**oule (J).

Quantity	Formula	Unit Expressed in Base Units	Name of Combination
Force	ma	kg·m·s^{-2}	N (newton)
Frequency	$\frac{1}{T}$	s^{-1}	Hz (hertz)
Work	Fs	kg·m^2·s^{-2}	J (joule)

Table 2: Some examples of combinations of SI base units assigned special names

TIP: When writing combinations of base SI units, place a dot (·) between the units to indicate that different base units are used. For example, the symbol for metres per second is correctly written as m·s^{-1}, and not as ms^{-1} or m/s. Although the last two options will be accepted in tests and exams, we will only use the first one in this book.

1.5 Rounding, Scientific Notation and Significant Figures

(section shortcode: U10004)

1.5.1 Rounding Off

Certain numbers may take an infinite amount of paper and ink to write out. Not only is that impossible, but writing numbers out to a high precision (many decimal places) is very inconvenient and rarely gives better answers. For this reason we often estimate the number to a certain number of decimal places. Rounding off or approximating a decimal number to a given number of decimal places is the quickest way to approximate a number. For example, if you wanted to round-off $2,6525272$ to three decimal places then you would first count three places after the decimal. $2,652|5272$ All numbers to the right of $|$ are ignored after you determine whether the number in the third decimal place must be rounded up or rounded down. You *round up* the final digit (make the digit one more) if the first digit after the $|$ was greater or equal to 5 and *round down* (leave the digit alone) otherwise. So, since the first digit after the $|$ is a 5, we must round up the digit in the third decimal place to a 3 and the final answer of $2,6525272$ rounded to three decimal places is 2,653.

Exercise 1: Rounding-off Round off $\pi = 3,141592654...$ to 4 decimal places.

Solution to Exercise

Step 1. $\pi = 3,1415|92654...$
Step 2. The last digit of $\pi = 3,1415|92654...$ must be rounded up because there is a 9 after the $|$.
Step 3. $\pi = 3,1416$ rounded to 4 decimal places.

Exercise 2: Rounding-off Round off $9,191919...$ to 2 decimal places

Solution to Exercise

Step 1. $9,19|1919...$
Step 2. The last digit of $9,19|1919...$ must be rounded down because there is a 1 after the $|$.
Step 3. Answer = 9,19 rounded to 2 decimal places.

1.5.2 Error Margins

In a calculation that has many steps, it is best to leave the rounding off right until the end. For example, Jack and Jill walk to school. They walk 0,9 kilometers to get to school and it takes them 17 minutes. We can calculate their speed in the following two ways.

Method 1:

$$\begin{aligned} \mathrm{time in hours} &= \frac{17\mathrm{min}}{60\mathrm{min}} \\ &= 0,283333333\mathrm{h} \end{aligned} \tag{1}$$

$$\begin{aligned} \mathrm{speed} &= \frac{\mathrm{Distance}}{\mathrm{Time}} \\ &= \frac{0{,}9\mathrm{km}}{0{,}28333333\mathrm{h}} \\ &= 3,176470588 km \cdot h^{-1} \\ &= 3,18 km \cdot h^{-1} \end{aligned} \tag{2}$$

Method 2:

$$\begin{aligned} \mathrm{time in hours} &= \frac{17\mathrm{min}}{60\mathrm{min}} \\ &= 0,28\mathrm{h} \end{aligned} \tag{3}$$

$$\begin{aligned} \text{speed} &= \frac{\text{Distance}}{\text{Time}} \\ &= \frac{0{,}9\text{km}}{0{,}28\text{h}} \\ &= 3,214285714 km \cdot h^{-1} \\ &= 3,21 km \cdot h^{-1} \end{aligned} \tag{4}$$

You will see that we get two different answers. In Method 1 no rounding was done, but in Method 2, the time was rounded to 2 decimal places. This made a big difference to the answer. The answer in Method 1 is more accurate because rounded numbers were not used in the calculation. Always only round off your final answer.

1.5.3 Scientific Notation

In Science one often needs to work with very large or very small numbers. These can be written more easily in scientific notation, in the general form

$$d \times 10^e \tag{5}$$

where d is a decimal number between 0 and 10 that is rounded off to a few decimal places. e is known as the *exponent* and is an integer. If $e > 0$ it represents how many times the decimal place in d should be moved to the right. If $e < 0$, then it represents how many times the decimal place in d should be moved to the left. For example $3,24 \times 10^3$ represents 3240 (the decimal moved three places to the right) and $3,24 \times 10^{-3}$ represents $0,00324$ (the decimal moved three places to the left).

If a number must be converted into scientific notation, we need to work out how many times the number must be multiplied or divided by 10 to make it into a number between 1 and 10 (i.e. the value of e) and what this number between 1 and 10 is (the value of d). We do this by counting the number of decimal places the decimal comma must move.

For example, write the speed of light in scientific notation, to two decimal places. The speed of light is 299 792 458 m·s^{-1}. First, find where the decimal comma must go for two decimal places (to find d) and then count how many places there are after the decimal comma to determine e.

In this example, the decimal comma must go after the first 2, but since the number after the 9 is 7, $d = 3,00$. $e = 8$ because there are 8 digits left after the decimal comma. So the speed of light in scientific notation, to two decimal places is 3,00 $\times$ 10^8 m·s^{-1}.

1.5.4 Significant Figures

In a number, each non-zero digit is a significant figure. Zeroes are only counted if they are between two non-zero digits or are at the end of the decimal part. For example, the number 2000 has 1 significant figure (the 2), but 2000,0 has 5 significant figures. You estimate a number like this by removing significant figures from the number (starting from the right) until you have the desired number of significant figures, rounding as you go. For example 6,827 has 4 significant figures, but if you wish to write it to 3 significant figures it would mean removing the 7 and rounding up, so it would be 6,83.

1.5.5 Using Significant Figures

1. Round the following numbers:
   - **a.** 123,517 ℓ to 2 decimal places
   - **b.** 14,328 km·h^{-1} to one decimal place
   - **c.** 0,00954 m to 3 decimal places
2. Write the following quantities in scientific notation:
   - **a.** 10130 Pa to 2 decimal places
   - **b.** 978,15 m·s^{-2} to one decimal place
   - **c.** 0,000001256 A to 3 decimal places
3. Count how many significant figures each of the quantities below has:
   - **a.** 2,590 km
   - **b.** 12,305 mℓ
   - **c.** 7800 kg

Find the answers with the shortcodes:

(1.) IOI (2.) IO5 (3.) ION

1.6 Prefixes of Base Units

(section shortcode: U10005)

Now that you know how to write numbers in scientific notation, another important aspect of units is the prefixes that are used with the units.

Definition: Prefix
A prefix is a group of letters that are placed in front of a word. The effect of the prefix is to change meaning of the word. For example, the prefix *un* is often added to a word to mean *not*, as in *un*necessary which means *not necessary*.

In the case of units, the prefixes have a special use. The kilogram (kg) is a simple example. 1 kg is equal to 1 000 g or 1×10^3 g. Grouping the 10^3 and the g together we can replace the 10^3 with the prefix k (kilo). Therefore the k takes the place of the 10^3. The kilogram is unique in that it is the only SI base unit containing a prefix.

In Science, all the prefixes used with units are some power of 10. Table 3 lists some of these prefixes. You will not use most of these prefixes, but those prefixes listed in **bold** should be learnt. The case of the prefix symbol is very important. Where a letter features twice in the table, it is written in uppercase for exponents bigger than one and in lowercase for exponents less than one. For example M means mega (10^{6}) and m means milli (10^{-3}).

Prefix	Symbol	Exponent	Prefix	Symbol	Exponent
yotta	Y	10^{24}	yocto	y	10^{-24}
zetta	Z	10^{21}	zepto	z	10^{-21}
exa	E	10^{18}	atto	a	10^{-18}
peta	P	10^{15}	femto	f	10^{-15}
tera	T	10^{12}	pico	p	10^{-12}
giga	G	10^{9}	**nano**	n	10^{-9}
mega	M	10^{6}	**micro**	μ	10^{-6}
kilo	k	10^{3}	**milli**	m	10^{-3}
hecto	h	10^{2}	**centi**	c	10^{-2}
deca	da	10^{1}	**deci**	d	10^{-1}

Table 3: Unit Prefixes

TIP: There is no space and no dot between the prefix and the symbol for the unit.

Here are some examples of the use of prefixes:

- 40000 m can be written as 40 km (kilometre)
- 0,001 g is the same as 1×10^{-3} g and can be written as 1 mg (milligram)
- $2,5 \times 10^{6}$ N can be written as 2,5 MN (meganewton)
- 250000 A can be written as 250 kA (kiloampere) or 0,250 MA (megaampere)
- 0,000000075 s can be written as 75 ns (nanoseconds)
- 3×10^{-7} mol can be rewritten as $0,3 \times 10^{-6}$ mol, which is the same as 0,3 μmol (micromol)

1.6.1 Using Scientific Notation

1. Write the following in scientific notation using Table 3 as a reference.

a. 0,511 MV
b. 10 cℓ
c. 0,5 μm
d. 250 nm
e. 0,00035 hg

2. Write the following using the prefixes in Table 3.

a. 1,602 $\times 10^{-19}$ C
b. 1,992 $\times 10^{6}$ J
c. 5,98 $\times 10^{4}$ N
d. 25 $\times 10^{-4}$ A

e. 0,0075 $\times 10^6$ m

Find the answers with the shortcodes:

(1.) lOR (2.) lOn

1.7 The Importance of Units

(section shortcode: U10006)

Without units much of our work as scientists would be meaningless. We need to express our thoughts clearly and units give meaning to the numbers we measure and calculate. Depending on which units we use, the numbers are different. For example if you have 12 water, it means nothing. You could have 12 ml of water, 12 litres of water, or even 12 bottles of water. Units are an essential part of the language we use. Units must be specified when expressing physical quantities. Imagine that you are baking a cake, but the units, like grams and millilitres, for the flour, milk, sugar and baking powder are not specified!

1.7.1 Investigation : Importance of Units

Work in groups of 5 to discuss other possible situations where using the incorrect set of units can be to your disadvantage or even dangerous. Look for examples at home, at school, at a hospital, when travelling and in a shop.

1.7.2 Case Study : The importance of units

Read the following extract from CNN News 30 September 1999 and answer the questions below.

NASA: Human error caused loss of Mars orbiter November 10, 1999

Failure to convert English measures to metric values caused the loss of the Mars Climate Orbiter, a spacecraft that smashed into the planet instead of reaching a safe orbit, a NASA investigation concluded Wednesday. The Mars Climate Orbiter, a key craft in the space agency's exploration of the red planet, vanished on 23 September after a 10 month journey. It is believed that the craft came dangerously close to the atmosphere of Mars, where it presumably burned and broke into pieces. An investigation board concluded that NASA engineers failed to convert English measures of rocket thrusts to newton, a metric system measuring rocket force. One English pound of force equals 4,45 newtons. A small difference between the two values caused the spacecraft to approach Mars at too low an altitude and the craft is thought to have smashed into the planet's atmosphere and was destroyed. The spacecraft was to be a key part of the exploration of the planet. From its station about the red planet, the Mars Climate Orbiter was to relay signals from the Mars Polar Lander, which is scheduled to touch down on Mars next month. "The root cause of the loss of the spacecraft was a failed translation of English units into metric units and a segment of ground-based, navigation-related mission software," said Arthus Stephenson, chairman of the investigation board. **Q**uestions:

1. Why did the Mars Climate Orbiter crash? Answer in your own words.
2. How could this have been avoided?
3. Why was the Mars Orbiter sent to Mars?
4. Do you think space exploration is important? Explain your answer.

1.8 How to Change Units

(section shortcode: U10007)

It is very important that you are aware that different systems of units exist. Furthermore, you must be able to convert between units. Being able to change between units (for example, converting from millimetres to metres) is a useful skill in Science.

The following conversion diagrams will help you change from one unit to another.

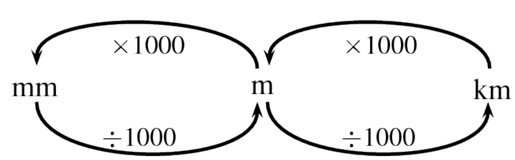


Figure 1: The distance conversion table

If you want to change millimetre to metre, you divide by 1000 (follow the arrow from mm to m); or if you want to change kilometre to millimetre, you multiply by 1000×1000.

The same method can be used to change millilitre to litre or kilolitre. Use Figure 2 to change volumes:

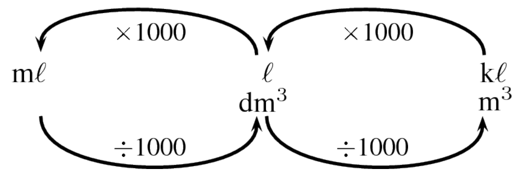


Figure 2: The volume conversion table

Exercise 3: Conversion 1 Express 3 800 mm in metres.

Solution to Exercise

Step 1. Use Figure 1 . Millimetre is on the left and metre in the middle.
Step 2. You need to go from mm to m, so you are moving from left to right.
Step 3. 3 800 mm ÷ 1000 = 3,8 m

Exercise 4: Conversion 2 Convert 4,56 kg to g.

Solution to Exercise

Step 1. Use Figure 1. Kilogram is the same as kilometre and gram the same as metre.
Step 2. You need to go from kg to g, so it is from right to left.
Step 3. 4,56 kg × 1000 = 4560 g

1.8.1 Two other useful conversions

Very often in Science you need to convert speed and temperature. The following two rules will help you do this:

Converting speed When converting km·h^{-1} to m·s^{-1}you divide by 3,6.
For example 72 km·h^{-1}$\div$ 3,6 = 20 m·s^{-1}.

When converting m·s^{-1}to km·h^{-1}, you multiply by 3,6. For example 30 m·s^{-1}$\times$3,6 = 108 km·h^{-1}.

Converting temperature Converting between the kelvin and celsius temperature scales is easy. To convert from celsius to kelvin add 273. To convert from kelvin to celsius subtract 273. Representing the kelvin temperature by T_K and the celsius temperature by $T_{^\circ C}$,

$$T_K = T_{^\circ C} + 273 \tag{6}$$

1.9 A sanity test

(section shortcode: U10008)

A sanity test is a method of checking whether an answer makes sense. All we have to do is to take a careful look at our answer and ask the question *Does the answer make sense?*

Imagine you were calculating the number of people in a classroom. If the answer you got was 1 000 000 people you would know it was wrong — it is not possible to have that many people in a classroom. That is all a sanity test is — is your answer insane or not?

It is useful to have an idea of some numbers before we start. For example, let us consider masses. An average person has a mass around 70 kg, while the heaviest person in medical history had a mass of 635 kg. If you ever have to calculate a person's mass and you get 7 000 kg, this should fail your sanity check — your answer is insane and you must have made a mistake somewhere. In the same way an answer of 0.01 kg should fail your sanity test.

The only problem with a sanity check is that you must know what typical values for things are. For example, finding the number of learners in a classroom you need to know that there are usually 20–50 people in a classroom. If you get and answer of 2500, you should realise that it is wrong.

1.9.1 The scale of the matter... :

Try to get an idea of the typical values for the following physical quantities and write your answers into the table:

Category	Quantity	Minimum	Maximum
People	mass		
	height		
Transport	speed of cars on freeways		
	speed of trains		
	speed of aeroplanes		
	distance between home and school		
General	thickness of a sheet of paper		
	height of a doorway		

Table 4

1.10 Summary

(section shortcode: U10009)

1. You need to know the seven base SI Units as listed in Table 1. Combinations of SI Units can have different names.
2. Unit names and abbreviations are written with lowercase letter unless it is named after a person.
3. Rounding numbers and using scientific notation is important.
4. Table 3 summarises the prefixes used in Science.
5. Use figures Figure 1 and Figure 2 to convert between units.

1.11 End of Chapter Exercises

(section shortcode: U10010)

1. Write down the SI unit for the each of the following quantities:
   a. length
   b. time
   c. mass
   d. quantity of matter
2. For each of the following units, write down the symbol and what power of 10 it represents:
   a. millimetre
   b. centimetre
   c. metre
   d. kilometre

3. For each of the following symbols, write out the unit in full and write what power of 10 it represents:
   a. μg
   b. mg
   c. kg
   d. Mg
4. Write each of the following in scientific notation, correct to 2 decimal places:
   a. 0,00000123 N
   b. 417 000 000 kg
   c. 246800 A
   d. 0,00088 mm
5. Rewrite each of the following, accurate to two decimal places, using the correct prefix where applicable:
   a. 0,00000123 N
   b. 417 000 000 kg
   c. 246800 A
   d. 0,00088 mm
6. For each of the following, write the measurement using the correct symbol for the prefix and the base unit:
   a. 1,01 microseconds
   b. 1 000 milligrams
   c. 7,2 megameters
   d. 11 nanolitre
7. The Concorde is a type of aeroplane that flies very fast. The top speed of the Concorde is 2 172 km·hr^{-1}. Convert the Concorde's top speed to m·s^{-1}.
8. The boiling point of water is 100 °C. What is the boiling point of water in kelvin?

Find the answers with the shortcodes:

(1.) IOQ (2.) IOU (3.) IOP (4.) IOE (5.) IOm (6.) IOy
(7.) IOV (8.) IOp

Part I

Chemistry

Classification of Matter

1.12 Introduction

(section shortcode: C10000)

All the objects that we see in the world around us, are made of **matter**. Matter makes up the air we breathe, the ground we walk on, the food we eat and the animals and plants that live around us. Even our own human bodies are made of matter!

Different objects can be made of different types of matter, or **materials**. For example, a cupboard (an *object*) is made of wood, nails and hinges (the *materials*). The **properties** of the materials will affect the properties of the object. In the example of the cupboard, the strength of the wood and metals make the cupboard strong and durable. In the same way, the raincoats that you wear during bad weather, are made of a material that is waterproof. The electrical wires in your home are made of metal because metals are a type of material that is able to conduct electricity. It is very important to understand the properties of materials, so that we can use them in our homes, in industry and in other applications. In this chapter, we will be looking at different types of materials and their properties.

Some of the properties of matter that you should know are:

- Materials can be strong and resist bending (e.g. iron rods, cement) or weak (e.g. fabrics)
- Materials that conduct heat (e.g. metals) are called thermal conductors. Materials that conduct electricity are electrical conductors.
- Brittle materials break easily. Materials that are malleable can be easily formed into different shapes. Ductile materials are able to be formed into long wires.
- Magnetic materials have a magnetic field.
- Density is the mass per unit volume. An example of a dense material is concrete.
- The boiling and melting points of substance help us to classify substances as solids, liquids or gases at a specific temperature.

The diagram below shows one way in which matter can be classified (grouped) according to its different properties. As you read further in this chapter, you will see that there are also other ways of classifying materials, for example according to whether or not they are good electrical conductors.

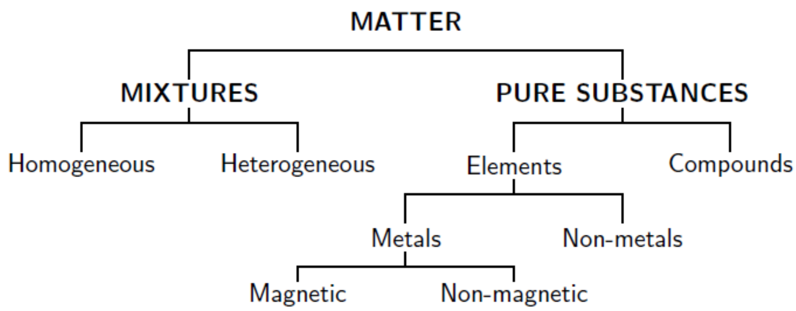


Figure 2.1: The classification of matter

Discussion: Everyday materials: In groups of 3 or 4 look at the labels of medicines, food items, and any other items that you use often. What can you tell about the material inside the container from the list of ingredients? Why is it important to have a list of ingredients on the materials that we use? Do some research on the safety data of the various compounds in the items that you looked at. Are the compounds in the items safe to use? In the food items, what preservatives and additives are there? Are these preservatives and additives good for you? Are there natural alternatives (natural alternatives are usually used by indigenous people groups)?

1.13 Mixtures

(section shortcode: C10001)

We see mixtures all the time in our everyday lives. A stew, for example, is a mixture of different foods such as meat and vegetables; sea water is a mixture of water, salt and other substances, and air is a mixture of gases such as carbon dioxide, oxygen and nitrogen.

Definition: Mixture
A **mixture** is a combination of two or more substances, where these substances are not bonded (or joined) to each other.

In a mixture, the substances that make up the mixture:

- *are not in a fixed ratio* Imagine, for example, that you have a 250 ml beaker of water. It doesn't matter whether you add 20 g, 40 g, 100 g or any other mass of sand to the water; it will still be called a mixture of sand and water.
- *keep their physical properties* In the example we used of the sand and water, neither of these substances has changed in any way when they are mixed together. Even though the sand is in water, it still has the same properties as when it was out of the water.

- *can be separated by mechanical means* To separate something by 'mechanical means', means that there is no chemical process involved. In our sand and water example, it is possible to separate the mixture by simply pouring the water through a filter. Something *physical* is done to the mixture, rather than something *chemical*.

Some other examples of mixtures include blood (a mixture of blood cells, platelets and plasma), steel (a mixture of iron and other materials) and the gold that is used to make jewellery. The gold in jewellery is not pure gold but is a mixture of metals. The amount of gold in the jewellery is measured in *karats* (24 karat would be pure gold, while 18 karat is only 75% gold).

We can group mixtures further by dividing them into those that are heterogeneous and those that are homogeneous.

1.13.1 Heterogeneous mixtures

A **heterogeneous** mixture does not have a definite composition. Think of a pizza, that has a topping of cheese, tomato, mushrooms and peppers (the topping is a mixture). Each slice will probably be slightly different from the next because the toppings (the tomato, cheese, mushrooms and peppers) are not evenly distributed. Another example would be granite, a type of rock. Granite is made up of lots of different mineral substances including quartz and feldspar. But these minerals are not spread evenly through the rock and so some parts of the rock may have more quartz than others. Another example is a mixture of oil and water. Although you may add one substance to the other, they will stay separate in the mixture. We say that these heterogeneous mixtures are *non-uniform*, in other words they are not exactly the same throughout.

Definition: Heterogeneous mixture
A heterogeneous mixture is one that is non-uniform and the different components of the mixture can be seen.

1.13.2 Homogeneous mixtures

A **homogeneous** mixture has a definite composition, and specific properties. In a homogeneous mixture, the different parts cannot be seen. A solution of salt dissolved in water is an example of a homogeneous mixture. When the salt dissolves, it will spread evenly through the water so that all parts of the solution are the same, and you can no longer see the salt as being separate from the water. Think also of a powdered drink that you mix with water. Provided you give the container a good shake after you have added the powder to the water, the drink will have the same sweet taste for anyone who drinks it, it won't matter whether they take a sip from the top or from the bottom. The air we breathe is another example of a homogeneous mixture since it is made up of different gases which are in a constant ratio, and which can't be distinguished from each other.

Definition: Homogeneous mixture
A homogeneous mixture is one that is uniform, and where the different components of the mixture cannot be seen.

An **alloy** is a homogeneous mixture of two or more elements, at least one of which is a metal, where the resulting material has metallic properties. Alloys are usually made to improve the properties of the elements that make them up. For example steel is much stronger than iron (which is the main component of steel).

Activity: Classifying materials: Look around your classroom or school. Make a list of all the different materials that you see around you. Try to work out why a particular material was used. Can you classify all the different materials used according to their properties? On your way to school or at home or in the shops, look at the different materials that are used. Why are these materials chosen over other materials?

Activity: Making mixtures: Make mixtures of sand and water, potassium dichromate and water, iodine and ethanol, iodine and water. Classify these as heterogeneous or homogeneous. Try to make mixtures using other substances. Are the mixtures that you have made heterogeneous or homogeneous? Give reasons for your choice.

1.13.3 Mixtures

1. Which of the following substances are *mixtures*?
   - **a.** tap water
   - **b.** brass (an alloy of copper and zinc)
   - **c.** concrete
   - **d.** aluminium
   - **e.** Coca cola
   - **f.** distilled water
2. In each of the examples above, say whether the mixture is homogeneous or heterogeneous.

www Find the answers with the shortcodes:

(1.) llm

1.14 Pure Substances: Elements and Compounds

www (section shortcode: C10002)

Any material that is not a mixture, is called a **pure substance**. Pure substances include **elements** and **compounds**. It is much more difficult to break down pure substances into their parts, and complex chemical methods are needed to do this.

One way to determine if a substance is pure is to look at its melting or boiling point. Pure substances will have a sharply defined melting or boiling point (i.e. the melting or boiling point will be a single temperature rather than a range of temperatures.) Impure substances have a temperature range over which they melt or boil. We can also use chromatography to determine if a substance is pure or not. Chromatography is the process of separating substances into their individual components. If a substance is pure then chromatography will only produce one substance at the end of the process. If a substance is impure then several substances will be seen at the end of the process.

1.14.1 Activity: Smartie Chromatography

You will need filter paper (or chromatography paper), some smarties in different colours, water and an eye dropper.

Place a smartie in the center of a piece of filter paper. Carefully drop a few drops of water onto the smartie. You should see rings of different colour forming around the smartie. Each colour is one of the individual colours that are used to make up the colour of the smartie.

1.14.2 Elements

An **element** is a chemical substance that can't be divided or changed into other chemical substances by any ordinary chemical means. The smallest unit of an element is the **atom**.

Definition: Element
An element is a substance that cannot be broken down into other substances through chemical means.

There are 112 officially named elements and about 118 known elements. Most of these are natural, but some are man-made. The elements we know are represented in the **Periodic Table of the Elements**, where each element is abbreviated to a **chemical symbol**. Examples of elements are magnesium (Mg), hydrogen (H), oxygen (O) and carbon (C). On the Periodic Table you will notice that some of the abbreviations do not seem to match the elements they represent. The element iron, for example, has the chemical formula Fe. This is because the elements were originally given Latin names. Iron has the abbreviation Fe because its Latin name is 'ferrum'. In the same way, sodium's Latin name is 'natrium' (Na) and gold's is 'aurum' (Au).

NOTE: Recently it was agreed that two more elements would be added to the list of officially named elements. These are elements number 114 and 116. The proposed name for element 114 is flerovium and for element 116 it is moscovium. This brings the total number of officially named elements to 114.

1.14.3 Compounds

A **compound** is a chemical substance that forms when two or more elements combine in a fixed ratio. Water (H_2O), for example, is a compound that is made up of two hydrogen atoms for every one oxygen atom. Sodium chloride ($NaCl$) is a compound made up of one sodium atom for every chlorine atom. An important characteristic of a compound is that it has a **chemical formula**, which describes the ratio in which the atoms of each element in the compound occur.

Definition: Compound
A substance made up of two or more elements that are joined together in a fixed ratio.

Figure 2.2 might help you to understand the difference between the terms *element*, *mixture* and *compound*. Iron (Fe) and sulphur (S) are two elements. When they are added together, they form a *mixture* of iron and sulphur. The iron and sulphur are not joined together. However, if the mixture is heated, a new *compound* is formed, which is called iron sulphide (FeS). In this compound, the iron and sulphur are joined to each other in a ratio of 1:1. In other words, one atom of iron is joined to one atom of sulphur in the compound iron sulphide.

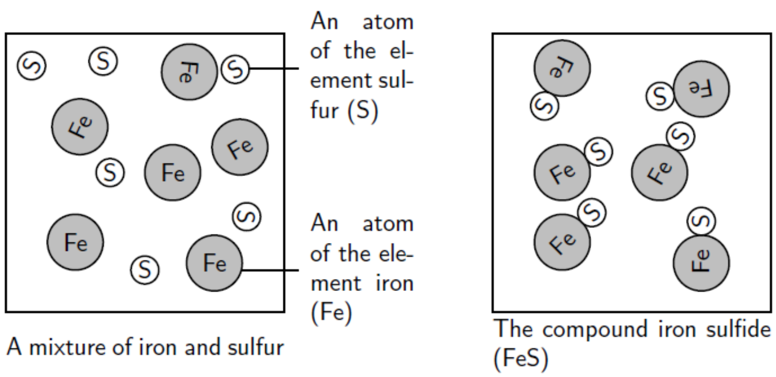


Figure 2.2: Understanding the difference between a mixture and a compound

Figure 2.2 shows the microscopic representation of mixtures and compounds. In a microscopic representation we use circles to represent different elements. To show a compound, we draw several circles joined together. Mixtures are simply shown as two or more individual elements in the same box. The circles are not joined for a mixture.

We can also use symbols to represent elements, mixtures and compounds. The symbols for the elements are all found on the periodic table. Compounds are shown as two or more element names written right next to each other. Subscripts may be used to show that there is more than one atom of a particular element. (e.g. H_2O or $NaCl$). Mixtures are written as: a mixture of element (or compound) A and element (or compound) B. (e.g. a mixture of Fe and S).

One way to think of mixtures and compounds is to think of buildings. The building is a mixture of different building materials (e.g. glass, bricks, cement, etc.). The building materials are all compounds. You can also think of the elements as Lego blocks. Each Lego block can be added to other Lego blocks to make new structures, in the same way that elements can combine to make compounds.

Activity: Using models to represent substances Use coloured balls and sticks to represent elements and compounds. Think about the way that we represent substances microscopically. Would you use just one ball to represent an element or many? Why?

Elements, mixtures and compounds

1. In the following table, tick whether each of the substances listed is a *mixture* or a *pure substance*. If it is a mixture, also say whether it is a homogeneous or heterogeneous mixture.

Substance	Mixture or pure	Homogeneous or heterogeneous mixture
fizzy colddrink		
steel		
oxygen		
iron filings		
smoke		
limestone ($CaCO_3$)		

Table 2.1

2. In each of the following cases, say whether the substance is an element, a mixture or a compound.
   - **a.** Cu
   - **b.** iron and sulphur
   - **c.** Al
   - **d.** H_2SO_4
   - **e.** SO_3

Find the answers with the shortcodes:

(1.) lly (2.) llV

1.15 Giving names and formulae to substances

(section shortcode: C10003)

Think about what you call your friends. Their full name is like the substances name and their nickname is like the substances formulae. Without these names your friends would have no idea which of them you are referring to. In the same way scientists like to have a consistent way of naming things and a short way of describing the thing being named. This helps scientists to communicate efficiently.

It is easy to describe elements and mixtures. We simply use the names that we find on the periodic table for elements and we use words to describe mixtures. But how are compounds named? In the example of iron sulphide that was used earlier, which element is named first, and which 'ending' is given to the compound name (in this case, the ending is -ide)?

The following are some guidelines for naming compounds:

1. The compound name will always include the **names of the elements** that are part of it.
   - A compound of **iron** (Fe) and *sulphur* (S) is **iron** *sulph*ide (FeS)
   - A compound of **potassium** (K) and *bromine* (Br) is **potassium** *brom*ide (KBr)
   - A compound of **sodium** (Na) and *chlorine* (Cl) is **sodium** *chlor*ide ($NaCl$)
2. In a compound, the element that is on the left of the Periodic Table, is used *first* when naming the compound. In the example of $NaCl$, sodium is a group 1 element on the left hand side of the table, while chlorine is in group 7 on the right of the table. Sodium therefore comes first in the compound name. The same is true for FeS and KBr.

3. The **symbols** of the elements can be used to represent compounds e.g. FeS, $NaCl$, KBr and H_2O. These are called **chemical formulae**. In the first three examples, the ratio of the elements in each compound is 1:1. So, for FeS, there is one atom of iron for every atom of sulphur in the compound. In the last example (H_2O) there are two atoms of hydrogen for every atom of oxygen in the compound.
4. A compound may contain **compound ions**. An ion is an atom that has lost (positive ion) or gained (negative ion) electrons. Some of the more common compound ions and their formulae are given below.

Name of compound ion	Formula
Carbonate	CO_3^{2-}
Sulphate	SO_4^{2-}
Hydroxide	OH^-
Ammonium	NH_4^+
Nitrate	NO_3^-
Hydrogen carbonate	HCO_3^-
Phosphate	PO_4^{3-}
Chlorate	ClO_3^-
Cyanide	CN^-
Chromate	CrO_4^{2-}
Permanganate	MnO_4^-

Table 2.2

5. When there are only two elements in the compound, the compound is often given a **suffix** (ending) of -ide. You would have seen this in some of the examples we have used so far. For compound ions, when a non-metal is combined with oxygen to form a negative ion (anion) which then combines with a positive ion (cation) from hydrogen or a metal, then the suffix of the name will be ...ate or ...ite. NO_3^- for example, is a negative ion, which may combine with a cation such as hydrogen (HNO_3) or a metal like potassium (KNO_3). The NO_3^- anion has the name nitr**ate**. SO_3^{2-} in a formula is sulph**ite**, e.g. sodium sulphite (Na_2SO_3). SO_4^{2-} is sulph**ate** and PO_4^{3-} is phosph**ate**.
6. **Prefixes** can be used to describe the ratio of the elements that are in the compound. You should know the following prefixes: 'mono' (one), 'di' (two) and 'tri' (three).
   - CO (carbon monoxide) - There is one atom of oxygen for every one atom of carbon
   - NO_2 (nitrogen dioxide) - There are two atoms of oxygen for every one atom of nitrogen
   - SO_3 (sulphur trioxide) - There are three atoms of oxygen for every one atom of sulphur

The above guidelines also help us to work out the formula of a compound from the name of the compound.

TIP: When numbers are written as 'subscripts' in compounds (i.e. they are written below and to the right of the element symbol), this tells us how many atoms of that element there are in relation to other elements in the compound. For example in nitrogen dioxide (NO_2) there are two oxygen atoms for every one atom of nitrogen. In sulphur trioxide (SO_3), there are three oxygen atoms for every one atom of sulphur in the compound. Later, when we start looking at chemical equations, you will notice that sometimes there are numbers *before* the compound name. For example, $2H_2O$ means that there are two molecules of water, and that in each molecule there are two hydrogen atoms for every one oxygen atom.

We can use these rules to help us name both ionic compounds and covalent compounds (more on these compounds will be covered in a later chapter). However, covalent compounds are often given other names by scientists to simplify the name. For example, if we have 2 hydrogen atoms and one oxygen atom the above naming rules would tell us that the substance is dihydrogen monoxide. But this compound is better known as water! Or if we had 1 carbon atom and 4 hydrogen atoms then the name would be carbon tetrahydride, but scientists call this compound methane.

Exercise 2.1: Naming compounds What is the chemical name for

a. $KMnO_4$
b. NH_4Cl

Solution to Exercise

Step 1. For a) we have potassium and the permanganate ion. For b) we have the ammonium ion and chlorine.
Step 2. For a) we list the potassium first and the permanganate ion second. So a) is potassium permanganate. For b) we list the ammonium ion first and change the ending of chlorine to -ide. So b) is ammonium chloride.

Exercise 2.2 Write the chemical formulae for:

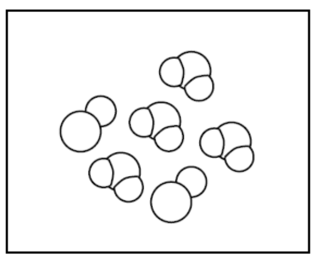
a. sodium sulphate
b. potassium chromate

Solution to Exercise

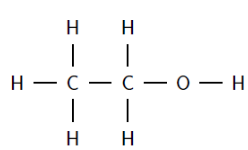
Step 1. In part a) we have Na^+ (sodium) and SO_4^{2-} (sulphate). In part b) we have K^+ (potassium) and CrO_4^{2-} (chromate)
Step 2. In part a) the charge on sodium is $+1$ and the charge on sulphate is -2, so we must have two sodiums for every sulphate. In part b) the charge on potassium is $+1$ and the charge on chromate is -2, so we must have two potassiums for every chromate.
Step 3. a) is Na_2SO_4 and b) is K_2CrO_4

1.15.1 Naming compounds

1. The formula for calcium carbonate is $CaCO_3$.
    a. Is calcium carbonate a mixture or a compound? Give a reason for your answer.
    b. What is the ratio of $Ca : C : O$ atoms in the formula?
2. Give the name of each of the following substances.
    a. KBr
    b. HCl
    c. $KMnO_4$
    d. NO_2
    e. NH_4OH
    f. Na_2SO_4
3. Give the chemical formula for each of the following compounds.
    a. potassium nitrate
    b. sodium iodide
    c. barium sulphate
    d. nitrogen dioxide
    e. sodium monosulphate
4. Refer to the diagram below, showing sodium chloride and water, and then answer the questions that follow.



    a. What is the chemical formula for water?
    b. What is the chemical formula for sodium chloride?
    c. Label the water and sodium chloride in the diagram.
    d. Give a description of the picture. Focus on whether there are elements or compounds and if it is a mixture or not.
5. What is the formula of this molecule?



    a. C_6H_2O
    b. C_2H_6O
    c. $2C6HO$
    d. $_2CH_6O$

Find the answers with the shortcodes:

(1.) llp (2.) lld (3.) llv (4.) llL (5.) llf

1.16 Metals, Metalloids and Non-metals

(section shortcode: C10004)

The elements in the Periodic Table can also be divided according to whether they are **metals**, **metalloids** or **non-metals**. On the right hand side of the Periodic Table you can draw a 'zigzag' line (This line starts with Boron (B) and goes down to Polonium (Po). This line separates all the elements that are metals from those that are non-metals. Metals are found on the left of the line, and non-metals are those on the right. Along the line you find the metalloids. You should notice that there are more metals then non-metals. Metals, metalloids and non-metals all have their own specific properties.

1.16.1 Metals

Examples of metals include copper (Cu), zinc (Zn), gold (Au) and silver (Ag). On the Periodic Table, the metals are on the left of the zig-zag line. There are a large number of elements that are metals. The following are some of the properties of metals:

- *Thermal conductors* Metals are good conductors of heat. This makes them useful in cooking utensils such as pots and pans.
- *Electrical conductors* Metals are good conductors of electricity. Metals can be used in electrical conducting wires.
- *Shiny metallic lustre* Metals have a characteristic shiny appearance and so are often used to make jewellery.
- *Malleable* This means that they can be bent into shape without breaking.
- *Ductile* Metals (such as copper) can be stretched into thin wires, which can then be used to conduct electricity.
- *Melting point* Metals usually have a high melting point and can therefore be used to make cooking pots and other equipment that needs to become very hot, without being damaged.

You can see how the properties of metals make them very useful in certain applications.

Group Work : Looking at metals

1. Collect a number of metal items from your home or school. Some examples are listed below:
   - hammer
   - wire
   - cooking pots
   - jewellery
   - nails
   - coins

2. In groups of 3-4, combine your collection of metal objects.
3. What is the function of each of these objects?
4. Discuss why you think metal was used to make each object. You should consider the properties of metals when you answer this question.

1.16.2 Non-metals

In contrast to metals, non-metals are poor thermal conductors, good electrical insulators (meaning that they do *not* conduct electrical charge) and are neither malleable nor ductile. The non-metals are found on the right hand side of the Periodic Table, and include elements such as sulphur (S), phosphorus (P), nitrogen (N) and oxygen (O).

1.16.3 Metalloids

Metalloids or semi-metals have mostly non-metallic properties. One of their distinguishing characteristics is that their conductivity increases as their temperature increases. This is the opposite of what happens in metals. The metalloids include elements such as silicon (Si) and germanium (Ge). Notice where these elements are positioned in the Periodic Table.

You should now be able to take any material and determine whether it is a metal, non-metal or metalloid simply by using its properties.

1.17 Electrical conductors, semi-conductors and insulators

(section shortcode: C10005)

An **electrical conductor** is a substance that allows an electrical current to pass through it. Electrical conductors are usually metals. *Copper* is one of the best electrical conductors, and this is why it is used to make conducting wire. In reality, *silver* actually has an even higher electrical conductivity than copper, but because silver is so expensive, it is not practical to use it for electrical wiring because such large amounts are needed. In the overhead power lines that we see above us, *aluminium* is used. The aluminium usually surrounds a steel core which adds tensile strength to the metal so that it doesn't break when it is stretched across distances. Occasionally gold is used to make wire, not because it is a particularly good conductor, but because it is very resistant to surface corrosion. *Corrosion* is when a material starts to deteriorate at the surface because of its reactions with the surroundings, for example oxygen and water in the air.

An **insulator** is a non-conducting material that does not carry any charge. Examples of insulators would be plastic and wood. Do you understand now why electrical wires are normally covered with plastic insulation? **Semi-conductors** behave like insulators when they are cold, and like conductors when they are hot. The elements silicon and germanium are examples of semi-conductors.

Definition: Conductors and insulators
A conductor allows the easy movement or flow of something such as heat or electrical charge through it. Insulators are the opposite to conductors because they *inhibit* or reduce the flow of heat, electrical charge, sound etc through them.

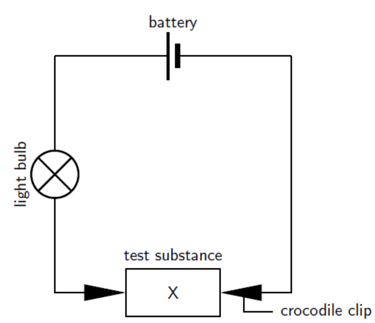
Think about the materials around you. Are they electrical conductors or not? Why are different materials used? Think about the use of semiconductors in electronics? Can you think of why they are used there?

1.17.1 Experiment : Electrical conductivity

Aim:
To investigate the electrical conductivity of a number of substances

Apparatus:

- two or three cells
- light bulb
- crocodile clips
- wire leads
- a selection of test substances (e.g. a piece of plastic, aluminium can, metal pencil sharpener, magnet, wood, chalk).



Method:

1. Set up the circuit as shown above, so that the test substance is held between the two crocodile clips. The wire leads should be connected to the cells and the light bulb should also be connected into the circuit.
2. Place the test substances one by one between the crocodile clips and see what happens to the light bulb.

Results:
Record your results in the table below:

Test substance	Metal/non-metal	Does the light bulb glow?	Conductor or insulator

Table 2.3

Conclusions:
In the substances that were tested, the metals were able to conduct electricity and the non-metals were not. Metals are good electrical conductors and non-metals are not.

The following simulation allows you to work through the above activity. For this simulation use the grab bag option to get materials to test. Set up the circuit as described in the activity.
www (Simulation: lbK)

1.18 Thermal Conductors and Insulators

www (section shortcode: C10006)

A **thermal conductor** is a material that allows energy in the form of heat, to be transferred within the material, without any movement of the material itself. An easy way to understand this concept is through a simple demonstration.

1.18.1 Demonstration : Thermal conductivity

Aim:
To demonstrate the ability of different substances to conduct heat.

Apparatus:
You will need two cups (made from the same material e.g. plastic); a metal spoon and a plastic spoon.

Method:

- Pour boiling water into the two cups so that they are about half full.
- At the same time, place a metal spoon into one cup and a plastic spoon in the other.
- Note which spoon heats up more quickly

Results:
The metal spoon heats up faster than the plastic spoon. In other words, the metal conducts heat well, but the plastic does not.

Conclusion: Metal is a good thermal conductor, while plastic is a poor thermal conductor. This explains why cooking pots are metal, but their handles are often plastic or wooden. The pot itself must be metal so that heat from the cooking surface can heat up the pot to cook the food inside it, but the handle is made from a poor thermal conductor so that the heat does not burn the hand of the person who is cooking.

An **insulator** is a material that does not allow a transfer of electricity or energy. Materials that are poor thermal conductors can also be described as being good thermal insulators.

> **NOTE:** Water is a better thermal conductor than air and conducts heat away from the body about 20 times more efficiently than air. A person who is not wearing a wetsuit, will lose heat very quickly to the water around them and can be vulnerable to hypothermia (this is when the body temperature drops very low). Wetsuits help to preserve body heat by trapping a layer of water against the skin. This water is then warmed by body heat and acts as an insulator. Wetsuits are made out of closed-cell, foam neoprene. Neoprene is a synthetic rubber that contains small bubbles of nitrogen gas when made for use as wetsuit material. Nitrogen gas has very low thermal conductivity, so it does not allow heat from the body (or the water trapped between the body and the wetsuit) to be lost to the water outside of the wetsuit. In this way a person in a wetsuit is able to keep their body temperature much higher than they would otherwise.

1.18.2 Investigation : A closer look at thermal conductivity

Look at the table below, which shows the thermal conductivity of a number of different materials, and then answer the questions that follow. The higher the number in the second column, the better the material is at conducting heat (i.e. it is a good thermal conductor). Remember that a material that conducts heat efficiently, will also lose heat more quickly than an insulating material.

Material	**Thermal Conductivity ($\mathrm{W \cdot m^{-1} \cdot K^{-1}}$)**
Silver	429
Stainless steel	16
Standard glass	1.05
Concrete	0.9 - 2
Red brick	0.69
Water	0.58
Snow	0.25 - 0.5
Wood	0.04 - 0.12
Polystyrene	0.03
Air	0.024

Table 2.4

Use this information to answer the following questions:

1. Name two materials that are good thermal conductors.
2. Name two materials that are good insulators.

3. Explain why:
   a. cooler boxes are often made of polystyrene
   b. homes that are made from wood need less internal heating during the winter months.
   c. igloos (homes made from snow) are so good at maintaining warm temperatures, even in freezing conditions.

NOTE: It is a known fact that well-insulated buildings need less energy for heating than do buildings that have no insulation. Two building materials that are being used more and more worldwide, are **mineral wool** and **polystyrene**. Mineral wool is a good insulator because it holds air still in the matrix of the wool so that heat is not lost. Since air is a poor conductor and a good insulator, this helps to keep energy within the building. Polystyrene is also a good insulator and is able to keep cool things cool and hot things hot. It has the added advantage of being resistant to moisture, mould and mildew.

Remember that concepts such as conductivity and insulation are not only relevant in the building, industrial and home environments. Think for example of the layer of blubber or fat that is found in some animals. In very cold environments, fat and blubber not only provide protection, but also act as an insulator to help the animal keep its body temperature at the right level. This is known as *thermoregulation*.

1.19 Magnetic and Non-magnetic Materials

(section shortcode: C10007)

We have now looked at a number of ways in which matter can be grouped, such as into metals, semi-metals and non-metals; electrical conductors and insulators, and thermal conductors and insulators. One way in which we can further group metals, is to divide them into those that are **magnetic** and those that are **non-magnetic.**

Definition: Magnetism
Magnetism is one of the phenomena by which materials exert attractive or repulsive forces on other materials.

A metal is said to be **ferromagnetic** if it can be magnetised (i.e. made into a magnet). If you hold a magnet very close to a metal object, it may happen that its own electrical field will be induced and the object becomes magnetic. Some metals keep their magnetism for longer than others. Look at iron and steel for example. Iron loses its magnetism quite quickly if it is taken away from the magnet. Steel on the other hand will stay magnetic for a longer time. Steel is often used to make permanent magnets that can be used for a variety of purposes.

Magnets are used to sort the metals in a scrap yard, in compasses to find direction, in the magnetic strips of video tapes and ATM cards where information must be stored, in computers and TV's, as well as in generators and electric motors.

1.19.1 Investigation : Magnetism

You can test whether an object is magnetic or not by holding another magnet close to it. If the object is attracted to the magnet, then it too is magnetic.

Find some objects in your classroom or your home and test whether they are magnetic or not. Then complete the table below:

Object	Magnetic or non-magnetic

Table 2.5

1.19.2 Group Discussion : Properties of materials

In groups of 4-5, discuss how our knowledge of the properties of materials has allowed society to:

- develop advanced computer technology
- provide homes with electricity
- find ways to conserve energy

1.20 Seperating mixtures - Not in CAPS - included for completeness

(section shortcode: C10008)

Sometimes it is important to be able to separate a mixture. There are lots of different ways to do this. These are some examples:

- *Filtration* A piece of filter paper in a funnel can be used to separate a mixture of sand and water.
- *Heating / evaporation* Heating a solution causes the liquid (normally water) to evaporate, leaving the other (solid) part of the mixture behind. You can try this using a salt solution.
- *Centrifugation* This is a laboratory process which uses the centrifugal force of spinning objects to separate out the heavier substances from a mixture. This process is used to separate the cells and plasma in blood. When the test tubes that hold the blood are spun round in the machine, the heavier cells sink to the bottom of the test tube. Can you think of a reason why it might be important to have a way of separating blood in this way?
- *Dialysis* This is an interesting way of separating a mixture because it can be used in some important applications. Dialysis works using a process called *diffusion*. Diffusion takes place when one substance

in a mixture moves from an area where it has a high concentration to an area where its concentration is lower. When this movement takes place across a semi-permeable membrane it is called osmosis. A semi-permeable membrane is a barrier that lets some things move across it, but not others. This process is very important for people whose kidneys are not functioning properly, an illness called *renal failure*.

NOTE: Normally, healthy kidneys remove waste products from the blood. When a person has renal failure, their kidneys cannot do this any more, and this can be life-threatening. Using dialysis, the blood of the patient flows on one side of a semi-permeable membrane. On the other side there will be a fluid that has no waste products but lots of other important substances such as potassium ions (K^+) that the person will need. Waste products from the blood diffuse from where their concentration is high (i.e. in the person's blood) into the 'clean' fluid on the other side of the membrane. The potassium ions will move in the opposite direction from the fluid into the blood. Through this process, waste products are taken out of the blood so that the person stays healthy.

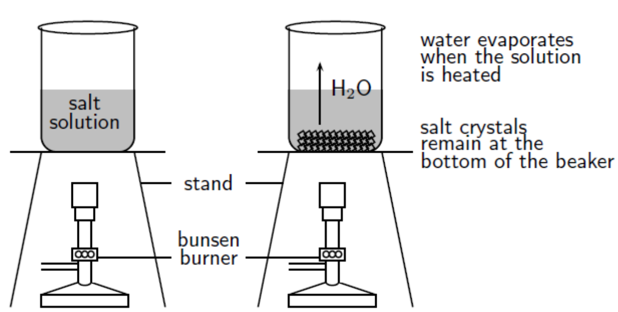
1.20.1 Investigation : The separation of a salt solution

Aim:
To demonstrate that a homogeneous salt solution can be separated using physical methods.

Apparatus:
glass beaker, salt, water, retort stand, bunsen burner.

Method:

1. Pour a small amount of water (about 20 ml) into a beaker.
2. Measure a teaspoon of salt and pour this into the water.
3. Stir until the salt dissolves completely. This is now called a *salt solution*. This salt solution is a homogeneous mixture.
4. Place the beaker on a retort stand over a bunsen burner and heat gently. You should increase the heat until the water almost boils.
5. Watch the beaker until all the water has evaporated. What do you see in the beaker?



Results:
The water evaporates from the beaker and tiny grains of salt remain at the bottom. (You may also observe grains of salt on the walls of the beaker.)

Conclusion:
The salt solution, which is a homogeneous mixture of salt and water, has been separated using heating and evaporation.

1.20.2 Discussion : Separating mixtures

Work in groups of 3-4

Imagine that you have been given a container which holds a mixture of sand, iron filings (small pieces of iron metal), salt and small stones of different sizes. Is this a homogeneous or a heterogeneous mixture? In your group, discuss how you would go about separating this mixture into the four materials that it contains. The following presentation provides a summary of the classification of matter.

www (Presentation: P10009)

1.21 Summary

www (section shortcode: C10010)

- All the objects and substances that we see in the world are made of **matter**.
- This matter can be classified according to whether it is a **mixture** or a **pure substance**.
- A **mixture** is a combination of one or more substances that are not chemically bonded to each other. Examples of mixtures are air (a mixture of different gases) and blood (a mixture of cells, platelets and plasma).
- The main **characteristics** of mixtures are that the substances that make them up are not in a fixed ratio, they keep their individual properties and they can be separated from each other using mechanical means.
- A **heterogeneous mixture** is non-uniform and the different parts of the mixture can be seen. An example would be a mixture of sand and water.
- A **homogeneous mixture** is uniform, and the different components of the mixture can't be seen. An example would be a salt solution. A salt solution is a mixture of salt and water. The salt dissolves in the water, meaning that you can't see the individual salt particles. They are interspersed between the water molecules. Another example is a metal **alloy** such as steel.
- Mixtures can be **separated** using a number of methods such as filtration, heating, evaporation, centrifugation and dialysis.
- Pure substances can be further divided into **elements** and **compounds**.
- An **element** is a substance that can't be broken down into simpler substances through chemical means.
- All the elements are recorded in the **Periodic Table of the Elements**. Each element has its own chemical symbol. Examples are iron (Fe), sulphur (S), calcium (Ca), magnesium (Mg) and fluorine (F).
- A **compound** is a substance that is made up of two or more elements that are chemically bonded to each other in a fixed ratio. Examples of compounds are sodium chloride (NaCl), iron sulphide (FeS), calcium carbonate ($\mathrm{CaCO_3}$) and water ($\mathrm{H_2O}$).
- When **naming compounds** and writing their **chemical formula**, it is important to know the elements that are in the compound, how many atoms of each of these elements will combine in the compound and where the elements are in the Periodic Table. A number of rules can then be followed to name the compound.
- Another way of classifying matter is into **metals** (e.g. iron, gold, copper), **semi-metals** (e.g. silicon and germanium) and **non-metals** (e.g. sulphur, phosphorus and nitrogen).

- **Metals** are good electrical and thermal conductors, they have a shiny lustre, they are malleable and ductile, and they have a high melting point. These properties make metals very useful in electrical wires, cooking utensils, jewellery and many other applications.
- A further way of classifying matter is into **electrical conductors**, **semi-conductors** and **insulators**.
- An **electrical conductor** allows an electrical current to pass through it. Most metals are good electrical conductors.
- An **electrical insulator** is not able to carry an electrical current. Examples are plastic, wood, cotton material and ceramic.
- Materials may also be classified as **thermal conductors** or **thermal insulators** depending on whether or not they are able to conduct heat.
- Materials may also be either **magnetic** or **non-magnetic**.

1.21.1 Summary

1. For each of the following **multiple choice** questions, choose *one* correct answer from the list provided.
   a. Which of the following can be classified as a mixture:
      a. sugar
      b. table salt
      c. air
      d. iron
   b. An element can be defined as:
      a. A substance that cannot be separated into two or more substances by ordinary chemical (or physical) means
      b. A substance with constant composition
      c. A substance that contains two or more substances, in definite proportion by weight
      d. A uniform substance
2. Classify each of the following substances as an *element*, a *compound*, a *solution* (homogeneous mixture), or a *heterogeneous mixture*: salt, pure water, soil, salt water, pure air, carbon dioxide, gold and bronze.

3. Look at the table below. In the first column (A) is a list of substances. In the second column (B) is a description of the group that each of these substances belongs in. Match up the *substance* in Column A with the *description* in Column B.

Column A	**Column B**
iron	a compound containing 2 elements
H_2S	a heterogeneous mixture
sugar solution	a metal alloy
sand and stones	an element
steel	a homogeneous mixture

Table 2.6

4. You are given a test tube that contains a mixture of iron filings and sulphur. You are asked to weigh the amount of iron in the sample.
   a. Suggest one method that you could use to separate the iron filings from the sulphur.
   b. What property of metals allows you to do this?

5. Given the following descriptions, write the chemical formula for each of the following substances:

- **a.** silver metal
- **b.** a compound that contains only potassium and bromine
- **c.** a gas that contains the elements carbon and oxygen in a ratio of 1:2

6. Give the names of each of the following compounds:

- **a.** $NaBr$
- **b.** $BaSO_4$
- **c.** SO_2

7. For each of the following materials, say what properties of the material make it important in carrying out its particular function.

- **a.** **tar** on roads
- **b.** **iron** burglar bars
- **c.** **plastic** furniture
- **d.** **metal** jewellery
- **e.** **clay** for building
- **f.** **cotton** clothing

Find the answers with the shortcodes:

(1.) ll6 (2.) llF (3.) llG (4.) ll7 (5.) llA (6.) llo
(7.) lls (8.) llH

States of matter and the kinetic molecular theory

1.22 Introduction

(section shortcode: C10011)

In this chapter we will explore the states of matter and then look at the kinetic molecular theory. Matter exists in three states: solid, liquid and gas. We will also examine how the kinetic theory of matter helps explain boiling and melting points as well as other properties of matter.

NOTE: When a gas is heated above a certain temperature the electrons in the atoms start to leave the atoms. The gas is said to be ionised. When a gas is ionised it is known as a plasma. Plasmas share many of the properties of gases (they have no fixed volume and fill the space they are in). This is a very high energy state and plasmas often glow. Ionisation is the process of moving from a gas to a plasma and deionisation is the reverse process. We will not consider plasmas further in this chapter.

1.23 States of matter

(section shortcode: C10012)

All matter is made up of particles. We can see this when we look at diffusion. Diffusion is the movement of particles from a high concentration to a low concentration. Diffusion can be seen as a spreading out of particles resulting in an even distribution of the particles. You can see diffusion when you place a drop of food colouring in water. The colour slowly spreads out through the water. If matter were not made of particles then we would only see a clump of colour when we put the food colouring in water, as there would be nothing that could move about and mix in with the water. The composition of matter will be looked at in What are the objects around us made of? (Chapter 6).

Diffusion is a result of the constant thermal motion of particles. In Section 3.3 (The Kinetic Theory of Matter) we will talk more about the thermal motion of particles.

In 1828 Robert Brown observed that pollen grains suspended in water moved about in a rapid, irregular motion. This motion has since become known as Brownian motion. Brownian motion is essentially diffusion of many particles.

Matter exists in one of three states, namely solid, liquid and gas. Matter can change between these states by either adding heat or removing heat. This is known as a change of state. As we heat an object (e.g. water) it goes from a solid to a liquid to a gas. As we cool an object it goes from a gas to a liquid to a solid. The changes of state that you should know are:

- Melting is the process of going from solid to liquid.
- Boiling (or evaporation) is the process of going from liquid to gas.
- Freezing is the process of going from liquid to solid.
- Condensation is the process of going from gas to liquid.
- Occasionally (e.g. for carbon dioxide) we can go directly from solid to gas in a process called sublimation.

A solid has a fixed shape and volume. A liquid takes on the shape of the container that it is in. A gas completely fills the container that it is in. See Section 2.3 (The Kinetic Theory of Matter) for more on changes of state.

If we know the melting and boiling point of a substance then we can say what state (solid, liquid or gas) it will be in at any temperature.

1.23.1 Experiment: States of matter

Aim To investigate the heating and cooling curve of water.

Apparatus beakers, ice, bunsen burner, thermometer, water.

Method

- Place some ice in a beaker
- Measure the temperature of the ice and record it.
- After 10 s measure the temperature again and record it. Repeat every 10 s, until at least 1 minute after the ice has melted.
- Heat some water in a beaker until it boils. Measure and record the temperature of the water.
- Remove the water from the heat and measure the temperature every 10 s, until the beaker is cool to touch

WARNING: Be careful when handling the beaker of hot water. Do not touch the beaker with your hands, you will burn yourself.

Results Record your results in the following table:

Temperature of ice	Time (s)	Temperature of water	Time (s)

Table 3.1: Table of results

Plot a graph of temperature against time for the ice melting and the boiling water cooling.

Discussion and conclusion Discuss your results with others in your class. What conclusions can you draw? You should find that the temperature of the ice increases until the first drops of liquid appear and then the temperature remains the same, until all the ice is melted. You should also find that when you cool water down from boiling, the temperature remains constant for a while, then starts decreasing.

In the above experiment, you investigated the heating and cooling curves of water. We can draw heating and cooling curves for any substance. A heating curve of a substance gives the changes in temperature as we move from a solid to a liquid to a gas. A cooling curve gives the changes in temperature as we move from gas to liquid to solid. An important observation is that as a substance melts or boils, the temperature remains constant until the substance has changed state. This is because all the heat energy goes into breaking or forming the forces between the molecules.

The above experiment is one way of demonstrating the changes of state of a substance. Ice melting or water boiling should be very familiar to you.

1.24 The Kinetic Theory of Matter

(section shortcode: C10013)

The **kinetic theory of matter** helps us to explain why matter exists in different *phases* (i.e. solid, liquid and gas), and how matter can change from one phase to the next. The kinetic theory of matter also helps us to understand other properties of matter. It is important to realise that what we will go on to describe is only a *theory*. It cannot be proved beyond doubt, but the fact that it helps us to explain our observations of changes in phase, and other properties of matter, suggests that it probably is more than just a theory.

Broadly, the Kinetic Theory of Matter says that:

- Matter is made up of **particles** that are constantly moving.
- All particles have **energy**, but the energy varies depending on whether the substance is a solid, liquid or gas. Solid particles have the least amount of energy and gas particles have the greatest amount of energy.

- The **temperature** of a substance is a measure of the *average kinetic energy* of the particles.
- A change in **phase** may occur when the energy of the particles is changed.
- There are **spaces** between the particles of matter.
- There are **attractive forces** between particles and these become stronger as the particles move closer together. These attractive forces will either be intramolecular forces (if the particles are atoms) or intermolecular forces (if the particles are molecules). When the particles are extremely close, repulsive forces start to act.

Table 3.2 summarises the characteristics of the particles that are in each phase of matter.

Property of matter	**Solid**	**Liquid**	**Gas**
Particles	Atoms or molecules	Atoms or molecules	Atoms or molecules
Energy and movement of particles	Low energy - particles vibrate around a fixed point	Particles have less energy than in the gas phase	Particles have high energy and are constantly moving
Spaces between particles	Very little space between particles. Particles are tightly packed together	Smaller spaces than in gases, but larger spaces than in solids	Large spaces because of high energy
Attractive forces between particles	Very strong forces. Solids have a fixed volume.	Stronger forces than in gas. Liquids can be poured.	Weak forces because of the large distance between particles
Changes in phase	Solids become liquids if their temperature is increased. In some cases a solid may become a gas if the temperature is increased.	A liquid becomes a gas if its temperature is increased. It becomes a solid if its temperature decreases.	In general a gas becomes a liquid when it is cooled. (In a few cases a gas becomes a solid when cooled). Particles have less energy and therefore move closer together so that the attractive forces become stronger, and the gas becomes a liquid (or a solid.)

Table 3.2: Table summarising the general features of solids, liquids and gases.

The following presentation is a brief summary of the above. Try to fill in the blank spaces before clicking onto the next slide.

(Presentation: P10014)

Let's look at an example that involves the three phases of water: ice (solid), water (liquid) and water vapour (gas). Note that in the Figure 3.2 below the molecules in the solid phase are represented by single spheres, but they would in reality look like the molecules in the liquid and gas phase. Sometimes we represent molecules as single spheres in the solid phase to emphasise the small amount of space between them and to make the drawing simpler.

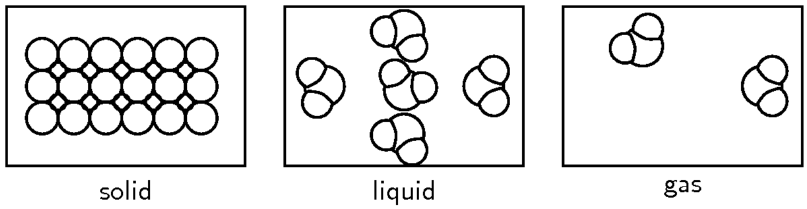


Figure 3.2: The three phases of matter

Taking water as an example we find that in the solid phase the water molecules have very little energy and can't move away from each other. The molecules are held closely together in a regular pattern called a *lattice*. If the ice is heated, the energy of the molecules increases. This means that some of the water molecules are able to overcome the intermolecular forces that are holding them together, and the molecules move further apart to form *liquid water*. This is why liquid water is able to flow, because the molecules are more free to move than they were in the solid lattice. If the molecules are heated further, the liquid water will become water vapour, which is a gas. Gas particles have lots of energy and are far away from each other. That is why it is difficult to keep a gas in a specific area! The attractive forces between the particles are very weak and they are only loosely held together. Figure 3.3 shows the changes in phase that may occur in matter, and the names that describe these processes.

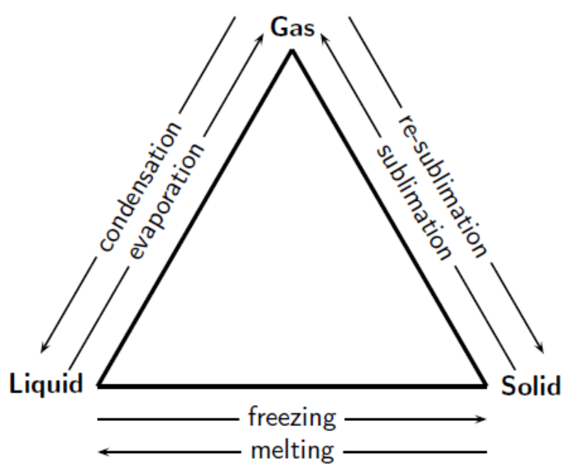


Figure 3.3: Changes in phase

1.25 Intramolecular and intermolecular forces (Not in CAPS - Included for Completeness)

(section shortcode: C10015)

When atoms join to form molecules, they are held together by **chemical bonds**. The type of bond, and the strength of the bond, depends on the atoms that are involved. These bonds are called **intramolecular forces** because they are bonding forces *inside* a molecule ('intra' means 'within' or 'inside'). Sometimes we simply call these intramolecular forces chemical bonds.

Definition: Intramolecular force
The force between the atoms of a molecule, which holds them together.

Examples of the types of chemical bonds that can exist between atoms inside a molecule are shown below. These will be looked at in more detail in Chemical bonding (Chapter 5).

- *Covalent bond* Covalent bonds exist between non-metal atoms e.g. There are covalent bonds between the carbon and oxygen atoms in a molecule of carbon dioxide.
- *Ionic bond* Ionic bonds occur between non-metal and metal atoms e.g. There are ionic bonds between the sodium and chlorine atoms in a molecule of sodium chloride.
- *Metallic bond* Metallic bonds join metal atoms e.g. There are metallic bonds between copper atoms in a piece of copper metal.

Intermolecular forces are those bonds that hold *molecules* together. A glass of water for example, contains many molecules of water. These molecules are held together by intermolecular forces. The strength of the intermolecular forces is important because they affect properties such as *melting point* and *boiling point*. For example, the stronger the intermolecular forces, the higher the melting point and boiling point for that substance. The strength of the intermolecular forces increases as the size of the molecule increases.

Definition: Intermolecular force
A force between molecules, which holds them together.

The following diagram may help you to understand the difference between intramolecular forces and intermolecular forces.

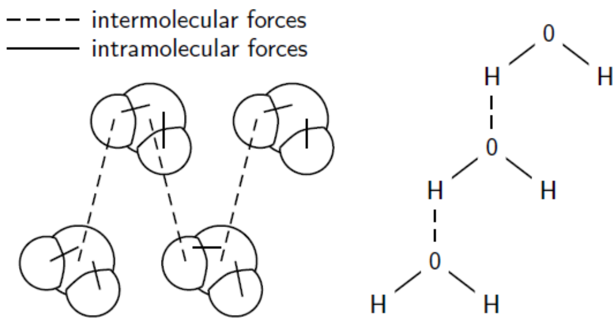


Figure 3.4: Two representations showing the intermolecular and intramolecular forces in water: space-filling model and structural formula.

It should be clearer now that there are two types of forces that hold matter together. In the case of water, there are intramolecular forces that hold the two hydrogen atoms to the oxygen atom *in each molecule of water* (these are the solid lines in the above diagram). There are also intermolecular forces *between each of these water molecules*. These intermolecular forces join the hydrogen atom from one molecule to the oxygen atom of **another molecule** (these are the dashed lines in the above figure). As mentioned earlier, these forces are very important because they affect many of the *properties of matter* such as boiling point, melting point and a number of other properties. Before we go on to look at some of these examples, it is important that we first take a look at the **Kinetic Theory of Matter**.

TIP: To help you remember that intermolecular means between molecules, remember that international means between nations.

1.25.1 Intramolecular and intermolecular forces

1. Using ammonia gas as an example...
   a. Explain what is meant by an *intramolecular* force or *chemical bond*.
   b. Explain what is meant by an *intermolecular* force.
2. Draw a diagram showing three molecules of carbon dioxide. On the diagram, show where the intramolecular and intermolecular forces are.
3. Why is it important to understand the types of forces that exist between atoms and between molecules? Try to use some practical examples in your answer.

Find the answers with the shortcodes:

(1.) lin (2.) liQ (3.) liU

1.26 The Properties of Matter

(section shortcode: C10016)

Let us now look at what we have learned about chemical bonds, intermolecular forces and the kinetic theory of matter, and see whether this can help us to understand some of the macroscopic properties of materials.

1. **Melting point**

   Definition: Melting point
   The temperature at which a *solid* changes its phase or state to become a *liquid*. The process is called melting and the reverse process (change in phase from liquid to solid) is called **freezing**.

   In order for a solid to melt, the energy of the particles must increase enough to overcome the bonds that are holding the particles together. It makes sense then that a solid which is held together by strong bonds will have a *higher* melting point than one where the bonds are weak, because more energy (heat) is needed to break the bonds. In the examples we have looked at metals, ionic solids and some atomic lattices (e.g. diamond) have high melting points, whereas the melting points for molecular solids and other atomic lattices (e.g. graphite) are much lower. Generally, the intermolecular forces between molecular solids are *weaker* than those between ionic and metallic solids.
2. **Boiling point**

   Definition: Boiling point
   The temperature at which a *liquid* changes its phase to become a *gas*. The process is called evaporation and the reverse process is called condensation

   When the temperature of a liquid increases, the average kinetic energy of the particles also increases and they are able to overcome the bonding forces that are holding them in the liquid. When boiling point is reached, *evaporation* takes place and some particles in the liquid become a gas. In other words, the energy of the particles is too great for them to be held in a liquid anymore. The stronger the bonds within a liquid, the higher the boiling point needs to be in order to break these bonds. Metallic and ionic compounds have high boiling points while the boiling point for molecular liquids is lower. The data in Table 3.3 below may help you to understand some of the concepts we have explained. Not all of the substances in the table are solids at room temperature, so for now, let's just focus on the *boiling points* for each of these substances. What do you notice?

Substance	Melting point ($^\circ\mathrm{C}$)	Boiling point ($^\circ\mathrm{C}$)
Ethanol (C_2H_6O)	- 114,3	78,4
Water	0	100
Mercury	-38,83	356,73
Sodium chloride	801	1465

Table 3.3: The melting and boiling points for a number of substances

You will have seen that substances such as ethanol, with relatively weak intermolecular forces, have the lowest boiling point, while substances with stronger intermolecular forces such as sodium chloride and mercury, must be heated much more if the particles are to have enough energy to overcome the forces that are holding them together in the liquid. See the section (Section 2.5.1: Exercise: Forces and boiling point) below for a further exercise on boiling point.

3. **Density and viscosity** The density of a solid is generally higher than that of a liquid because the particles are held much more closely together and therefore there are more particles packed together in a particular volume. In other words, there is a greater mass of the substance in a particular volume. In general, density increases as the strength of the intermolecular forces increases.

Definition: Density
Density is a measure of the mass of a substance per unit volume.

Definition: Viscosity
Viscosity is a measure of how resistant a liquid is to flowing (in other words, how easy it is to pour the liquid from one container to another).

Viscosity is also sometimes described as the 'thickness' of a fluid. Think for example of syrup and how slowly it pours from one container into another. Now compare this to how easy it is to pour water. The viscosity of syrup is greater than the viscosity of water. Once again, the stronger the intermolecular forces in the liquid, the greater its viscosity.

It should be clear now that we can explain a lot of the **macroscopic properties** of matter (i.e. the characteristics we can see or observe) by understanding their **microscopic structure** and the way in which the atoms and molecules that make up matter are held together.

1.26.1 Exercise: Forces and boiling point

The table below gives the molecular formula and the boiling point for a number of organic compounds called *alkanes* (more on these compounds in grade 12). Refer to the table and then answer the questions that follow.

Organic compound	Molecular formula	Boiling point ($^\circ C$)
Methane	CH_4	-161.6
Ethane	C_2H_6	- 88.6
Propane	C_3H_8	-45
Butane	C_4H_{10}	-0.5
Pentane	C_5H_{12}	36.1
Hexane	C_6H_{14}	69
Heptane	C_7H_{16}	98.42
Octane	C_8H_{18}	125.52

Table 3.4

www Data from: http://www.wikipedia.com

1. Draw a graph to show the relationship between the number of carbon atoms in each alkane and its boiling point. (Number of carbon atoms will go on the x-axis and boiling point on the y-axis).
2. Describe what you see.
3. Suggest a reason for what you have observed.
4. Why was it enough for us to use 'number of carbon atoms' as a measure of the molecular weight of the molecules?

www Find the answers with the shortcodes: (1.) liP

1.26.2 Investigation : Determining the density of liquids:

Density is a very important property because it helps us to identify different materials. Every material, depending on the elements that make it up and the arrangement of its atoms, will have a different density.

The equation for density is:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \tag{3.1}$$

Discussion questions:
To calculate the density of liquids and solids, we need to be able to first determine their mass and volume. As a group, think about the following questions:

- How would you determine the mass of a liquid?
- How would you determine the volume of an irregular solid?

Apparatus:
Laboratory mass balance, $10\ \text{ml}$ and $100\ \text{ml}$ graduated cylinders, thread, distilled water, two different liquids.

Method:
Determine the density of the distilled water and two liquids as follows:

1. Measure and record the mass of a 10 ml graduated cyclinder.
2. Pour an amount of distilled water into the cylinder.
3. Measure and record the combined mass of the water and cylinder.
4. Record the volume of distilled water in the cylinder
5. Empty, clean and dry the graduated cylinder.
6. Repeat the above steps for the other two liquids you have.
7. Complete the table below.

Liquid	**Mass (g)**	**Volume (ml)**	**Density ($\text{g} \cdot \text{ml}^{-1}$)**
Distilled water			
Liquid 1			
Liquid 2			

Table 3.5

1.26.3 Investigation : Determining the density of irregular solids:

Apparatus:
Use the same materials and equpiment as before (for the liquids). Also find a number of solids that have an irregular shape.

Method:
Determine the density of irregular solids as follows:

1. Measure and record the mass of one of the irregular solids.
2. Tie a piece of thread around the solid.
3. Pour some water into a 100 ml graduated cylinder and record the volume.
4. Gently lower the solid into the water, keeping hold of the thread. Record the combined volume of the solid and the water.
5. Determine the volume of the solid by subtracting the combined volume from the original volume of the water only.
6. Repeat these steps for the second object.
7. Complete the table below.

Solid	**Mass (g)**	**Volume (ml)**	**Density ($\text{g} \cdot \text{ml}^{-1}$)**
Solid 1			
Solid 2			
Solid 3			

Table 3.6

The following presentation provides a summary of the work covered in this chapter.

(Presentation: P10017)

1.27 Summary

(section shortcode: C10018)

- There are three states of matter: solid, liquid and gas.
- Diffusion is the movement of particles from a high concentration to a low concentration. Brownian motion is the diffusion of many particles.
- The **kinetic theory of matter** attempts to explain the behaviour of matter in different phases.
- The kinetic theory of matter says that all matter is composed of **particles** which have a certain amount of **energy** which allows them to **move** at different speeds depending on the temperature (energy). There are **spaces** between the particles and also **attractive forces** between particles when they come close together.
- Intramolecular force is the force between the atoms of a molecule, which holds them together. Intermolecular force is a force between molecules, which holds them together.
- Understanding chemical bonds, intermolecular forces and the kinetic theory of matter can help to explain many of the **macroscopic properties** of matter.
- **Melting point** is the temperature at which a *solid* changes its phase to become a *liquid*. The reverse process (change in phase from liquid to solid) is called **freezing**. The stronger the chemical bonds and intermolecular forces in a substance, the higher the melting point will be.
- **Boiling point** is the temperature at which a liquid changes phase to become a gas. The reverse process (change in phase from gas to liquid) is called **condensing**. The stronger the chemical bonds and intermolecular forces in a substance, the higher the boiling point will be.
- **Density** is a measure of the mass of a substance per unit volume.
- **Viscosity** is a measure of how resistant a liquid is to flowing.

1.28 End of chapter exercises

(section shortcode: C10019)

1. Give one word or term for each of the following descriptions.
   - **a.** The property that determines how easily a liquid flows.
   - **b.** The change in phase from liquid to gas.
2. If one substance A has a melting point that is *lower* than the melting point of substance B, this suggests that...
   - **a.** A will be a liquid at room temperature.
   - **b.** The chemical bonds in substance A are weaker than those in substance B.
   - **c.** The chemical bonds in substance A are stronger than those in substance B.
   - **d.** B will be a gas at room temperature.

3. Boiling point is an important concept to understand.
   a. Define 'boiling point'.
   b. What change in phase takes place when a liquid reaches its boiling point?
   c. What is the boiling point of water?
   d. Use the kinetic theory of matter and your knowledge of intermolecular forces to explain why water changes phase at this temperature.
4. Describe a solid in terms of the kinetic molecular theory.
5. Refer to the table below which gives the melting and boiling points of a number of elements and then answer the questions that follow. (*Data from http://www.chemicalelements.com*)

Element	**Melting point**	**Boiling point ($^\circ$C)**
copper	1083	2567
magnesium	650	1107
oxygen	-218,4	-183
carbon	3500	4827
helium	-272	-268,6
sulphur	112,8	444,6

Table 3.7

   a. What state of matter (i.e. solid, liquid or gas) will each of these elements be in at room temperature?
   b. Which of these elements has the strongest forces between its atoms? Give a reason for your answer.
   c. Which of these elements has the weakest forces between its atoms? Give a reason for your answer.

Find the answers with the shortcodes:

(1.) l2t (2.) lip (3.) lim (4.) lgf (5.) liy

The Atom

1.29 Introduction

(section shortcode: C10020)

The following video covers some of the properties of an atom.

Veritasium video on the atom - 1 www (Video: P10021)

We have now looked at many examples of the types of matter and materials that exist around us and we have investigated some of the ways that materials are classified. But what is it that makes up these materials? And what makes one material different from another? In order to understand this, we need to take a closer look at the building block of matter - the **atom**. Atoms are the basis of all the structures and organisms in the universe. The planets, sun, grass, trees, air we breathe and people are all made up of different combinations of atoms.

1.30 Project: Models of the atom

www (section shortcode: C10022)

Our current understanding of the atom came about over a long period of time, with many different people playing a role. Conduct some research into the development of the different ideas of the atom and the people who contributed to it. Some suggested people to look at are: JJ Thomson, Ernest Rutherford, Marie Curie, JC Maxwell, Max Planck, Albert Einstein, Niels Bohr, Lucretius, LV de Broglie, CJ Davisson, LH Germer, Chadwick, Werner Heisenberg, Max Born, Erwin Schrodinger, John Dalton, Empedocles, Leucippus, Democritus, Epicurus, Zosimos, Maria the Jewess, Geber, Rhazes, Robert Boyle, Henry Cavendish, A Lavoisier and H Becquerel. You do not need to find information on all these people, but try to find information about as many of them as possible.

Make a list of the key contributions to a model of the atom that each of these people made and then make a timeline of this information. (You can use an online tool such as Dipity[1] to make a timeline.) Try to get a feel for how it all eventually fit together into the modern understanding of the atom.

1.31 Models of the Atom

(section shortcode: C10023)

It is important to realise that a lot of what we know about the structure of atoms has been developed over a long period of time. This is often how scientific knowledge develops, with one person building on the ideas of someone else. We are going to look at how our modern understanding of the atom has evolved over time.

The idea of atoms was invented by two Greek philosophers, Democritus and Leucippus in the fifth century BC. The Greek word $\alpha\tau o\mu o\nu$ (atom) means **indivisible** because they believed that atoms could not be broken into smaller pieces.

Nowadays, we know that atoms are made up of a **positively charged nucleus** in the centre surrounded by **negatively charged electrons**. However, in the past, before the structure of the atom was properly understood, scientists came up with lots of different **models** or **pictures** to describe what atoms look like.

Definition: Model
A model is a representation of a system in the real world. Models help us to understand systems and their properties. For example, an *atomic model* represents what the structure of an atom *could* look like, based on what we know about how atoms behave. It is not necessarily a true picture of the exact structure of an atom.

1.31.1 The Plum Pudding Model

After the electron was discovered by J.J. Thomson in 1897, people realised that atoms were made up of even smaller particles than they had previously thought. However, the atomic nucleus had not been discovered yet and so the 'plum pudding model' was put forward in 1904. In this model, the atom is made up of negative electrons that float in a soup of positive charge, much like plums in a pudding or raisins in a fruit cake (Figure 4.2). In 1906, Thomson was awarded the Nobel Prize for his work in this field. However, even with the Plum Pudding Model, there was still no understanding of how these electrons in the atom were arranged.

[1] http://www.dipity.com/

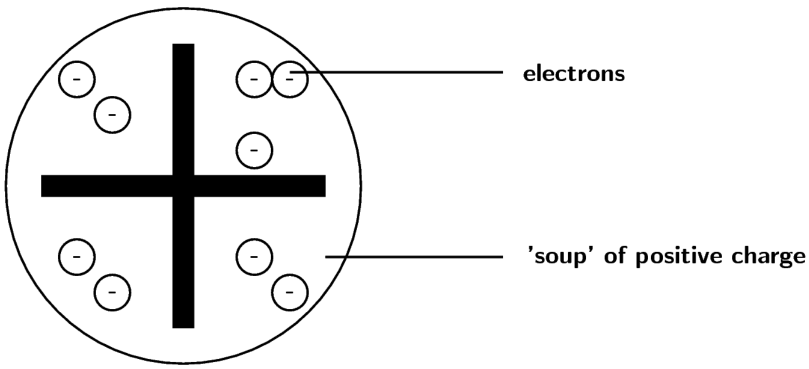


Figure 4.2: A schematic diagram to show what the atom looks like according to the Plum Pudding model

The discovery of **radiation** was the next step along the path to building an accurate picture of atomic structure. In the early twentieth century, Marie Curie and her husband Pierre, discovered that some elements (the *radioactive* elements) emit particles, which are able to pass through matter in a similar way to X-rays (read more about this in Grade 11). It was Ernest Rutherford who, in 1911, used this discovery to revise the model of the atom.

1.31.2 Rutherford's model of the atom

Radioactive elements emit different types of particles. Some of these are positively charged alpha (α) particles. Rutherford carried out a series of experiments where he bombarded sheets of gold foil with these particles, to try to get a better understanding of where the positive charge in the atom was. A simplified diagram of his experiment is shown in Figure 4.3.

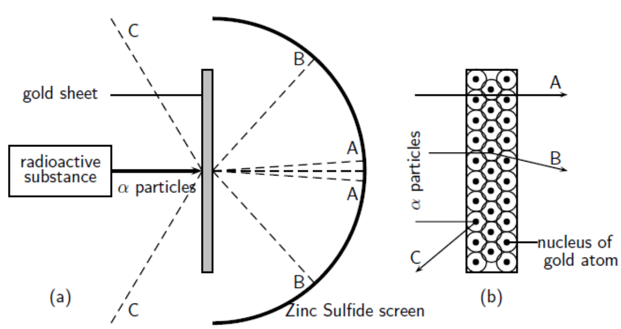


Figure 4.3: Rutherford's gold foil experiment. Figure (a) shows the path of the α particles after they hit the gold sheet. Figure (b) shows the arrangement of atoms in the gold sheets and the path of the α particles in relation to this.

Rutherford set up his experiment so that a beam of alpha particles was directed at the gold sheets. Behind the gold sheets was a screen made of zinc sulphide. This screen allowed Rutherford to see where the alpha particles were landing. Rutherford knew that the *electrons* in the gold atoms would not really affect the path of the alpha particles, because the mass of an electron is so much smaller than that of a proton. He reasoned that the positively charged *protons* would be the ones to *repel* the positively charged alpha particles and alter their path.

What he discovered was that most of the alpha particles passed through the foil undisturbed and could be detected on the screen directly behind the foil (A). Some of the particles ended up being slightly deflected onto other parts of the screen (B). But what was even more interesting was that some of the particles were deflected straight back in the direction from where they had come (C)! These were the particles that had been repelled by the positive protons in the gold atoms. If the Plum Pudding model of the atom were true then Rutherford would have expected much more repulsion, since the positive charge according to that model is distributed throughout the atom. But this was not the case. The fact that most particles passed straight through suggested that the positive charge was concentrated in one part of the atom only.

Rutherford's work led to a change in ideas around the atom. His new model described the atom as a tiny, dense, positively charged core called a nucleus surrounded by lighter, negatively charged electrons. Another way of thinking about this model was that the atom was seen to be like a mini solar system where the electrons orbit the nucleus like planets orbiting around the sun. A simplified picture of this is shown in Figure 4.4. This model is sometimes known as the planetary model of the atom.

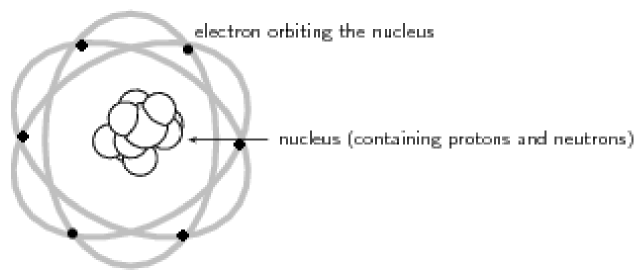


Figure 4.4: Rutherford's model of the atom

1.31.3 The Bohr Model

There were, however, some problems with this model: for example it could not explain the very interesting observation that atoms only emit light at certain wavelengths or frequencies. Niels Bohr solved this problem by proposing that the electrons could only orbit the nucleus in certain special orbits at different energy levels around the nucleus. The exact energies of the orbitals in each energy level depends on the type of atom. Helium for example, has different energy levels to Carbon. If an electron jumps down from a higher energy level to a lower energy level, then light is emitted from the atom. The energy of the light emitted is the same as the gap in the energy between the two energy levels. You can read more about this in "Energy quantisation and electron configuration" (Section 3.8: Electron configuration). The distance between the nucleus and the electron in the lowest energy level of a hydrogen atom is known as the **Bohr radius**.

NOTE: Light has the properties of both a particle **and** a wave! Einstein discovered that light comes in energy packets which are called **photons**. When an electron in an atom changes energy levels, a photon of light is emitted. This photon has the same energy as the difference between the two electron energy levels.

1.31.4 Other models of the atom

Although the most common model of the atom is the Bohr model, scientists have not stopped thinking about other ways to describe atoms. One of the most important contributions to atomic theory (the field of science that looks at atoms) was the development of quantum theory. Schrodinger, Heisenberg, Born and many others have had a role in developing quantum theory. The description of an atom by quantum theory is very complex and is only covered at university level.

1.31.5 Models of the atom

Match the information in column A, with the key discoverer in column B.

Column A	Column B
Discovery of electrons and the plum pudding model	Niels Bohr
Arrangement of electrons	Marie Curie and her husband, Pierre
Atoms as the smallest building block of matter	Ancient Greeks
Discovery of the nucleus	JJ Thomson
Discovery of radiation	Rutherford

Table 4.1

Find the answers with the shortcodes:

(1.) l4g

1.32 The size of atoms

(section shortcode: C10024)

It is difficult sometimes to imagine the size of an atom, or its mass, because we cannot see an atom and also because we are not used to working with such small measurements.

1.32.1 How heavy is an atom?

It is possible to determine the mass of a single atom in kilograms. But to do this, you would need very modern *mass spectrometers* and the values you would get would be very clumsy and difficult to use. The mass of a carbon atom, for example, is about $1,99 \times 10^{-26}\,\mathrm{kg}$, while the mass of an atom of hydrogen is about $1,67 \times 10^{-27}\,\mathrm{kg}$. Looking at these very small numbers makes it difficult to compare how much bigger the mass of one atom is when compared to another.

To make the situation simpler, scientists use a different unit of mass when they are describing the mass of an atom. This unit is called the **atomic mass unit** (amu). We can abbreviate (shorten) this unit to just 'u'. Scientists use the **carbon standard** to determine amu. The carbon standard assigns carbon an atomic mass of 12 u. Using the carbon standard the mass of an atom of hydrogen will be 1 u. You can check this by dividing the mass of a carbon atom in kilograms (see above) by the mass of a hydrogen atom in kilograms (you will need to use a calculator for this!). If you do this calculation, you will see that the mass of a carbon atom is twelve times greater than the mass of a hydrogen atom. When we use atomic mass units instead of kilograms, it becomes easier to see this. Atomic mass units are therefore not giving us the *actual* mass of an atom, but rather its mass *relative* to the mass of one (carefully chosen) atom in the Periodic Table. Although carbon is the usual element to compare other elements to, oxygen and hydrogen have also been used. The important thing to remember here is that the atomic mass unit is relative to one (carefully chosen) element. The atomic masses of some elements are shown in the table (Table 4.2) below.

Element	**Atomic mass (u)**
Carbon (C)	12
Nitrogen (N)	14
Bromine (Br)	80
Magnesium (Mg)	24
Potassium (K)	39
Calcium (Ca)	40
Oxygen (O)	16

Table 4.2: The atomic mass number of some of the elements

The actual value of 1 atomic mass unit is $1,67 \times 10^{-24}\,\mathrm{g}$ or $1,67 \times 10^{-27}\,\mathrm{kg}$. This is a very tiny mass!

1.32.2 How big is an atom?

TIP: *pm* stands for *picometres*. $1\,\mathrm{pm} = 10^{-12}\,\mathrm{m}$

Atomic radius also varies depending on the element. On average, the radius of an atom ranges from $32\,\mathrm{pm}$ (Helium) to $225\,\mathrm{pm}$ (Caesium). Using different units, $100\,\mathrm{pm} = 1\,\mathrm{Angstrom}$, and $1\,\mathrm{Angstrom} = 10^{-10}\,\mathrm{m}$. That is the same as saying that $1\,\mathrm{Angstrom} = 0,0000000010\,\mathrm{m}$ or that $100\,\mathrm{pm} = 0,0000000010\,\mathrm{m}$! In other words, the diameter of an atom ranges from $0,0000000010\,m$ to $0,0000000067\,m$. This is very small indeed.

The atomic radii given above are for the whole atom (nucleus and electrons). The nucleus itself is even smaller than this by a factor of about 23 000 in uranium and 145 000 in hydrogen. If the nucleus were the size of a golf ball, then the nearest electrons would be about one kilometer away! This should give help you realise that the atom is mostly made up of empty space.

1.33 Atomic structure

(section shortcode: C10025)

As a result of the work done by previous scientists on atomic models (that we discussed in "Models of the Atom" (Section 3.3: Models of the Atom)), scientists now have a good idea of what an atom looks like. This knowledge is important because it helps us to understand why materials have different properties and why some materials bond with others. Let us now take a closer look at the microscopic structure of the atom.

So far, we have discussed that atoms are made up of a positively charged **nucleus** surrounded by one or more negatively charged **electrons**. These electrons orbit the nucleus.

1.33.1 The Electron

The electron is a very light particle. It has a mass of $9,11 \times 10^{-31}\ \mathrm{kg}$. Scientists believe that the electron can be treated as a **point particle** or **elementary particle** meaning that it can't be broken down into anything smaller. The electron also carries one unit of **negative** electric charge which is the same as $1,6 \times 10^{-19}\ \mathrm{C}$ (Coulombs).

The electrons determine the charge on an atom. If the number of electrons is the same as the number of protons then the atom will be neutral. If the number of electrons is greater than the number of protons then the atom will be negatively charged. If the number of electrons is less than the number of protons then the atom will be positively charged. Atoms that are not neutral are called ions. Ions will be covered in more detail in a later chapter. For now all you need to know is that for each electron you remove from an atom you loose -1 of charge and for each electron that you add to an atom you gain $+1$ of charge. For example, the charge on an atom of sodium after removing one electron is -1.

1.33.2 The Nucleus

Unlike the electron, the nucleus **can** be broken up into smaller building blocks called **protons** and **neutrons**. Together, the protons and neutrons are called **nucleons**.

The Proton

Each proton carries one unit of **positive** electric charge. Since we know that atoms are **electrically neutral**, i.e. do not carry any extra charge, then the number of protons in an atom has to be the same as the number of electrons to balance out the positive and negative charge to zero. The total positive charge of a nucleus is equal to the number of protons in the nucleus. The proton is much heavier than the electron (10 000 times heavier!) and has a mass of $1,6726 \times 10^{-27}\ \mathrm{kg}$. When we talk about the atomic mass of an atom, we are mostly referring to the combined mass of the protons and neutrons, i.e. the nucleons.

The Neutron

The neutron is electrically neutral i.e. it carries no charge at all. Like the proton, it is much heavier than the electron and its mass is $1,6749 \times 10^{-27}$ kg (slightly heavier than the proton).

NOTE: Rutherford predicted (in 1920) that another kind of particle must be present in the nucleus along with the proton. He predicted this because if there were only positively charged protons in the nucleus, then it should break into bits because of the repulsive forces between the like-charged protons! Also, if protons were the only particles in the nucleus, then a helium nucleus (atomic number 2) would have two protons and therefore only twice the mass of hydrogen. However, it is actually **four** times heavier than hydrogen. This suggested that there must be something else inside the nucleus as well as the protons. To make sure that the atom stays electrically neutral, this particle would have to be neutral itself. In 1932 James Chadwick discovered the neutron and measured its mass.

	proton	neutron	electron
Mass (kg)	$1,6726 \times 10^{-27}$	$1,6749 \times 10^{-27}$	$9,11 \times 10^{-31}$
Units of charge	$+1$	0	-1
Charge (C)	$1,6 \times 10^{-19}$	0	$-1,6 \times 10^{-19}$

Table 4.3: Summary of the particles inside the atom

NOTE: Unlike the electron which is thought to be a **point particle** and unable to be broken up into smaller pieces, the proton and neutron **can** be divided. Protons and neutrons are built up of smaller particles called **quarks**. The proton and neutron are made up of 3 quarks each.

1.34 Atomic number and atomic mass number

(section shortcode: C10026)

The chemical properties of an element are determined by the charge of its nucleus, i.e. by the **number of protons**. This number is called the **atomic number** and is denoted by the letter **Z**.

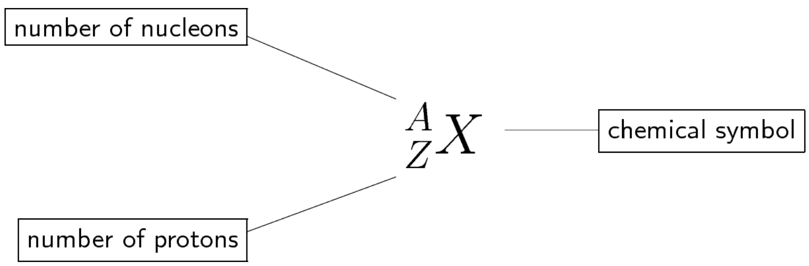
Definition: Atomic number (Z)
The number of protons in an atom

You can find the atomic number on the periodic table. The atomic number is an integer and ranges from 1 to about 118.

The mass of an atom depends on how many nucleons its nucleus contains. The number of nucleons, i.e. the total number of protons **plus** neutrons, is called the **atomic mass number** and is denoted by the letter **A**.

Definition: Atomic mass number (A)
The number of protons and neutrons in the nucleus of an atom

Standard notation shows the chemical symbol, the atomic mass number and the atomic number of an element as follows:



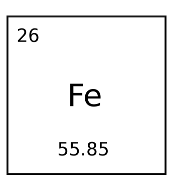
NOTE: A nuclide is a distinct kind of atom or nucleus characterized by the number of protons and neutrons in the atom. To be absolutely correct, when we represent atoms like we do here, then we should call them nuclides.

For example, the iron nucleus which has 26 protons and 30 neutrons, is denoted as:

$$^{56}_{26}\mathrm{Fe} \tag{4.1}$$

where the atomic number is $Z = 26$ and the mass number $A = 56$. The number of neutrons is simply the difference $N = A - Z$.

TIP: Don't confuse the notation we have used above with the way this information appears on the Periodic Table. On the Periodic Table, the atomic number usually appears in the top lefthand corner of the block or immediately above the element's symbol. The number below the element's symbol is its **relative atomic mass**. This is not exactly the same as the atomic mass number. This will be explained in "Isotopes" (Section 3.7: Isotopes). The example of iron is shown below.



You will notice in the example of iron that the atomic mass number is more or less the same as its atomic mass. Generally, an atom that contains *n* nucleons (protons and neutrons), will have a mass approximately equal to nu. For example the mass of a $\mathrm{C} - 12$ atom which has 6 protons, 6 neutrons and 6 electrons is 12u, where the protons and neutrons have about the same mass and the electron mass is negligible.

Exercise 4.1 Use standard notation to represent sodium and give the number of protons, neutrons and electrons in the element.

Solution to Exercise

Step 1. Sodium is given by Na
Step 2. Sodium has 11 protons, so we have: ${}_{11}\mathrm{Na}$
Step 3. Sodium has 12 neutrons.
Step 4. $A = N + Z = 12 + 11 = 23$
Step 5. In standard notation sodium is given by: ${}^{23}_{11}\mathrm{Na}$. The number of protons is 11, the number of neutrons is 12 and the number of electrons is 11.

1.34.1 The structure of the atom

1. Explain the meaning of each of the following terms:
   - **a.** nucleus
   - **b.** electron
   - **c.** atomic mass
2. Complete the following table: (Note: You will see that the atomic masses on the Periodic Table are not *whole numbers*. This will be explained later. For now, you can round off to the nearest whole number.)

Element	**Atomic mass**	**Atomic number**	**Number of protons**	**Number of electrons**	**Number of neutrons**
Mg	24	12			
O			8		
		17			
Ni				28	
	40				20
Zn					
					0

continued on next page

C	12			6	

Table 4.4

3. Use standard notation to represent the following elements:
   a. potassium
   b. copper
   c. chlorine
4. For the element $^{35}_{17}\mathrm{Cl}$, give the number of ...
   a. protons
   b. neutrons
   c. electrons

   ... in the atom.

5. Which of the following atoms has 7 electrons?
   a. $^{5}_{2}\mathrm{He}$
   b. $^{13}_{6}\mathrm{C}$
   c. $^{7}_{3}\mathrm{Li}$
   d. $^{15}_{7}\mathrm{N}$
6. In each of the following cases, give the number or the element symbol represented by 'X'.
   a. $^{40}_{18}\mathrm{X}$
   b. $^{x}_{20}\mathrm{Ca}$
   c. $^{31}_{x}\mathrm{P}$
7. Complete the following table:

	A	Z	N
$^{235}_{92}\mathrm{U}$			
$^{238}_{92}\mathrm{U}$			

Table 4.5

   In these two different forms of Uranium...
   a. What is the *same*?
   b. What is *different*?

   Uranium can occur in different forms, called *isotopes*. You will learn more about isotopes in "Isotopes" (Section 3.7: Isotopes).

Find the answers with the shortcodes:

(1.) lI0 (2.) lI8 (3.) lI9 (4.) lIX (5.) lIk (6.) lIK
(7.) lIB

1.35 Isotopes

(section shortcode: C10027)

1.35.1 What is an isotope?

The chemical properties of an element depend on the number of protons and electrons inside the atom. So if a neutron or two is added or removed from the nucleus, then the chemical properties will not change. This means that such an atom would remain in the same place in the Periodic Table. For example, no matter how many neutrons we add or subtract from a nucleus with 6 protons, that element will **always** be called carbon and have the element symbol C (see the Table of Elements). Atoms which have the same number of protons, but a different number of neutrons, are called **isotopes**.

Definition: Isotope
The **isotope** of a particular element is made up of atoms which have the same number of protons as the atoms in the original element, but a different number of neutrons.

The different isotopes of an element have the same atomic number Z but different mass numbers A because they have a different number of neutrons N. The chemical properties of the different isotopes of an element are the same, but they might vary in how stable their nucleus is. Note that we can also write elements as $\mathrm{X} - \mathrm{A}$ where the X is the element symbol and the A is the atomic mass of that element. For example, $\mathrm{C}{-}12$ has an atomic mass of 12 and $\mathrm{Cl}{-}35$ has an atomic mass of 35 u, while $\mathrm{Cl}{-}37$ has an atomic mass of 37 u.

NOTE: In Greek, "same place" reads as $\acute{\iota}\ \sigma o\varsigma\tau\ \acute{o}\ \pi o\varsigma$ (isos topos). This is why atoms which have the same number of protons, but different numbers of neutrons, are called *isotopes*. They are in the same place on the Periodic Table!

The following worked examples will help you to understand the concept of an isotope better.

Exercise 4.2: Isotopes For the element $^{234}_{92}\mathrm{U}$ (uranium), use standard notation to describe:

1. the isotope with 2 fewer neutrons
2. the isotope with 4 more neutrons

Solution to Exercise

Step 1. We know that isotopes of any element have the **same** number of protons (same atomic number) in each atom, which means that they have the same chemical symbol. However, they have a different number of neutrons, and therefore a different mass number.

Step 2. Therefore, any isotope of uranium will have the symbol:

$$\mathrm{U} \tag{4.2}$$

Also, since the number of protons in uranium isotopes is always the same, we can write down the atomic number:

$$_{92}\mathrm{U} \tag{4.3}$$

Now, if the isotope we want has 2 fewer neutrons than $^{234}_{92}\mathrm{U}$, then we take the original mass number and subtract 2, which gives:

$$^{232}_{92}\mathrm{U} \tag{4.4}$$

Following the steps above, we can write the isotope with 4 more neutrons as:

$$^{238}_{92}\mathrm{U} \tag{4.5}$$

Exercise 4.3: Isotopes Which of the following are isotopes of $^{40}_{20}\mathrm{Ca}$?

- $^{40}_{19}\mathrm{K}$
- $^{42}_{20}\mathrm{Ca}$
- $^{40}_{18}\mathrm{Ar}$

Solution to Exercise

Step 1. We know that isotopes have the same atomic number but different mass numbers.

Step 2. You need to look for the element that has the same atomic number but a different atomic mass number. The only element is $^{42}_{20}Ca$. What is different is that there are 2 more neutrons than in the original element.

Exercise 4.4: Isotopes For the sulphur isotope $^{33}_{16}\mathrm{S}$, give the number of...

a. protons
b. nucleons
c. electrons
d. neutrons

Solution to Exercise

Step 1. $Z = 16$, therefore the number of protons is 16 (answer to (a)).
Step 2. $A = 33$, therefore the number of nucleons is 33 (answer to (b)).
Step 3. The atom is neutral, and therefore the number of electrons is the same as the number of protons. The number of electrons is 16 (answer to (c)).
Step 4.

$$N = A - Z = 33 - 16 = 17 \tag{4.6}$$

The number of neutrons is 17 (answer to (d)).

Isotopes

1. Atom A has 5 protons and 5 neutrons, and atom B has 6 protons and 5 neutrons. These atoms are...
   - **a.** allotropes
   - **b.** isotopes
   - **c.** isomers
   - **d.** atoms of different elements
2. For the sulphur isotopes, ${}^{32}_{16}\mathrm{S}$ and ${}^{34}_{16}\mathrm{S}$, give the number of...
   - **a.** protons
   - **b.** nucleons
   - **c.** electrons
   - **d.** neutrons
3. Which of the following are isotopes of ${}^{35}_{17}\mathrm{Cl}$?
   - **a.** ${}^{17}_{35}\mathrm{Cl}$
   - **b.** ${}^{35}_{17}\mathrm{Cl}$
   - **c.** ${}^{37}_{17}\mathrm{Cl}$
4. Which of the following are isotopes of $\mathrm{U}-235$? (X represents an element symbol)
   - **a.** ${}^{238}_{92}\mathrm{X}$
   - **b.** ${}^{238}_{90}\mathrm{X}$
   - **c.** ${}^{235}_{92}\mathrm{X}$

www Find the answers with the shortcodes:

(1.) ll4 (2.) llZ (3.) llW (4.) llD

1.35.2 Relative atomic mass

It is important to realise that the atomic mass of isotopes of the same element will be different because they have a different number of nucleons. Chlorine, for example, has two common isotopes which are chlorine-35 and chlorine-37. Chlorine-35 has an atomic mass of 35 u, while chlorine-37 has an atomic mass of 37 u. In the world around us, both of these isotopes occur naturally. It doesn't make sense to say that the element chlorine has an atomic mass of 35 u, or that it has an atomic mass of 37 u. Neither of these are absolutely true since the mass varies depending on the form in which the element occurs. We need to look at how much more common one is than the other in order to calculate the **relative atomic mass** for the element chlorine. This is the number that you find on the Periodic Table.

Definition: Relative atomic mass
Relative atomic mass is the average mass of one atom of all the naturally occurring isotopes of a particular chemical element, expressed in atomic mass units.

NOTE: The relative atomic mass of some elements depends on where on Earth the element is found. This is because the isotopes can be found in varying ratios depending on certain factors such as geological composition, etc. The International Union of Pure and Applied Chemistry (IUPAC) has decided to give the relative atomic mass of some elements as a range to better represent the varying isotope ratios on the Earth. For the calculations that you will do at high school, it is enough to simply use one number without worrying about these ranges.

Exercise 4.5: The relative atomic mass of an isotopic element The element chlorine has two isotopes, chlorine-35 and chlorine-37. The abundance of these isotopes when they occur naturally is 75% chlorine-35 and 25% chlorine-37. Calculate the *average* relative atomic mass for chlorine.

Solution to Exercise

Step 1. Contribution of $\mathrm{Cl}-35 = \left(\frac{75}{100} \times 35\right) = 26,25\,\mathrm{u}$

Step 2. Contribution of $\mathrm{Cl}-37 = \left(\frac{25}{100} \times 37\right) = 9,25\,\mathrm{u}$

Step 3. $\text{Relative atomic mass of chlorine} = 26,25\,\mathrm{u} + 9,25\,\mathrm{u} = 35,5\,\mathrm{u}$
If you look on the periodic table, the average relative atomic mass for chlorine is $35,5\,\mathrm{u}$. You will notice that for many elements, the relative atomic mass that is shown is not a whole number. You should now understand that this number is the *average* relative atomic mass for those elements that have naturally occurring isotopes.

This simulation allows you to see how isotopes and relative atomic mass are inter related.
(Simulation: IbT)

Isotopes

1. Complete the table below:

Isotope	Z	A	Protons	Neutrons	Electrons
Carbon-12					
Carbon-14					
Chlorine-35					
Chlorine-37					

Table 4.6

2. If a sample contains 90% carbon-12 and 10% carbon-14, calculate the relative atomic mass of an atom in that sample.
3. If a sample contains 22,5% Cl-37 and 77,5% Cl-35, calculate the relative atomic mass of an atom in that sample.

www Find the answers with the shortcodes:

(1.) llj (2.) llb (3.) llT

Group Discussion : The changing nature of scientific knowledge

Scientific knowledge is not static: it changes and evolves over time as scientists build on the ideas of others to come up with revised (and often improved) theories and ideas. In this chapter for example, we saw how peoples' understanding of atomic structure changed as more information was gathered about the atom. There are many more examples like this one in the field of science. For example, think about our knowledge of the solar system and the origin of the universe, or about the particle and wave nature of light.

Often, these changes in scientific thinking can be very controversial because they disturb what people have come to know and accept. It is important that we realise that what we know *now* about science may also change. An important part of being a scientist is to be a *critical thinker*. This means that you need to question information that you are given and decide whether it is accurate and whether it can be accepted as true. At the same time, you need to learn to be open to new ideas and not to become stuck in what you believe is right... there might just be something new waiting around the corner that you have not thought about!

In groups of 4-5, discuss the following questions:

- Think about some other examples where scientific knowledge has changed because of new ideas and discoveries:
  - What were these new ideas?
  - Were they controversial? If so, why?
  - What role (if any) did *technology* play in developing these new ideas?
  - How have these ideas affected the way we understand the world?

- Many people come up with their own ideas about how the world works. The same is true in science. So how do we, and other scientists, know what to believe and what not to? How do we know when new ideas are 'good' science or 'bad' science? In your groups, discuss some of the things that would need to be done to check whether a new idea or theory was worth listening to, or whether it was not.
- Present your ideas to the rest of the class.

1.36 Electron configuration

(section shortcode: C10028)

1.36.1 The energy of electrons

You will remember from our earlier discussions that an atom is made up of a central nucleus, which contains protons and neutrons and that this nucleus is surrounded by electrons. Although these electrons all have the same charge and the same mass, each electron in an atom has a different amount of *energy*. Electrons that have the *lowest* energy are found closest to the nucleus where the attractive force of the positively charged nucleus is the greatest. Those electrons that have *higher* energy, and which are able to overcome the attractive force of the nucleus, are found further away.

1.36.2 Energy quantisation and line emission spectra (Not in CAPS, included for completeness)

If the energy of an atom is increased (for example when a substance is heated), the energy of the electrons inside the atom can be increased (when an electron has a higher energy than normal it is said to be "excited"). For the excited electron to go back to its original energy (called the ground state), it needs to release energy. It releases energy by emitting light. If one heats up different elements, one will see that for each element, light is emitted only at certain frequencies (or wavelengths). Instead of a smooth continuum of frequencies, we see lines (called emission lines) at particular frequencies. These frequencies correspond to the energy of the emitted light. If electrons could be excited to any energy and lose any amount of energy, there would be a continuous spread of light frequencies emitted. However, the sharp lines we see mean that there are only certain particular energies that an electron can be excited to, or can lose, for each element.

You can think of this like going up a flight of steps: you can't lift your foot by *any* amount to go from the ground to the first step. If you lift your foot too low you'll bump into the step and be stuck on the ground level. You have to lift your foot just the right amount (the height of the step) to go to the next step, and so on. The same goes for electrons and the amount of energy they can have. This is called **quantisation of energy** because there are only certain quantities of energy that an electron can have in an atom. Like steps, we can think of these quantities as **energy levels** in the atom. The energy of the light released when an electron drops down from a higher energy level to a lower energy level is the same as the difference in energy between the two levels.

1.36.3 Electron configuration

We will start with a very simple view of the arrangement or configuration of electrons around an atom. This view simply states that electrons are arranged in energy levels (or shells) around the nucleus of an atom. These energy levels are numbered 1, 2, 3, etc. Electrons that are in the first energy level (energy level 1) are closest to the nucleus and will have the lowest energy. Electrons further away from the nucleus will have a higher energy.

In the following examples, the energy levels are shown as concentric circles around the central nucleus. The important thing to know for these diagrams is that the first energy level can hold 2 electrons, the second energy level can hold 8 electrons and the third energy level can hold 8 electrons.

1. **Lithium** Lithium (Li) has an atomic number of 3, meaning that in a neutral atom, the number of electrons will also be 3. The first two electrons are found in the first energy level, while the third electron is found in the second energy level (Figure 4.8).

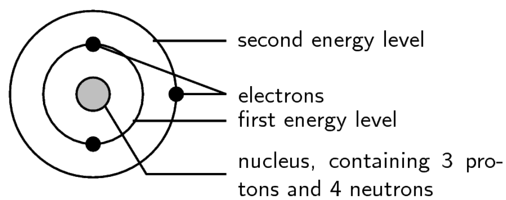


Figure 4.8: The arrangement of electrons in a lithium atom.

2. **Fluorine** Fluorine (F) has an atomic number of 9, meaning that a neutral atom also has 9 electrons. The first 2 electrons are found in the first energy level, while the other 7 are found in the second energy level (Figure 4.9).

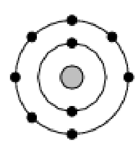


Figure 4.9: The arrangement of electrons in a fluorine atom.

3. **Argon** Argon has an atomic number of 18, meaning that a neutral atom also has 18 electrons. The first 2 electrons are found in the first energy level, the next 8 are found in the second energy level, and the last 8 are found in the third energy level (Figure 4.10).

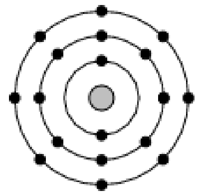
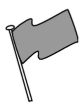


Figure 4.10: The arrangement of electrons in an argon atom.

But the situation is slightly more complicated than this. Within each energy level, the electrons move in **orbitals**. An orbital defines the spaces or regions where electrons move.



Definition: Atomic orbital
An atomic orbital is the region in which an electron may be found around a single atom.

There are different orbital shapes, but we will be mainly dealing with only two. These are the 's' and 'p' orbitals (there are also 'd' and 'f' orbitals). The 's' orbitals are spherical and the 'p' orbitals are dumbbell shaped.

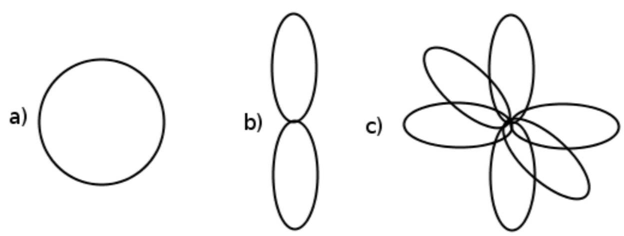


Figure 4.11: The shapes of orbitals. a) shows an 's' orbital, b) shows a single 'p' orbital and c) shows the three 'p' orbitals.

The first energy level contains only one 's' orbital, the second energy level contains one 's' orbital and three 'p' orbitals and the third energy level contains one 's' orbital and three 'p' orbitals (as well as 5 'd' orbitals). Within each energy level, the 's' orbital is at a lower energy than the 'p' orbitals. This arrangement is shown in Figure 4.12.

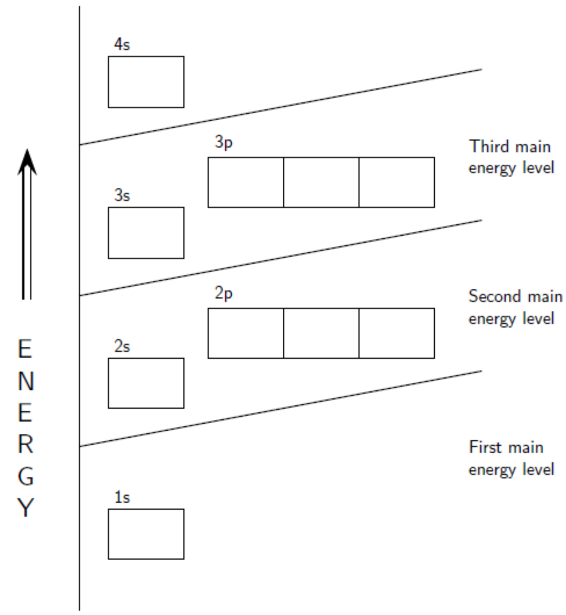


Figure 4.12: The positions of the first ten orbitals of an atom on an energy diagram. Note that each block is able to hold two electrons.

This diagram also helps us when we are working out the electron configuration of an element. The electron configuration of an element is the arrangement of the electrons in the shells and subshells. There are a few guidelines for working out the electron configuration. These are:

- Each orbital can only hold **two electrons**. Electrons that occur together in an orbital are called an **electron pair**.
- An electron will always try to enter an orbital with the lowest possible energy.
- An electron will occupy an orbital on its own, rather than share an orbital with another electron. An electron would also rather occupy a lower energy orbital *with* another electron, before occupying a higher energy orbital. In other words, within one energy level, electrons will fill an 's' orbital before starting to fill 'p' orbitals.

- The s subshell can hold 2 electrons
- The p subshell can hold 6 electrons

In the examples you will cover, you will mainly be filling the s and p subshells. Occasionally you may get an example that has the d subshell. The f subshell is more complex and is not covered at this level.

The way that electrons are arranged in an atom is called its **electron configuration**.

Definition: Electron configuration
Electron configuration is the arrangement of electrons in an atom, molecule or other physical structure.

An element's electron configuration can be represented using **Aufbau diagrams** or energy level diagrams. An Aufbau diagram uses arrows to represent electrons. You can use the following steps to help you to draw an Aufbau diagram:

1. Determine the number of electrons that the atom has.
2. Fill the 's' orbital in the first energy level (the $1\mathrm{s}$ orbital) with the first two electrons.
3. Fill the 's' orbital in the second energy level (the $2\mathrm{s}$ orbital) with the second two electrons.
4. Put one electron in each of the three 'p' orbitals in the second energy level (the $2\mathrm{p}$ orbitals) and then if there are still electrons remaining, go back and place a second electron in each of the $2\mathrm{p}$ orbitals to complete the electron pairs.
5. Carry on in this way through each of the successive energy levels until all the electrons have been drawn.

TIP: When there are two electrons in an orbital, the electrons are called an **electron pair**. If the orbital only has one electron, this electron is said to be an **unpaired electron**. Electron pairs are shown with arrows pointing in opposite directions. You may hear people talking of the Pauli exclusion principle. This principle says that electrons have a property known as spin and two electrons in an orbital will not spin the same way. This is why we use arrows pointing in opposite directions. An arrow pointing up denotes an electron spinning one way and an arrow pointing downwards denotes an electron spinning the other way.

Interesting Fact

NOTE: Aufbau is the German word for 'building up'. Scientists used this term since this is exactly what we are doing when we work out electron configuration, we are building up the atoms structure.

Sometimes people refer to Hund's rule for electron configuration. This rule simply says that electrons would rather be in a subshell on it's own then share a subshell. This is why, when you are filling the subshells you put one electron in each subshell and only if there are extra electrons do you go back and fill the subshell, before moving onto the next energy level.

An Aufbau diagram for the element Lithium is shown in Figure 4.13.

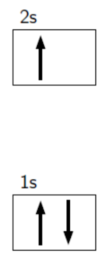


Figure 4.13: The electron configuration of Lithium, shown on an Aufbau diagram

A special type of notation is used to show an atom's electron configuration. The notation describes the energy levels, orbitals and the number of electrons in each. For example, the electron configuration of lithium is $1s^2 2s^1$. The number and letter describe the energy level and orbital and the number above the orbital shows how many electrons are in that orbital.

Aufbau diagrams for the elements fluorine and argon are shown in Figure 4.14 and Figure 4.15 respectively. Using standard notation, the electron configuration of fluorine is $1s^2 2s^2 2p^5$ and the electron configuration of argon is $1s^2 2s^2 2p^6$.

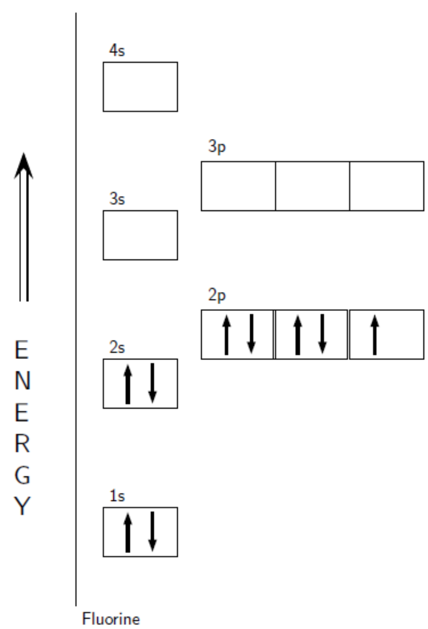


Figure 4.14: An Aufbau diagram showing the electron configuration of fluorine

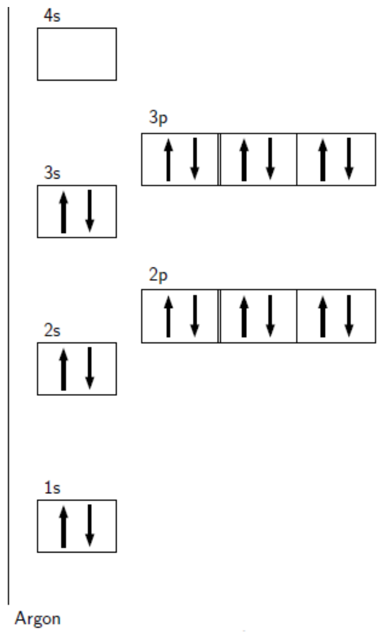
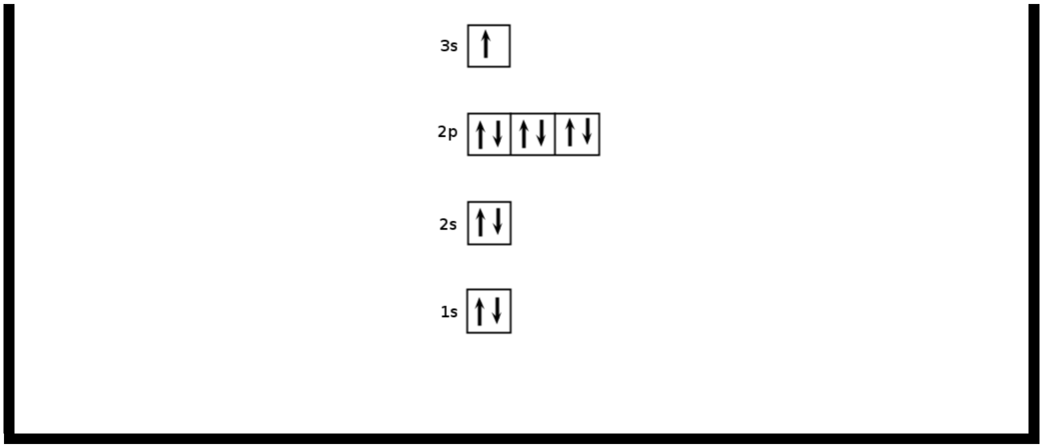


Figure 4.15: An Aufbau diagram showing the electron configuration of argon

Exercise 4.6: Aufbau diagrams Give the electron configuration for sodium (Na) and draw an aufbau diagram.

Solution to Exercise

Step 1. Sodium has 11 electrons.
Step 2. We start by placing two electrons in the $1s$ orbital: $1s^2$. Now we have 9 electrons left to place in orbitals, so we put two in the $2s$ orbital: $2s^2$. There are now 7 electrons to place in orbitals so we place 6 of them in the $2p$ orbital: $2p^6$. The last electron goes into the $3s$ orbital: $3s^1$.
Step 3. The electron configuration is: $1s^2 2s^2 2p^6 3s^1$
Step 4. Using the electron configuration we get the following diagram:



1.36.4 Core and valence electrons

Electrons in the outermost energy level of an atom are called **valence electrons**. The electrons that are in the energy shells closer to the nucleus are called **core electrons**. Core electrons are all the electrons in an atom, excluding the valence electrons. An element that has its valence energy level full is *more stable* and *less likely to react* than other elements with a valence energy level that is not full.

Definition: Valence electrons
The electrons in the outer energy level of an atom

Definition: Core electrons
All the electrons in an atom, excluding the valence electrons

1.36.5 The importance of understanding electron configuration

By this stage, you may well be wondering why it is important for you to understand how electrons are arranged around the nucleus of an atom. Remember that during chemical reactions, when atoms come into contact with one another, it is the *electrons* of these atoms that will interact first. More specifically, it is the **valence electrons** of the atoms that will determine how they react with one another.

To take this a step further, an atom is at its most stable (and therefore *unreactive*) when all its orbitals are full. On the other hand, an atom is least stable (and therefore most *reactive*) when its valence electron orbitals are not full. This will make more sense when we go on to look at chemical bonding in a later chapter. To put it simply, the valence electrons are largely responsible for an element's chemical behaviour and elements that have the same number of valence electrons often have similar chemical properties.

One final point to note about electron configurations is stability. Which configurations are stable and which are not? Very simply, the most stable configurations are the ones that have full energy levels. These configurations occur in the noble gases. The noble gases are very stable elements that do not react easily (if at all) with any other elements. This is due to the full energy levels. All elements would like to reach the most stable electron configurations, i.e. all elements want to be noble gases. This principle of stability is sometimes referred to as the octet rule. An octet is a set of 8, and the number of electrons in a full energy level is 8.

Experiment: Flame tests

Aim:
To determine what colour a metal cation will cause a flame to be.

Apparatus:
Watch glass, bunsen burner, methanol, bamboo sticks, metal salts (e.g. $NaCl$, $CuCl_2$, $CaCl_2$, KCl, etc.) and metal powders (e.g. copper, magnesium, zinc, iron, etc.)

Method:
For each salt or powder do the following:

1. Dip a clean bamboo stick into the methanol
2. Dip the stick into the salt or powder
3. Wave the stick through the flame from the bunsen burner. DO NOT hold the stick in the flame, but rather wave it back and forth through the flame.
4. Observe what happens

Results:
Record your results in a table, listing the metal salt and the colour of the flame.

Conclusion:
You should have observed different colours for each of the metal salts and powders that you tested.

The above experiment on flame tests relates to the line emission spectra of the metals. These line emission spectra are a direct result of the arrangement of the electrons in metals.

Energy diagrams and electrons

1. Draw Aufbau diagrams to show the electron configuration of each of the following elements:
   a. magnesium
   b. potassium
   c. sulphur
   d. neon
   e. nitrogen

2. Use the Aufbau diagrams you drew to help you complete the following table:

Element	**No. of energy levels**	**No. of core electrons**	**No. of valence electrons**	**Electron configuration (standard notation)**
Mg				
K				
S				
Ne				
N				

Table 4.7

3. Rank the elements used above in order of *increasing reactivity*. Give reasons for the order you give.
www Find the answers with the shortcodes: ll2

Group work : Building a model of an atom

Earlier in this chapter, we talked about different 'models' of the atom. In science, one of the uses of models is that they can help us to understand the structure of something that we can't see. In the case of the atom, models help us to build a picture in our heads of what the atom looks like.
Models are often simplified. The small toy cars that you may have played with as a child are models. They give you a good idea of what a real car looks like, but they are much smaller and much simpler. A model cannot always be absolutely accurate and it is important that we realise this so that we don't build up a false idea about something.
In groups of 4-5, you are going to build a model of an atom. Before you start, think about these questions:

- What information do I know about the structure of the atom? (e.g. what parts make it up? how big is it?)
- What materials can I use to represent these parts of the atom as accurately as I can?
- How will I put all these different parts together in my model?

As a group, share your ideas and then plan how you will build your model. Once you have built your model, discuss the following questions:

- Does our model give a good idea of what the atom actually looks like?
- In what ways is our model *inaccurate*? For example, we know that electrons *move* around the atom's nucleus, but in your model, it might not have been possible for you to show this.
- Are there any ways in which our model could be improved?

Now look at what other groups have done. Discuss the same questions for each of the models you see and record your answers.
The following simulation allows you to build an atom
www (Simulation: lbb)
This is another simulation that allows you to build an atom. This simulation also provides a summary of what you have learnt so far.
www (Simulation: lbj)

1.37 Summary

(section shortcode: C10029)

- Much of what we know today about the atom, has been the result of the work of a number of scientists who have added to each other's work to give us a good understanding of atomic structure.
- Some of the important scientific contributors include **J.J.Thomson** (discovery of the electron, which led to the Plum Pudding Model of the atom), **Ernest Rutherford** (discovery that positive charge is concentrated in the centre of the atom) and **Niels Bohr** (the arrangement of electrons around the nucleus in energy levels).
- Because of the very small mass of atoms, their mass is measured in **atomic mass units** (u). $1\,u = 1,67 \times 10^{-24}\,g$.
- An atom is made up of a central **nucleus** (containing **protons** and **neutrons**), surrounded by **electrons**.
- The **atomic number** (Z) is the number of protons in an atom.
- The **atomic mass number** (A) is the number of protons and neutrons in the nucleus of an atom.
- The **standard notation** that is used to write an element, is ${}^{A}_{Z}\mathrm{X}$, where X is the element symbol, A is the atomic mass number and Z is the atomic number.
- The **isotope** of a particular element is made up of atoms which have the same number of protons as the atoms in the original element, but a different number of neutrons. This means that not all atoms of an element will have the same atomic mass.
- The **relative atomic mass** of an element is the average mass of one atom of all the naturally occurring isotopes of a particular chemical element, expressed in atomic mass units. The relative atomic mass is written under the elements' symbol on the Periodic Table.
- The energy of electrons in an atom is **quantised**. Electrons occur in specific energy levels around an atom's nucleus.
- Within each energy level, an electron may move within a particular shape of **orbital**. An orbital defines the space in which an electron is most likely to be found. There are different orbital shapes, including s, p, d and f orbitals.
- Energy diagrams such as **Aufbau diagrams** are used to show the electron configuration of atoms.
- The electrons in the outermost energy level are called **valence electrons**.
- The electrons that are not valence electrons are called **core electrons**.
- Atoms whose outermost energy level is full, are less chemically reactive and therefore more stable, than those atoms whose outer energy level is not full.

(Presentation: P10030)

1.37.1 End of chapter exercises

1. Write down only the word/term for each of the following descriptions.
   a. The sum of the number of protons and neutrons in an atom
   b. The defined space around an atom's nucleus, where an electron is most likely to be found
2. For each of the following, say whether the statement is True or False. If it is False, re-write the statement correctly.
   a. ${}^{20}_{10}\mathrm{Ne}$ and ${}^{22}_{10}\mathrm{Ne}$ each have 10 protons, 12 electrons and 12 neutrons.
   b. The atomic mass of any atom of a particular element is always the same.
   c. It is safer to use helium gas rather than hydrogen gas in balloons.
   d. Group 1 elements readily form negative ions.

3. Multiple choice questions: In each of the following, choose the **one** correct answer.
   a. The three basic components of an atom are:
      a. protons, neutrons, and ions
      b. protons, neutrons, and electrons
      c. protons, neutrinos, and ions
      d. protium, deuterium, and tritium
   b. The charge of an atom is...
      a. positive
      b. neutral
      c. negative
   c. If Rutherford had used neutrons instead of alpha particles in his scattering experiment, the neutrons would...
      a. not deflect because they have no charge
      b. have deflected more often
      c. have been attracted to the nucleus easily
      d. have given the same results
   d. Consider the isotope $^{234}_{92}\mathrm{U}$. Which of the following statements is *true*?
      a. The element is an isotope of $^{234}_{94}\mathrm{Pu}$
      b. The element contains 234 neutrons
      c. The element has the same electron configuration as $^{238}_{92}\mathrm{U}$
      d. The element has an atomic mass number of 92
   e. The electron configuration of an atom of chlorine can be represented using the following notation:
      a. $1s^2 2s^8 3s^7$
      b. $1s^2 2s^2 2p^6 3s^2 3p^5$
      c. $1s^2 2s^2 2p^6 3s^2 3p^6$
      d. $1s^2 2s^2 2p^5$
4. Give the standard notation for the following elements:
   a. beryllium
   b. carbon-12
   c. titanium-48
   d. fluorine
5. Give the electron configurations and aufbau diagrams for the following elements:
   a. aluminium
   b. phosphorus
   c. carbon
6. Use standard notation to represent the following elements:
   a. argon
   b. calcium
   c. silver-107
   d. bromine-79
7. For each of the following elements give the number of protons, neutrons and electrons in the element:
   a. $^{195}_{78}\mathrm{Pt}$
   b. $^{40}_{18}\mathrm{Ar}$
   c. $^{59}_{27}\mathrm{Co}$
   d. $^{7}_{3}\mathrm{Li}$
   e. $^{11}_{5}\mathrm{B}$

8. For each of the following elements give the element or number represented by 'x':

a. $^{103}_{45}\mathrm{X}$
b. $^{35}_{x}\mathrm{Cl}$
c. $^{x}_{4}\mathrm{Be}$

9. Which of the following are isotopes of $^{24}_{12}\mathrm{Mg}$:

a. $^{12}_{25}\mathrm{Mg}$
b. $^{26}_{12}\mathrm{Mg}$
c. $^{24}_{13}\mathrm{Al}$

10. If a sample contains 69% of copper-63 and 31% of copper-65, calculate the relative atomic mass of an atom in that sample.

11. Complete the following table:

Element	Electron configuration	Core electrons	Valence electrons
Boron (B)			
Calcium (Ca)			
Silicon (Si)			
Lithium (Li)			
Neon (Ne)			

Table 4.8

12. Draw aufbau diagrams for the following elements:

a. beryllium
b. sulphur
c. argon

Find the answers with the shortcodes:

(1.) lif (2.) liG (3.a) li7 (3.b) liA (3.c) lio (3.d) lis
(3.e) liH (4.) lg7 (5.) lgG (6.) l44 (7.) l42 (8.) l4T
(9.) l4b (10.) l4j (11.) l4D (12.) lgW

The Periodic Table

1.38 The arrangement of atoms in the periodic table

(section shortcode: C10031)

The **periodic table of the elements** is a method of showing the chemical elements in a table. The elements are arranged in order of increasing atomic number. Most of the work that was done to arrive at the periodic table that we know, can be attributed to a man called **Dmitri Mendeleev** in 1869. Mendeleev was a Russian chemist who designed the table in such a way that recurring ("periodic") trends in the properties of the elements could be shown. Using the trends he observed, he even left gaps for those elements that he thought were 'missing'. He even predicted the properties that he thought the missing elements would have when they were discovered. Many of these elements were indeed discovered and Mendeleev's predictions were proved to be correct.

To show the recurring properties that he had observed, Mendeleev began new rows in his table so that elements with similar properties were in the same vertical columns, called **groups**. Each row was referred to as a **period**. One important feature to note in the periodic table is that all the non-metals are to the right of the zig-zag line drawn under the element boron. The rest of the elements are metals, with the exception of hydrogen which occurs in the first block of the table despite being a non-metal.

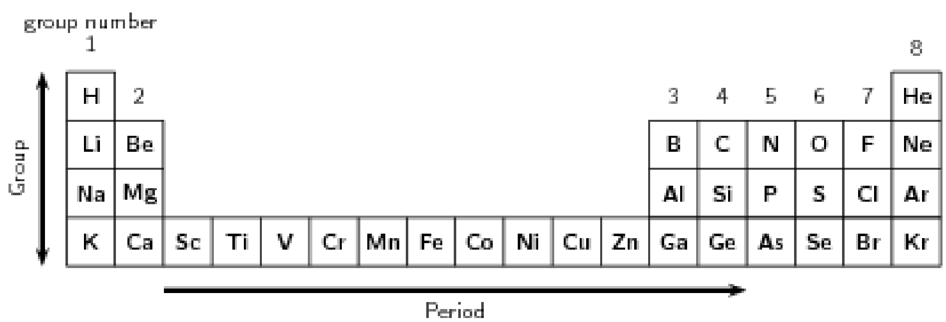


Figure 5.1: A simplified diagram showing part of the Periodic Table

You can view the periodic table online[2] . The full periodic table is also reproduced at the front of this book.

[2] http://periodictable.com/

1.38.1 Activity: Inventing the periodic table

You are the official chemist for the planet Zog. You have discovered all the same elements that we have here on Earth, but you don't have a periodic table. The citizens of Zog want to know how all these elements relate to each other. How would you invent the periodic table? Think about how you would organize the data that you have and what properties you would include. Do not simply copy Mendeleev's ideas, be creative and come up with some of your own. Research other forms of the periodic table and make one that makes sense to you. Present your ideas to your class.

1.38.2 Groups in the periodic table

A *group* is a vertical column in the periodic table and is considered to be the most important way of classifying the elements. If you look at a periodic table, you will see the groups numbered at the top of each column. The groups are numbered from left to right starting with 1 and ending with 18. This is the convention that we will use in this book. On some periodic tables you may see that the groups are numbered from left to right as follows: 1, 2, then an open space which contains the **transition elements**, followed by groups 3 to 8. Another way to label the groups is using Roman numerals. In some groups, the elements display very similar chemical properties and the groups are even given separate names to identify them.
The characteristics of each group are mostly determined by the electron configuration of the atoms of the element.

- *Group 1:* These elements are known as the **alkali metals** and they are very reactive.

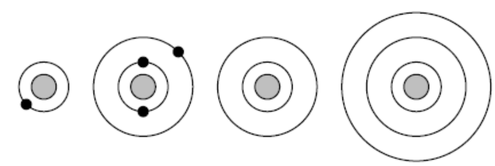


Figure 5.2: Electron diagrams for some of the Group 1 elements, with sodium and potasium incomplete; to be completed as an excersise.

- *Group 2:* These elements are known as the **alkali earth metals**. Each element only has two valence electrons and so in chemical reactions, the group 2 elements tend to *lose* these electrons so that the energy shells are complete. These elements are less reactive than those in group 1 because it is more difficult to lose two electrons than it is to lose one.
- *Group 13* elements have three valence electrons.
- *Group 16:* These elements are sometimes known as the chalcogens. These elements are fairly reactive and tend to gain electrons to fill their outer shell.
- *Group 17:* These elements are known as the **halogens**. Each element is missing just one electron from its outer energy shell. These elements tend to *gain* electrons to fill this shell, rather than losing them. These elements are also very reactive.
- *Group 18:* These elements are the **noble gases**. All of the energy shells of the halogens are full and so these elements are very unreactive.

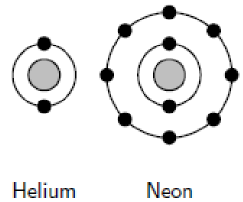


Figure 5.3: Electron diagrams for two of the noble gases, helium (He) and neon (Ne).

- *Transition metals:* The differences between groups in the transition metals are not usually dramatic.

TIP: The number of valence electrons of an element corresponds to its group number on the periodic table.

NOTE: Group 15 on the periodic table is sometimes called the pnictogens.

Investigation : The properties of elements

Refer to Figure 5.2.

1. Use a periodic table to help you to complete the last two diagrams for sodium (Na) and potassium (K).
2. What do you notice about the number of electrons in the valence energy level in each case?
3. Explain why elements from group 1 are more reactive than elements from group 2 on the periodic table (Hint: Think about the 'ionisation energy').

It is worth noting that in each of the groups described above, the **atomic diameter** of the elements increases as you move down the group. This is because, while the number of valence electrons is the same in each element, the number of core electrons increases as one moves down the group.

Khan academy video on the periodic table - 1 (Video: P10032)

1.38.3 Periods in the periodic table

A **period** is a horizontal row in the periodic table of the elements. Some of the trends that can be observed within a period are highlighted below:

- As you move from one group to the next within a period, the number of valence electrons increases by one each time.
- Within a single period, all the valence electrons occur in the same energy shell. If the period increases, so does the energy shell in which the valence electrons occur.

- In general, the diameter of atoms decreases as one moves from left to right across a period. Consider the attractive force between the positively charged nucleus and the negatively charged electrons in an atom. As you move across a period, the number of protons in each atom increases. The number of electrons also increases, but these electrons will still be in the same energy shell. As the number of protons increases, the force of attraction between the nucleus and the electrons will increase and the atomic diameter will decrease.
- Ionisation energy increases as one moves from left to right across a period. As the valence electron shell moves closer to being full, it becomes more difficult to remove electrons. The opposite is true when you move down a *group* in the table because more energy shells are being added. The electrons that are closer to the nucleus 'shield' the outer electrons from the attractive force of the positive nucleus. Because these electrons are not being held to the nucleus as strongly, it is easier for them to be removed and the ionisation energy decreases.
- In general, the reactivity of the elements decreases from left to right across a period.
- The formation of halides follows the general pattern: XCl_n (where X is any element in a specific group and n is the number of that specific group.). For example, the formula for the halides of group 1 will be $\mathrm{X}Cl$, for the second group the halides have the formula XCl_2 and in the third group the halides have the formula XCl_3. This should be easy to see if you remember the valency of the group and of the halides.

The formation of oxides show a trend as you move across a period. This should be easy to see if you think about valency. In the first group all the elements lose an electron to form a cation. So the formula for an oxide will be $\mathrm{X_2O}$. In the second group (moving from left to right across a period) the oxides have the formula XO. In the third group the oxides have the formula $\mathrm{X_2O_3}$.
Several other trends may be observed across a period such as density, melting points and boiling points. These trends are not as obvious to see as the above trends and often show variations to the general trend. Electron affinity and electronegativity also show some general trends across periods. Electron affinity can be thought of as how much an element wants electrons. Electron affinity generally increases from left to right across a period. Electronegativity is the tendency of atoms to attract electrons. The higher the electronegativity, the greater the atom attracts electrons. Electronegativity generally increases across a period (from left to right). Electronegativity and electron affinity will be covered in more detail in a later grade.
You may see periodic tables labeled with s-block, p-block, d-block and f-block. This is simply another way to group the elements. When we group elements like this we are simply noting which orbitals are being filled in each block. This method of grouping is not very useful to the work covered at this level.
Using the properties of the groups and the trends that we observe in certain properties (ionization energy, formation of halides and oxides, melting and boiling points, atomic diameter) we can predict the the properties of unknown elements. For example, the properties of the unfamiliar elements Francium (Fr), Barium (Ba), Astatine (At), and Xenon (Xe) can be predicted by knowing their position on the periodic table. Using the periodic table we can say: Francium (Group 1) is an alkali metal, very reactive and needs to lose 1 electron to obtain a full outer energy shell; Barium (Group 2) is an alkali earth metal and needs to lose 2 electrons to achieve stability; Astatine (Group 7) is a halogen, very reactive and needs to gain 1 electron to obtain a full outer energy shell; and Xenon (Group 8) is a noble gas and thus stable due to its full outer energy shell. This is how scientists are able to say what sort of properties the atoms in the last period have. Almost all of the elements in this period do not occur naturally on earth and are made in laboratories. These atoms do not exist for very long (they are very unstable and break apart easily) and so measuring their properties is difficult.

Exercise: Elements in the periodic table

Refer to the elements listed below:

- Lithium (Li)
- Chlorine (Cl)
- Magnesium (Mg)
- Neon (Ne)
- Oxygen (O)
- Calcium (Ca)
- Carbon (C)

Which of the elements listed above:

1. belongs to Group 1
2. is a halogen
3. is a noble gas
4. is an alkali metal
5. has an atomic number of 12
6. has 4 neutrons in the nucleus of its atoms
7. contains electrons in the 4th energy level
8. has only one valence electron
9. has all its energy orbitals full
10. will have chemical properties that are most similar
11. will form positive ions

Find the answers with the shortcodes:

(1.) liw

1.39 Ionisation Energy and the Periodic Table

(section shortcode: C10033)

1.39.1 Ions

In the previous section, we focused our attention on the electron configuration of *neutral* atoms. In a neutral atom, the number of protons is the same as the number of electrons. But what happens if an atom *gains* or *loses* electrons? Does it mean that the atom will still be part of the same element?

A change in the number of electrons of an atom does not change the type of atom that it is. However, the *charge* of the atom will change. If electrons are added, then the atom will become *more negative*. If electrons are taken away, then the atom will become *more positive*. The atom that is formed in either of these cases is called an **ion**. Put simply, an ion is a charged atom.

Definition: Ion

An ion is a charged atom. A positively charged ion is called a **cation** e.g. Na^{+}, and a negatively charged ion is called an **anion** e.g. F^{-}. The charge on an ion depends on the number of electrons that have been lost or gained.

But how do we know how many electrons an atom will gain or lose? Remember what we said about stability? We said that all atoms are trying to get a full outer shell. For the elements on the left hand side of the periodic table the easiest way to do this is to lose electrons and for the elements on the right of the periodic table the easiest way to do this is to gain electrons. So the elements on the left of the periodic table will form cations and the elements on the right hand side of the periodic table will form anions. By doing this the elements can be in the most stable electronic configuration and so be as stable as the noble gases.

Look at the following examples. Notice the number of valence electrons in the neutral atom, the number of electrons that are lost or gained and the final charge of the ion that is formed.

Lithium: A lithium atom loses one electron to form a positive ion:

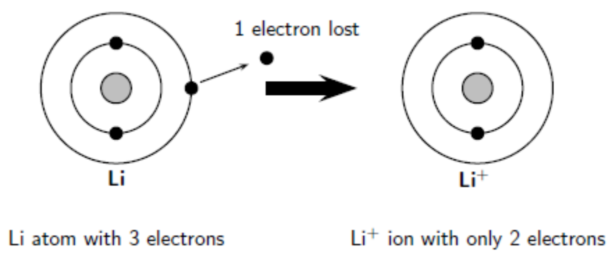


Figure 5.5: The arrangement of electrons in a lithium ion.

In this example, the lithium atom loses an electron to form the cation Li^{+}.

Fluorine: A fluorine atom gains one electron to form a negative ion:

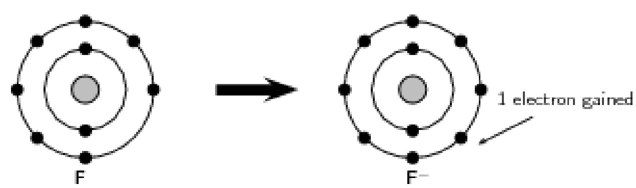


Figure 5.6: The arrangement of electrons in a fluorine ion.

You should have noticed in both these examples that each element lost or gained electrons to make a full outer shell.

Investigation : The formation of ions

1. Use the diagram for lithium as a guide and draw similar diagrams to show how each of the following ions is formed:
   a. Mg^{2+}
   b. Na^{+}
   c. Cl^{-}
   d. O^{2+}
2. Do you notice anything interesting about the charge on each of these ions? Hint: Look at the number of valence electrons in the neutral atom and the charge on the final ion.

Observations:
Once you have completed the activity, you should notice that:

- In each case the number of electrons that is either gained or lost, is the same as the number of electrons that are needed for the atoms to achieve a full outer energy level.
- If you look at an energy level diagram for sodium (Na), you will see that in a neutral atom, there is only one valence electron. In order to achieve a full outer energy level, and therefore a more stable state for the atom, this electron will be *lost*.
- In the case of oxygen (O), there are six valence electrons. To achieve a full energy level, it makes more sense for this atom to *gain* two electrons. A negative ion is formed.

Exercise: The formation of ions

Match the information in column A with the information in column B by writing only the letter (A to I) next to the question number (1 to 7)

1. A positive ion that has 3 less electrons than its neutral atom	A. Mg^{2+}
2. An ion that has 1 more electron than its neutral atom	B. Cl^{-}
3. The anion that is formed when bromine gains an electron	C. CO_3^{2-}
4. The cation that is formed from a magnesium atom	D. Al^{3+}
5. An example of a compound ion	E. Br^{2-}
6. A positive ion with the electron configuration of argon	F. K^{+}
7. A negative ion with the electron configuration of neon	G. Mg^{+}
	H. O^{2-}
	I. Br^{-}

Table 5.1

Find the answers with the shortcodes:

(1.) lid

1.39.2 Ionisation Energy

Ionisation energy is the energy that is needed to remove one electron from an atom. The ionisation energy will be different for different atoms.
The second ionisation energy is the energy that is needed to remove a second electron from an atom, and so on. As an energy level becomes more full, it becomes more and more difficult to remove an electron and the ionisation energy *increases*. On the Periodic Table of the Elements, a *group* is a vertical column of the elements, and a *period* is a horizontal row. In the periodic table, ionisation energy *increases* across a period, but *decreases* as you move down a group. The lower the ionisation energy, the more reactive the element will be because there is a greater chance of electrons being involved in chemical reactions. We will look at this in more detail in the next section.

Trends in ionisation energy

Refer to the data table below which gives the ionisation energy (in $kJ \cdot mol^{-1}$) and atomic number (Z) for a number of elements in the periodic table:

Z	Ionisation energy	Z	Ionisation energy
1	1310	10	2072
2	2360	11	494
3	517	12	734
4	895	13	575
5	797	14	783
6	1087	15	1051
7	1397	16	994
8	1307	17	1250
9	1673	18	1540

Table 5.2

1. Draw a line graph to show the relationship between atomic number (on the x-axis) and ionisation energy (y-axis).
2. Describe any trends that you observe.
3. Explain why...
   a. the ionisation energy for $Z = 2$ is higher than for $Z = 1$
   b. the ionisation energy for $Z = 3$ is lower than for $Z = 2$
   c. the ionisation energy increases between $Z = 5$ and $Z = 7$

www Find the answers with the shortcodes:

(1.) liv

Khan academy video on periodic table - 2 www (Video: P10034)

By now you should have an appreciation of what the periodic table can tell us. The periodic table does not just list the elements, but tells chemists what the properties of elements are, how the elements will combine and many other useful facts. The periodic table is truly an amazing resource. Into one simple table, chemists have packed so many facts and data that can easily be seen with a glance. The periodic table is a crucial part of chemistry and you should never go to science class without it. The following presentation provides a summary of the periodic table

www (Presentation: P10035)

1.40 Summary

(section shortcode: C10036)

- Elements are arranged in periods and groups on the periodic table. The elements are arranged according to increasing atomic number.
- A **group** is a column on the periodic table containing elements with similar properties. A **period** is a row on the periodic table.
- The groups on the periodic table are labeled from 1 to 8. The first group is known as the alkali metals, the second group is known as the alkali earth metals, the seventh group is known as the halogens and the eighth group is known as the noble gases. Each group has the same properties.
- Several trends such as ionisation energy and atomic diameter can be seen across the periods of the periodic table
- An **ion** is a charged atom. A **cation** is a positively charged ion and an **anion** is a negatively charged ion.
- When forming an ion, an atom will lose or gain the number of electrons that will make its valence energy level full.
- An element's **ionisation energy** is the energy that is needed to remove one electron from an atom.
- Ionisation energy increases across a **period** in the periodic table.
- Ionisation energy decreases down a **group** in the periodic table.

1.41 End of chapter exercises

(section shortcode: C10037)

1. For the following questions state whether they are true or false. If they are false, correct the statement.

   a. The group 1 elements are sometimes known as the alkali earth metals.
   b. The group 2 elements tend to lose 2 electrons to form cations.
   c. The group 8 elements are known as the noble gases.
   d. Group 7 elements are very unreactive.
   e. The transition elements are found between groups 3 and 4.
2. Give one word or term for each of the following:
   a. A positive ion
   b. The energy that is needed to remove one electron from an atom
   c. A horizontal row on the periodic table
   d. A very reactive group of elements that is missing just one electron from their outer shells.
3. For each of the following elements give the ion that will be formed:
   a. sodium
   b. bromine
   c. magnesium
   d. oxygen

4. The following table shows the first ionisation energies for the elements of period 1 and 2.

Period	Element	First ionisation energy (kJ.mol^{-1})
1	H	1312
	He	2372
	Li	520
	Be	899
	B	801
	C	1086
2	N	1402
	O	1314
	F	1681
	Ne	2081

Table 5.3

a. What is the meaning of the term *first ionisation energy*?
b. Identify the pattern of first ionisation energies in a period.
c. Which TWO elements exert the strongest attractive forces on their electrons? Use the data in the table to give a reason for your answer.
d. Draw Aufbau diagrams for the TWO elements you listed in the previous question and explain why these elements are so stable.
e. It is safer to use helium gas than hydrogen gas in balloons. Which property of helium makes it a safer option?
f. 'Group 1 elements readily form positive ions'. Is this statement correct? Explain your answer by referring to the table.

www Find the answers with the shortcodes:

(1.) l4Z (2.) l4K (3.) l4B (4.) li6

Chemical Bonding

1.42 Chemical Bonding

www (section shortcode: C10038)

When you look at the matter, or physical substances, around you, you will realise that atoms seldom exist on their own. More often, the things around us are made up of different atoms that have been joined together. This is called **chemical bonding.** Chemical bonding is one of the most important processes in chemistry because it allows all sorts of different molecules and combinations of atoms to form, which then make up the objects in the complex world around us.

1.43 What happens when atoms bond?

www (section shortcode: C10039)

A **chemical bond** is formed when atoms are held together by attractive forces. This attraction occurs when electrons are *shared* between atoms, or when electrons are *exchanged* between the atoms that are involved in the bond. The sharing or exchange of electrons takes place so that the outer energy levels of the atoms involved are filled and the atoms are more stable. If an electron is **shared**, it means that it will spend its time moving in the electron orbitals around *both* atoms. If an electron is **exchanged** it means that it is transferred from one atom to another, in other words one atom *gains* an electron while the other *loses* an electron.

Definition: Chemical bond
A chemical bond is the physical process that causes atoms and molecules to be attracted to each other, and held together in more stable chemical compounds.

The type of bond that is formed depends on the elements that are involved. In this chapter, we will be looking at three types of chemical bonding: **covalent**, **ionic** and **metallic bonding**.

You need to remember that it is the *valence electrons* that are involved in bonding and that atoms will try to fill their outer energy levels so that they are more stable (or are more like the noble gases which are very stable).

1.44 Covalent Bonding

(section shortcode: C10040)

1.44.1 The nature of the covalent bond

Covalent bonding occurs between the atoms of **non-metals**. The outermost orbitals of the atoms overlap so that unpaired electrons in each of the bonding atoms can be shared. By overlapping orbitals, the outer energy shells of all the bonding atoms are filled. The shared electrons move in the orbitals around *both* atoms. As they move, there is an attraction between these negatively charged electrons and the positively charged nuclei, and this force holds the atoms together in a covalent bond.

Definition: Covalent bond
Covalent bonding is a form of chemical bonding where pairs of electrons are shared between atoms.

Below are a few examples. Remember that it is only the *valence electrons* that are involved in bonding, and so when diagrams are drawn to show what is happening during bonding, it is only these electrons that are shown. Circles and crosses are used to represent electrons in different atoms.

Exercise 6.1: Covalent bonding How do hydrogen and chlorine atoms bond covalently in a molecule of hydrogen chloride?

Solution to Exercise

Step 1. A chlorine atom has 17 electrons, and an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^5$. A hydrogen atom has only 1 electron, and an electron configuration of $1s^1$.

Step 2. Chlorine has 7 valence electrons. One of these electrons is unpaired. Hydrogen has 1 valence electron and it is unpaired.

Step 3. The hydrogen atom needs one more electron to complete its valence shell. The chlorine atom also needs one more electron to complete its shell. Therefore *one pair of electrons* must be shared between the two atoms. In other words, one electron from the chlorine atom will spend some of its time orbiting the hydrogen atom so that hydrogen's valence shell is full. The hydrogen electron will spend some of its time orbiting the chlorine atom so that chlorine's valence shell is also full. A molecule of hydrogen chloride is formed (Figure 6.1). Notice the shared electron pair in the overlapping orbitals.

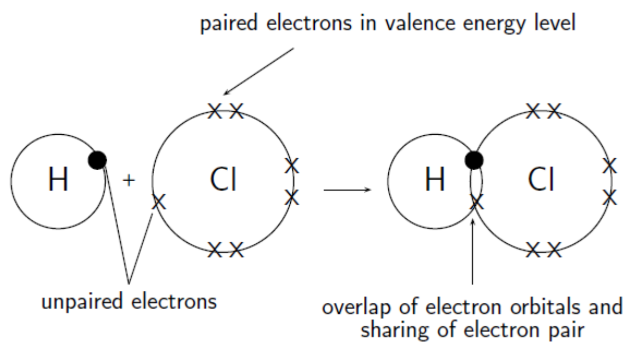


Figure 6.1: Covalent bonding in a molecule of hydrogen chloride

Exercise 6.2: Covalent bonding involving multiple bonds How do nitrogen and hydrogen atoms bond to form a molecule of ammonia (NH_3)?

Solution to Exercise

Step 1. A nitrogen atom has 7 electrons, and an electron configuration of $1s^2 2s^2 2p^3$. A hydrogen atom has only 1 electron, and an electron configuration of $1s^1$.

Step 2. Nitrogen has 5 valence electrons meaning that 3 electrons are unpaired. Hydrogen has 1 valence electron and it is unpaired.

Step 3. Each hydrogen atom needs one more electron to complete its valence energy shell. The nitrogen atom needs three more electrons to complete its valence energy shell. Therefore *three pairs of electrons* must be shared between the four atoms involved. The nitrogen atom will share three of its electrons so that each of the hydrogen atoms now have a complete valence shell. Each of the hydrogen atoms will share its electron with the nitrogen atom to complete its valence shell (Figure 6.2).

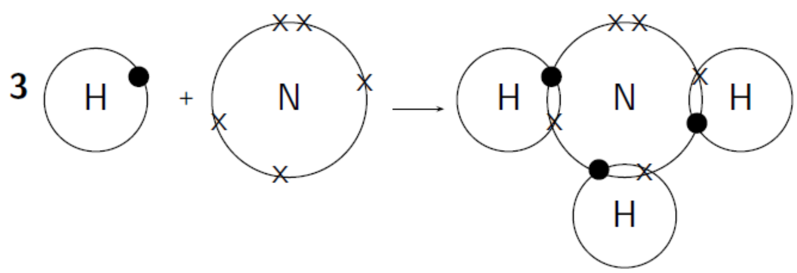


Figure 6.2: Covalent bonding in a molecule of ammonia

The above examples all show **single covalent bonds**, where only one pair of electrons is shared between *the same two atoms*. If two pairs of electrons are shared between the same two atoms, this is called a **double bond**. A **triple bond** is formed if three pairs of electrons are shared.

Exercise 6.3: Covalent bonding involving a double bond How do oxygen atoms bond covalently to form an oxygen molecule?

Solution to Exercise

Step 1. Each oxygen atom has 8 electrons, and their electron configuration is $1s^2 2s^2 2p^4$.

Step 2. Each oxygen atom has 6 valence electrons, meaning that each atom has 2 unpaired electrons.

Step 3. Each oxygen atom needs two more electrons to complete its valence energy shell. Therefore *two pairs of electrons* must be shared between the two oxygen atoms so that both valence shells are full. Notice that the two electron pairs are being shared between *the same two* atoms, and so we call this a **double bond** (Figure 6.3).

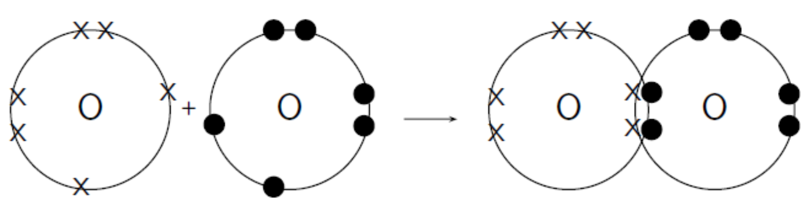


Figure 6.3: A double covalent bond in an oxygen molecule

You will have noticed in the above examples that the number of electrons that are involved in bonding varies between atoms. We say that the **valency** of the atoms is different.

Definition: Valency
The number of electrons in the outer shell of an atom which are able to be used to form bonds with other atoms.

In the first example, the valency of both hydrogen and chlorine is one, therefore there is a single covalent bond between these two atoms. In the second example, nitrogen has a valency of three and hydrogen has a valency of one. This means that three hydrogen atoms will need to bond with a single nitrogen atom. There are three *single* covalent bonds in a molecule of ammonia. In the third example, the valency of oxygen is two. This means that each oxygen atom will form two bonds with another atom. Since there is only one other atom in a molecule of O_2, a *double covalent* bond is formed between these two atoms.

TIP: There is a relationship between the valency of an element and its position on the Periodic Table. For the elements in groups 1 to 4, the valency is the same as the group number. For elements in groups 5 to 7, the valency is calculated by subtracting the group number from 8. For example, the valency of fluorine (group 7) is $8 - 7 = 1$, while the valency of calcium (group 2) is 2. Some elements have more than one possible valency, so you always need to be careful when you are writing a chemical formula. Often, if there is more than one possibility in terms of valency, the valency will be written in a bracket after the element symbol e.g. carbon (IV) oxide, means that in this molecule carbon has a valency of 4.

Covalent bonding and valency

1. Explain the difference between the *valence electrons* and the *valency* of an element.
2. Complete the table below by filling in the number of valence electrons and the valency for each of the elements shown:

Element	No. of valence electrons	No. of electrons needed to fill outer shell	Valency
F			
Ar			
C			
N			
O			

Table 6.1

3. Draw simple diagrams to show how electrons are arranged in the following covalent molecules:
   a. Water (H_2O)
   b. Chlorine (Cl_2)

Find the answers with the shortcodes:

(1.) IOY (2.) IOr (3.) IO1

1.44.2 Properties of covalent compounds

Covalent compounds have several properties that distinguish them from ionic compounds and metals. These properties are:

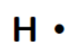
1. The melting and boiling points of covalent compounds is generally lower than that for ionic compounds.
2. Covalent compounds are generally more flexible than ionic compounds. The molecules in covalent compounds are able to move around to some extent and can sometimes slide over each other (as is the case with graphite, this is why the lead in your pencil feels slightly slippery). In ionic compounds all the ions are tightly held in place.
3. Covalent compounds generally are not very soluble in water.
4. Covalent compounds generally do not conduct electricity when dissolved in water. This is because they do not dissociate as ionic compounds do.

1.45 Lewis notation and molecular structure

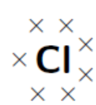
(section shortcode: C10041)

Although we have used diagrams to show the structure of molecules, there are other forms of notation that can be used, such as **Lewis notation** and **Couper notation**. **Lewis notation** uses dots and crosses to represent the **valence electrons** on different atoms. The chemical symbol of the element is used to represent the nucleus and the core electrons of the atom.

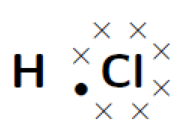
So, for example, a hydrogen atom would be represented like this:



A chlorine atom would look like this:



A molecule of hydrogen chloride would be shown like this:

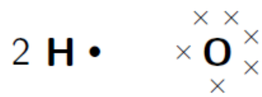


The dot and cross in between the two atoms, represent the pair of electrons that are shared in the covalent bond.

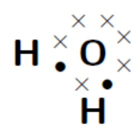
Exercise 6.4: Lewis notation: Simple molecules Represent the molecule H_2O using Lewis notation

Solution to Exercise

Step 1. The electron configuration of hydrogen is $1s^1$ and the electron configuration for oxygen is $1s^2 2s^2 2p^4$. Each hydrogen atom has one valence electron, which is unpaired, and the oxygen atom has six valence electrons with two unpaired.



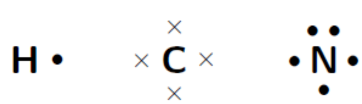
Step 2. The water molecule is represented below.



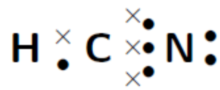
Exercise 6.5: Lewis notation: Molecules with multiple bonds Represent the molecule HCN (hydrogen cyanide) using Lewis notation

Solution to Exercise

Step 1. The electron configuration of hydrogen is $1s^1$, the electron configuration of nitrogen is $1s^2 2s^2 2p^3$ and for carbon is $1s^2 2s^2 2p^2$. This means that hydrogen has one valence electron which is unpaired, carbon has four valence electrons, all of which are unpaired, and nitrogen has five valence electrons, three of which are unpaired.



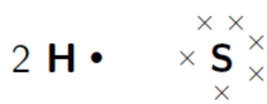
Step 2. The HCN molecule is represented below. Notice the three electron pairs between the nitrogen and carbon atom. Because these three covalent bonds are between the same two atoms, this is a *triple* bond.



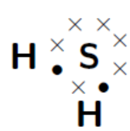
Exercise 6.6: Lewis notation: Atoms with variable valencies Represent the molecule $\mathrm{H_2S}$ (hydrogen sulphide) using Lewis notation

Solution to Exercise

Step 1. Hydrogen has an electron configuration of $\mathrm{1s^1}$ and sulphur has an electron configuration of $\mathrm{1s^2 2s^2 2}p^6\mathrm{3s^2 3p^4}$. Each hydrogen atom has one valence electron which is unpaired, and sulphur has six valence electrons. Although sulphur has a variable valency, we know that the sulphur will be able to form 2 bonds with the hydrogen atoms. In this case, the valency of sulphur must be two.



Step 2. The $\mathrm{H_2S}$ molecule is represented below.



Another way of representing molecules is using **Couper notation**. In this case, only the electrons that are involved in the bond between the atoms are shown. A line is used for each covalent bond. Using Couper notation, a molecule of water and a molecule of HCN would be represented as shown in Figure 6.13 and Figure 6.14 below.

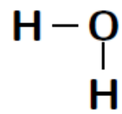


Figure 6.13: A water molecule represented using Couper notation

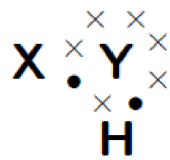
Figure 6.14: A molecule of HCN represented using Couper notation

1.45.1 Atomic bonding and Lewis notation

1. Represent each of the following *atoms* using Lewis notation:
   a. beryllium
   b. calcium
   c. lithium
2. Represent each of the following *molecules* using Lewis notation:
   a. bromine gas (Br_2)
   b. carbon dioxide (CO_2)
3. Which of the two molecules listed above contains a double bond?
4. Two chemical reactions are described below.
   - nitrogen and hydrogen react to form NH_3
   - carbon and hydrogen bond to form a molecule of CH_4

   For each reaction, give:
   a. the valency of each of the atoms involved in the reaction
   b. the Lewis structure of the product that is formed
   c. the chemical formula of the product
   d. the name of the product

5. A chemical compound has the following Lewis notation:



a. How many valence electrons does element Y have?
b. What is the valency of element Y?
c. What is the valency of element X?
d. How many covalent bonds are in the molecule?
e. Suggest a name for the elements X and Y.

Find the answers with the shortcodes:

(1.) IOC (2.) IOa (3.) IOa (4.) IOx (5.) IOc

1.46 Ionic Bonding

(section shortcode: C10042)

1.46.1 The nature of the ionic bond

You will remember that when atoms bond, electrons are either *shared* or they are *transferred* between the atoms that are bonding. In covalent bonding, electrons are shared between the atoms. There is another type of bonding, where electrons are *transferred* from one atom to another. This is called **ionic bonding**.

Ionic bonding takes place when the difference in electronegativity between the two atoms is more than 1,7. This usually happens when a metal atom bonds with a non-metal atom. When the difference in electronegativity is large, one atom will attract the shared electron pair much more strongly than the other, causing electrons to be transferred from one atom to the other.

Definition: Ionic bond
An ionic bond is a type of chemical bond based on the electrostatic forces between two oppositely-charged ions. When ionic bonds form, a metal donates one or more electrons, due to having a low electronegativity, to form a positive ion or cation. The non-metal atom has a high electronegativity, and therefore readily gains electrons to form a negative ion or anion. The two ions are then attracted to each other by electrostatic forces.

Example 1:

In the case of $NaCl$, the difference in electronegativity is 2,1. Sodium has only one valence electron, while chlorine has seven. Because the electronegativity of chlorine is higher than the electronegativity of sodium, chlorine will attract the valence electron of the sodium atom very strongly. This electron from sodium is transferred to chlorine. Sodium loses an electron and forms an Na^+ ion. Chlorine gains an electron and forms an Cl^- ion. The attractive force between the positive and negative ion holds the molecule together.

The balanced equation for the reaction is:

$$\mathrm{Na} + \mathrm{Cl} \rightarrow \mathrm{NaCl} \tag{6.1}$$

This can be represented using Lewis notation:

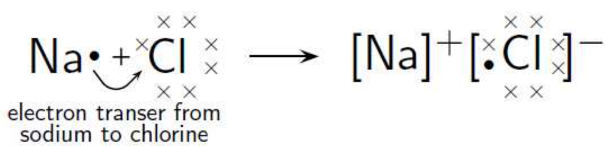


Figure 6.16: Ionic bonding in sodium chloride

Example 2:

Another example of ionic bonding takes place between magnesium (Mg) and oxygen (O) to form magnesium oxide (MgO). Magnesium has two valence electrons and an electronegativity of 1,2, while oxygen has six valence electrons and an electronegativity of 3,5. Since oxygen has a higher electronegativity, it attracts the two valence electrons from the magnesium atom and these electrons are transferred from the magnesium atom to the oxygen atom. Magnesium loses two electrons to form Mg^{2+}, and oxygen gains two electrons to form O^{2-}. The attractive force between the oppositely charged ions is what holds the molecule together.

The balanced equation for the reaction is:

$$2\mathrm{Mg} + \mathrm{O_2} \rightarrow 2\mathrm{MgO} \tag{6.2}$$

Because oxygen is a diatomic molecule, two magnesium atoms will be needed to combine with one oxygen molecule (which has two oxygen atoms) to produce two molecules of magnesium oxide (MgO).

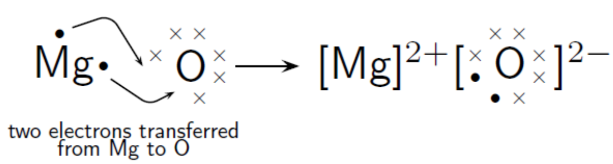


Figure 6.17: Ionic bonding in magnesium oxide

TIP: Notice that the number of electrons that is either lost or gained by an atom during ionic bonding, is the same as the **valency** of that element

Ionic compounds

1. Explain the difference between a *covalent* and an *ionic* bond.

2. Magnesium and chlorine react to form magnesium chloride.
   a. What is the difference in electronegativity between these two elements?
   b. Give the chemical formula for:
      i. a magnesium ion
      ii. a chloride ion
      iii. the ionic compound that is produced during this reaction
   c. Write a balanced chemical equation for the reaction that takes place.
3. Draw Lewis diagrams to represent the following ionic compounds:
   a. sodium iodide (NaI)
   b. calcium bromide ($CaBr_2$)
   c. potassium chloride (KCl)

Find the answers with the shortcodes:

(1.) IOq (2.) IOl (3.) IOi

1.46.2 The crystal lattice structure of ionic compounds

Ionic substances are actually a combination of lots of ions bonded together into a giant molecule. The arrangement of ions in a regular, geometric structure is called a **crystal lattice**. So in fact $NaCl$ does not contain one Na and one Cl ion, but rather a lot of these two ions arranged in a crystal lattice where the ratio of Na to Cl ions is 1:1. The structure of a crystal lattice is shown in Figure 6.18.

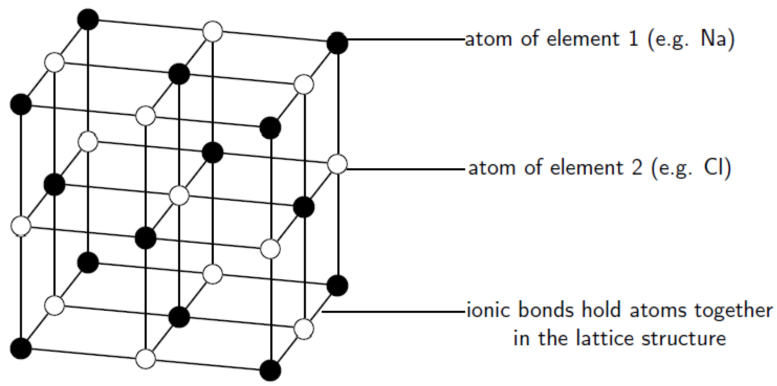


Figure 6.18: The crystal lattice arrangement in an ionic compound (e.g. $NaCl$)

1.46.3 Properties of Ionic Compounds

Ionic compounds have a number of properties:

- Ions are arranged in a lattice structure
- Ionic solids are crystalline at room temperature
- The ionic bond is a strong electrical attraction. This means that ionic compounds are often hard and have high melting and boiling points
- Ionic compounds are brittle, and bonds are broken along planes when the compound is stressed
- Solid crystals don't conduct electricity, but ionic solutions do

1.47 Metallic bonds

(section shortcode: C10043)

1.47.1 The nature of the metallic bond

The structure of a metallic bond is quite different from covalent and ionic bonds. In a metal bond, the valence electrons are *delocalised*, meaning that an atom's electrons do not stay around that one nucleus. In a metallic bond, the positive atomic nuclei (sometimes called the 'atomic kernels') are surrounded by a sea of delocalised electrons which are attracted to the nuclei (Figure 6.19).

Definition: Metallic bond
Metallic bonding is the electrostatic attraction between the positively charged atomic nuclei of metal atoms and the delocalised electrons in the metal.

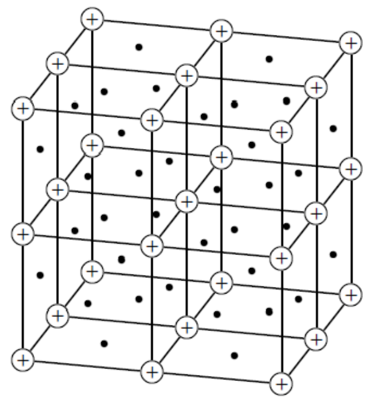


Figure 6.19: Positive atomic nuclei (+) surrounded by delocalised electrons (●)

1.47.2 The properties of metals

Metals have several unique properties as a result of this arrangement:

- *Thermal conductors* Metals are good conductors of heat and are therefore used in cooking utensils such as pots and pans. Because the electrons are loosely bound and are able to move, they can transport heat energy from one part of the material to another.
- *Electrical conductors* Metals are good conductors of electricity, and are therefore used in electrical conducting wires. The loosely bound electrons are able to move easily and to transfer charge from one part of the material to another.
- *Shiny metallic lustre* Metals have a characteristic shiny appearance and are often used to make jewellery. The loosely bound electrons are able to absorb and reflect light at all frequencies, making metals look polished and shiny.
- *Malleable and ductile* This means that they can be bent into shape without breaking (malleable) and can be stretched into thin wires (ductile) such as copper, which can then be used to conduct electricity. Because the bonds are not fixed in a particular direction, atoms can slide easily over one another, making metals easy to shape, mould or draw into threads.
- *Melting point* Metals usually have a high melting point and can therefore be used to make cooking pots and other equipment that needs to become very hot, without being damaged. The high melting point is due to the high strength of metallic bonds.
- *Density* Metals have a high density because their atoms are packed closely together.

1.47.3 Activity: Building models

Using coloured balls and sticks (or any other suitable materials) build models of each type of bonding. Think about how to represent each kind of bonding. For example, covalent bonding could be represented by simply connecting the balls with sticks to represent the molecules, while for ionic bonding you may wish to construct part of the crystal lattice. Do some research on types of crystal lattices (although the section on ionic bonding only showed the crystal lattice for sodium chloride, many other types of lattices exist) and try to build some of these. Share your findings with your class and compare notes to see what types of crystal lattices they found. How would you show metallic bonding?

You should spend some time doing this activity as it will really help you to understand how atoms combine to form molecules and what the differences are between the types of bonding.

Khan academy video on bonding - 1 www (Video: P10044)

1.47.4 Chemical bonding

1. Give two examples of everyday objects that contain..
   a. covalent bonds
   b. ionic bonds
   c. metallic bonds
2. Complete the table which compares the different types of bonding:

	Covalent	**Ionic**	**Metallic**
Types of atoms involved			
Nature of bond between atoms			
Melting Point (high/low)			
Conducts electricity? (yes/no)			
Other properties			

Table 6.2

3. Complete the table below by identifying the type of bond (covalent, ionic or metallic) in each of the compounds:

Molecular formula	**Type of bond**
H_2SO_4	
FeS	
NaI	
$MgCl_2$	
Zn	

Table 6.3

4. Which of these substances will conduct electricity most effectively? Give a reason for your answer.

5. Use your knowledge of the different types of bonding to explain the following statements:
   a. Swimming during an electric storm (i.e. where there is lightning) can be very dangerous.
   b. Most jewellery items are made from metals.
   c. Plastics are good insulators.

Find the answers with the shortcodes:

(1.) l3h (2.) l3u (3.) l3J (4.) l3J (5.) l3S

1.48 Writing chemical formulae

(section shortcode: C10045)

1.48.1 The formulae of covalent compounds

To work out the formulae of covalent compounds, we need to use the valency of the atoms in the compound. This is because the valency tells us how many bonds each atom can form. This in turn can help to work out how many atoms of each element are in the compound, and therefore what its formula is. The following are some examples where this information is used to write the chemical formula of a compound.

Exercise 6.7: Formulae of covalent compounds Write the chemical formula for water

Solution to Exercise

Step 1. A molecule of water contains the elements *hydrogen* and *oxygen*.

Step 2. The valency of hydrogen is 1 and the valency of oxygen is 2. This means that oxygen can form two bonds with other elements and each of the hydrogen atoms can form one.

Step 3. Using the valencies of hydrogen and oxygen, we know that in a single water molecule, two hydrogen atoms will combine with one oxygen atom. The chemical formula for water is therefore: H_2O.

Exercise 6.8: Formulae of covalent compounds Write the chemical formula for magnesium oxide

Solution to Exercise

Step 1. A molecule of magnesium oxide contains the elements *magnesium* and *oxygen*.

Step 2. The valency of magnesium is 2, while the valency of oxygen is also 2. In a molecule of magnesium oxide, one atom of magnesium will combine with one atom of oxygen.

Step 3. The chemical formula for magnesium oxide is therefore:
MgO

Exercise 6.9: Formulae of covalent compounds Write the chemical formula for copper (II) chloride.

Solution to Exercise

Step 1. A molecule of copper (II) chloride contains the elements *copper* and *chlorine*.

Step 2. The valency of copper is 2, while the valency of chlorine is 1. In a molecule of copper (II) chloride, two atoms of chlorine will combine with one atom of copper.

Step 3. The chemical formula for copper (II) chloride is therefore:
$CuCl_2$

1.48.2 The formulae of ionic compounds

The overall charge of an ionic compound will always be zero and so the negative and positive charge must be the same size. We can use this information to work out what the chemical formula of an ionic compound is if we know the charge on the individual ions. In the case of $NaCl$ for example, the charge on the sodium is $+1$ and the charge on the chlorine is -1. The charges balance ($+1 - 1 = 0$) and therefore the ionic compound is neutral. In MgO, magnesium has a charge of $+2$ and oxygen has a charge of -2. Again, the charges balance and the compound is neutral. Positive ions are called **cations** and negative ions are called **anions**.

Some ions are made up of groups of atoms, and these are called **compound ions**. It is a good idea to learn the compound ions that are shown in Table 6.4

Name of compound ion	**formula**	**Name of compound ion**	**formula**
Carbonate	CO_3^{2-}	Nitrate	NO_2^-
Sulphate	SO_4^{2-}	Hydrogen sulphite	HSO_3^-
Hydroxide	OH^-	Hydrogen sulphate	HSO_4^-
Ammonium	NH_4^+	Dihydrogen phosphate	$H_2PO_4^-$
Nitrate	NO_3^-	Hypochlorite	ClO^-
Hydrogen carbonate	HCO_3^-	Acetate (ethanoate)	CH_3COO^-
Phosphate	PO_4^{3-}	Oxalate	$C_2O_4^{2-}$
Chlorate	ClO_3^-	Oxide	O^{2-}
Cyanide	CN^-	Peroxide	O_2^{2-}
Chromate	CrO_4^{2-}	Sulphide	S^{2-}
Permanganate	MnO_4^-	Sulphite	SO_3^{2-}
Thiosulphate	$S_2O_3^{2-}$	Manganate	MnO_4^{2-}
Phosphide	P^{3-}	Hydrogen phosphate	HPO_4^{3-}

Table 6.4: Table showing common compound ions and their formulae

In the case of ionic compounds, the valency of an ion is the same as its charge (Note: valency is always expressed as a *positive* number e.g. valency of the chloride ion is 1 and not -1). Since an ionic compound is always *neutral*, the positive charges in the compound must balance out the negative. The following are some examples:

Exercise 6.10: Formulae of ionic compounds Write the chemical formula for potassium iodide.

Solution to Exercise

Step 1. Potassium iodide contains potassium and iodide ions.

Step 2. Potassium iodide contains the ions K^+ $(\text{valency} = 1; \text{charge} = +1)$ and I^- $(\text{valency} = 1; \text{charge} = -1)$. In order to balance the charge in a single molecule, one atom of potassium will be needed for every one atom of iodine.

Step 3. The chemical formula for potassium iodide is therefore: KI

Exercise 6.11: Formulae of ionic compounds Write the chemical formula for sodium sulphate.

Solution to Exercise

Step 1. Sodium sulphate contains sodium ions and sulphate ions.

Step 2. $\mathrm{Na^+}$ (valency = 1; charge = +1) and $\mathrm{SO_4^{2-}}(\mathrm{valency} = 2; \mathrm{charge} = -2)$.

Step 3. Two sodium ions will be needed to balance the charge of the sulphate ion. The chemical formula for sodium sulphate is therefore: $\mathrm{Na_2SO_4}$

Exercise 6.12: Formulae of ionic compounds Write the chemical formula for calcium hydroxide.

Solution to Exercise

Step 1. Calcium hydroxide contains calcium ions and hydroxide ions.

Step 2. Calcium hydroxide contains the ions $\mathrm{Ca^{2+}}$ $(\mathrm{charge} = +2)$ and $\mathrm{OH^-}$ $(\mathrm{charge} = -1)$. In order to balance the charge in a single molecule, two hydroxide ions will be needed for every ion of calcium.

Step 3. The chemical formula for calcium hydroxide is therefore: $\mathrm{Ca\,(OH)_2}$

NOTE: Notice how in the last example we wrote $\mathrm{OH^-}$ inside brackets. We do this to indicate that $\mathrm{OH^-}$ is a complex ion and that there are two of these ions bonded to one calcium ion.

Chemical formulae

1. Copy and complete the table below:

Compound	Cation	Anion	Formula
	$\mathrm{Na^+}$	$\mathrm{Cl^-}$	
potassium bromide		$\mathrm{Br^-}$	
	$\mathrm{NH_4^+}$	$\mathrm{Cl^-}$	
potassium chromate			
			PbI
potassium permanganate			
calcium phosphate			

Table 6.5

2. Write the chemical formula for each of the following compounds:
   a. hydrogen cyanide
   b. carbon dioxide
   c. sodium carbonate
   d. ammonium hydroxide
   e. barium sulphate

Find the answers with the shortcodes:

(1.) l3t (2.) l3z

1.49 Chemical compounds: names and masses

(section shortcode: C10046)

In Giving names and formulae to substances (Section 1.4: Giving names and formulae to substances) the names of chemical compounds was revised. The relative molecular mass for covalent molecules is simply the sum of the relative atomic masses of each of the individual atoms in that compound. For ionic compounds we use the formula of the compound to work out a relative formula mass. We ignore the fact that there are many molecules linked together to form a crystal lattice. For example NaCl has a relative formula mass of $58\ \mathrm{g} \cdot \mathrm{mol}^{-1}$.

(Presentation: P10047)

1.50 Summary

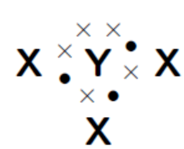
(section shortcode: C10048)

- A **chemical bond** is the physical process that causes atoms and molecules to be attracted together and to be bound in new compounds.
- Atoms are more **reactive**, and therefore more likely to bond, when their outer electron orbitals are not full. Atoms are less reactive when these outer orbitals contain the maximum number of electrons. This explains why the noble gases do not combine to form molecules.
- When atoms bond, electrons are either shared or exchanged.
- **Covalent bonding** occurs between the atoms of non-metals and involves a sharing of electrons so that the orbitals of the outermost energy levels in the atoms are filled.
- The **valency** of an atom is the number of electrons in the outer shell of that atom and valence electrons are able to form bonds with other atoms.
- A **double** or **triple bond** occurs if there are two or three electron pairs that are shared between the same two atoms.
- A **dative covalent bond** is a bond between two atoms in which both the electrons that are shared in the bond come from the same atom.
- **Lewis** and **Couper** notation are two ways of representing molecular structure. In Lewis notation, dots and crosses are used to represent the valence electrons around the central atom. In Couper notation, lines are used to represent the bonds between atoms.

- An **ionic bond** occurs between atoms where the difference in electronegativity is greater than 1,7. An exchange of electrons takes place and the atoms are held together by the electrostatic force of attraction between oppositely-charged ions.
- Ionic solids are arranged in a **crystal lattice** structure.
- Ionic compounds have a number of specific **properties**, including their high melting and boiling points, brittle nature, the lattice structure of solids and the ability of ionic solutions to conduct electricity.
- A **metallic bond** is the electrostatic attraction between the positively charged nuclei of metal atoms and the delocalised electrons in the metal.
- Metals also have a number of properties, including their ability to conduct heat and electricity, their metallic lustre, the fact that they are both malleable and ductile, and their high melting point and density.
- The valency of atoms, and the way they bond, can be used to determine the **chemical formulae** of compounds.

1.50.1 End of chapter exercises

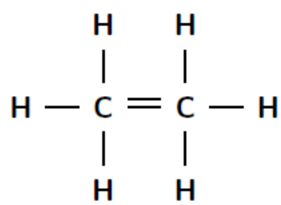
1. Explain the meaning of each of the following terms
   a. Valency
   b. Covalent bond
2. Which ONE of the following best describes the bond formed between an H^+ ion and the NH_3 molecule?
   a. Covalent bond
   b. Dative covalent (coordinate covalent) bond
   c. Ionic Bond
   d. Hydrogen Bond
3. Which of the following reactions will *not* take place? Explain your answer.
   a. $H + H \rightarrow H_2$
   b. $Ne + Ne \rightarrow Ne_2$
   c. $Cl + Cl \rightarrow Cl_2$
4. Draw the Lewis structure for each of the following:
   a. calcium
   b. iodine (Hint: Which group is it in? It will be similar to others in that group)
   c. hydrogen bromide (HBr)
   d. nitrogen dioxide (NO_2)
5. Given the following Lewis structure, where X and Y each represent a different element...



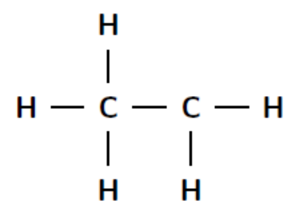
   a. What is the valency of X?
   b. What is the valency of Y?
   c. Which elements could X and Y represent?

6. A molecule of ethane has the formula C_2H_6. Which of the following diagrams (Couper notation) accurately represents this molecule?

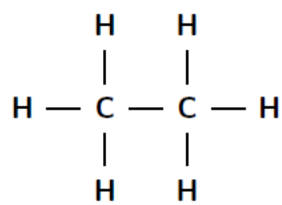
(a)



(b)



(c)



7. Potassium dichromate is dissolved in water.

a. Give the name and chemical formula for each of the ions in solution.

b. What is the chemical formula for potassium dichromate?

Find the answers with the shortcodes:

(1.) lgV (2.) l37 (3.) l36 (4.) l3H (5.) l3s (6.) l3o
(7.) l3A

What are the objects around us made of

1.51 Introduction: The atom as the building block of matter

(section shortcode: C10049)

We have now seen that different materials have different properties. Some materials are metals and some are non-metals; some are electrical or thermal conductors, while others are not. Depending on the properties of these materials, they can be used in lots of useful applications. But what is it exactly that makes up these materials? In other words, if we were to break down a material into the parts that make it up, what would we find? And how is it that a material's microscopic structure (the small parts that make up the material) is able to give it all these different properties?

The answer lies in the smallest building block of matter: the **atom**. It is the *type* of atoms, and the way in which they are *arranged* in a material, that affects the properties of that substance. This is similar to building materials. We can use bricks, steel, cement, wood, straw (thatch), mud and many other things to build structures from. In the same way that the choice of building material affects the properties of the structure, so the atoms that make up matter affect the properties of matter.

It is not often that substances are found in atomic form (just as you seldom find a building or structure made from one building material). Normally, atoms are bonded (joined) to other atoms to form **compounds** or **molecules**. It is only in the *noble gases* (e.g. helium, neon and argon) that atoms are found individually and are not bonded to other atoms. We looked at some of the reasons for this in earlier chapters.

1.52 Compounds

(section shortcode: C10050)

Definition: Compound
A compound is a group of two or more atoms that are attracted to each other by relatively strong forces or bonds.

Almost everything around us is made up of molecules. The only substances that are not made of molecules, but instead are individual atoms are the noble gases. *Water* is made up of molecules, each of which has two hydrogen atoms joined to one oxygen atom. *Oxygen* is a molecule that is made up of two oxygen atoms that are joined to one another. Even the food that we eat is made up of molecules that contain atoms of elements such as carbon, hydrogen and oxygen that are joined to one another in different ways. All of these are known as **small molecules** because there are only a few atoms in each molecule. **Giant molecules** are those where there may be millions of atoms per molecule. Examples of giant molecules are *diamonds*, which are made up of millions of carbon atoms bonded to each other and *metals*, which are made up of millions of metal atoms bonded to each other.

As we learnt in Chapter 5 atoms can share electrons to form covalent bonds or exchange electrons to form ionic bonds. Covalently bonded substances are known as molecular compounds. Ionically bonded substances are known as ionic compounds. We also learnt about metallic bonding. In a metal the atoms lose their outermost electrons to form positively charged ions that are arranged in a lattice, while the outermost electrons are free to move amongst the spaces of the lattice.

We can classify covalent molecules into covalent molecular structures and covalent network structures. Covalent molecular structures are simply individual covalent molecules and include water, oxygen, sulphur (S_8) and buckminsterfullerene (C_{60}). All covalent molecular structures are simple molecules. Covalent network structures are giant lattices of covalently bonded molecules, similar to the ionic lattice. Examples include diamond, graphite and silica (SiO_2). All covalent network structures are giant molecules.

Examples of ionic substances are sodium chloride ($NaCl$) and potassium permanganate ($KMnO_4$). Examples of metals are copper, zinc, titanium, gold, etc.

1.52.1 Representing molecules

The structure of a molecule can be shown in many different ways. Sometimes it is easiest to show what a molecule looks like by using different types of **diagrams**, but at other times, we may decide to simply represent a molecule using its **chemical formula** or its written name.

1. **Using formulae to show the structure of a molecule.** A **chemical formula** is an abbreviated (shortened) way of describing a molecule, or some other chemical substance. In the chapter on classification of matter, we saw how chemical compounds can be represented using element symbols from the Periodic Table. A chemical formula can also tell us the *number* of atoms of each element that are in a molecule and their *ratio* in that molecule. For example, the chemical formula for a molecule of carbon dioxide is CO_2 The formula above is called the **molecular formula** of that compound. The formula tells us that in one molecule of carbon dioxide, there is one atom of carbon and two atoms of oxygen. The ratio of carbon atoms to oxygen atoms is 1:2.

Definition: Molecular formula
This is a concise way of expressing information about the atoms that make up a particular chemical compound. The molecular formula gives the exact number of each type of atom in the molecule.

A molecule of glucose has the molecular formula: $C_6H_{12}O_6$. In each glucose molecule, there are six carbon atoms, twelve hydrogen atoms and six oxygen atoms. The ratio of carbon:hydrogen:oxygen is 6:12:6. We can simplify this ratio to write 1:2:1, or if we were to use the element symbols, the formula would be written as CH_2O. This is called the **empirical formula** of the molecule.

Definition: Empirical formula
This is a way of expressing the *relative* number of each type of atom in a chemical compound. In most cases, the empirical formula does not show the exact number of atoms, but rather the simplest *ratio* of the atoms in the compound.

The empirical formula is useful when we want to write the formula for a *giant molecule*. Since giant molecules may consist of millions of atoms, it is impossible to say exactly how many atoms are in each molecule. It makes sense then to represent these molecules using their empirical formula. So, in the case of a metal such as copper, we would simply write Cu, or if we were to represent a molecule of sodium chloride, we would simply write NaCl. Chemical formulae therefore tell us something about the *types* of atoms that are in a molecule and the *ratio* in which these atoms occur in the molecule, but they don't give us any idea of what the molecule actually looks like, in other words its *shape*. To show the shape of molecules we can represent molecules using diagrams. Another type of formula that can be used to describe a molecule is its **structural formula**. A structural formula uses a graphical representation to show a molecule's structure (Figure 7.1).

(a) C_4H_{10} (b) C_2H_5

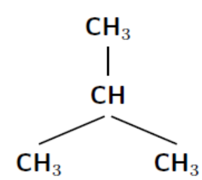


Figure 7.1: Diagram showing (a) the molecular, (b) the empirical and (c) the structural formula of isobutane

2. **Using diagrams to show the structure of a molecule** Diagrams of molecules are very useful because they help us to picture how the atoms are arranged in the molecule and they help us to see the shape of the molecule. There are two types of diagrams that are commonly used:
   - *Ball and stick models* This is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them. The centres of the atoms (the balls) are connected by straight lines which represent the bonds between them. A simplified example is shown in Figure 7.2.

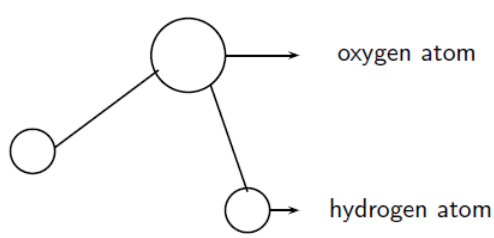


Figure 7.2: A ball and stick model of a water molecule

- *Space-filling model* This is also a 3-dimensional molecular model. The atoms are represented by spheres. Figure 7.3 and Figure 7.4 are some examples of **simple molecules** that are represented in different ways.

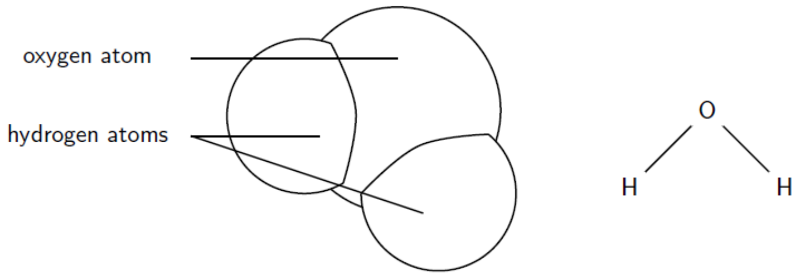


Figure 7.3: A space-filling model and structural formula of a water molecule. Each molecule is made up of two hydrogen atoms that are attached to one oxygen atom. This is a simple molecule.

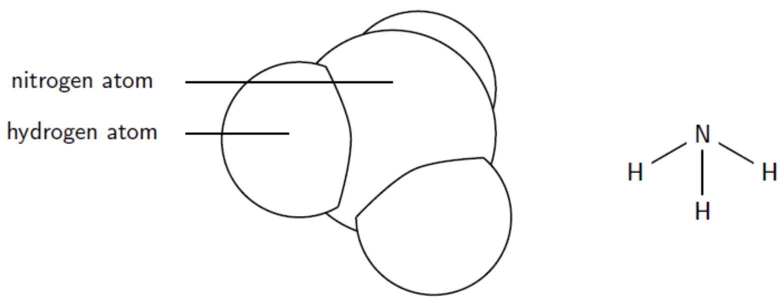


Figure 7.4: A space-filling model and structural formula of a molecule of ammonia. Each molecule is made up of one nitrogen atom and three hydrogen atoms. This is a simple molecule.

Figure 7.5 shows the bonds between the carbon atoms in diamond, which is a **giant molecule**. Each carbon atom is joined to four others, and this pattern repeats itself until a complex *lattice* structure is formed. Each black ball in the diagram represents a carbon atom, and each line represents the bond between two carbon atoms. Note that the carbon atoms on the edges are actually bonded to four carbon atoms, but some of these carbon atoms have been omitted.

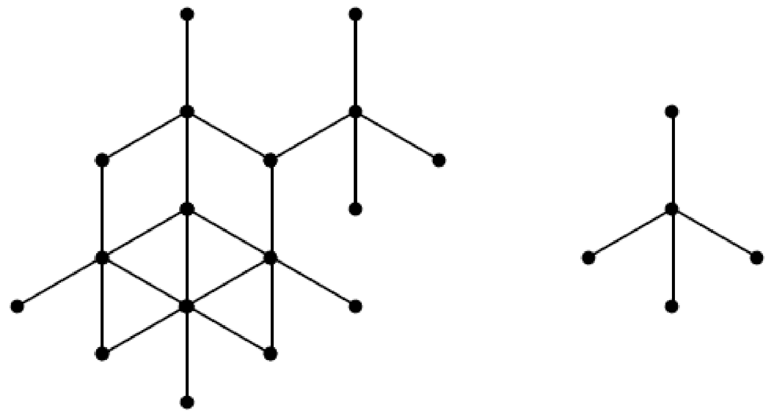


Figure 7.5: Diagrams showing the microscopic structure of diamond. The diagram on the left shows part of a diamond lattice, made up of numerous carbon atoms. The diagram on the right shows how each carbon atom in the lattice is joined to four others. This forms the basis of the lattice structure. Diamond is a giant molecule.

NOTE: Diamonds are most often thought of in terms of their use in the jewellery industry. However, about 80% of mined diamonds are unsuitable for use as gemstones and are therefore used in industry because of their strength and hardness. These properties of diamonds are due to the strong covalent bonds (covalent bonding will be explained later) between the carbon atoms in diamond. The most common uses for diamonds in industry are in cutting, drilling, grinding, and polishing.

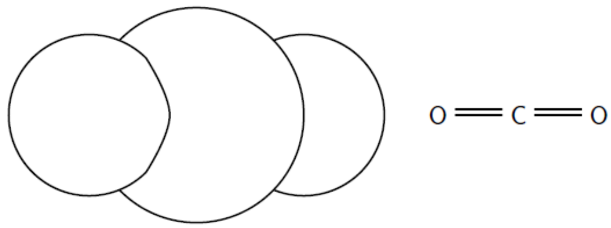
Alteredqualia is a website[3] that allows you to view several molecules. You do not need to know these molecules, this is simply to allow you to see one way of representing molecules.

Atoms and molecules

1. In each of the following, say whether the chemical substance is made up of single atoms, simple molecules or giant molecules.
   - **a.** ammonia gas (NH_3)
   - **b.** zinc metal (Zn)
   - **c.** graphite (C)
   - **d.** nitric acid (HNO_3)
   - **e.** neon gas (He)

[3] http://alteredqualia.com/canvasmol/

2. Refer to the diagram below and then answer the questions that follow:



   a. Identify the molecule.
   b. Write the molecular formula for the molecule.
   c. Is the molecule a simple or giant molecule?
3. Represent each of the following molecules using its *chemical formula*, *structural formula* and *ball and stick model*.
   a. Hydrogen
   b. Ammonia
   c. sulphur dioxide

Find the answers with the shortcodes:

(1.) li5 (2.) liN (3.) liR

1.53 Summary

(section shortcode: C10051)

- The smallest unit of matter is the **atom**. Atoms can combine to form **molecules**.
- A **compound** is a group of two or more atoms that are attracted to each other by chemical bonds.
- A **small molecule** consists of a few atoms per molecule. A **giant molecule** consists of millions of atoms per molecule, for example metals and diamonds.
- The structure of a molecule can be represented in a number of ways.
- The **chemical formula** of a molecule is an abbreviated way of showing a molecule, using the symbols for the elements in the molecule. There are two types of chemical formulae: molecular and empirical formula.
- The **molecular formula** of a molecule gives the exact number of atoms of each element that are in the molecule.
- The **empirical formula** of a molecule gives the relative number of atoms of each element in the molecule.
- Molecules can also be represented using **diagrams**.
- A **ball and stick** diagram is a 3-dimensional molecular model that uses 'balls' to represent atoms and 'sticks' to represent the bonds between them.
- A **space-filling model** is also a 3-dimensional molecular model. The atoms are represented by spheres.
- In a molecule, atoms are held together by **chemical bonds** or **intramolecular forces**. Covalent bonds, ionic bonds and metallic bonds are examples of chemical bonds.
- A **covalent bond** exists between non-metal atoms. An **ionic bond** exists between non-metal and metal atoms and a **metallic bond** exists between metal atoms.
- **Intermolecular forces** are the bonds that hold *molecules* together.

1.53.1 End of chapter exercises

1. Give one word or term for each of the following descriptions.
   a. A composition of two or more atoms that act as a unit.
   b. Chemical formula that gives the relative number of atoms of each element that are in a molecule.
2. Give a definition for each of the following terms: descriptions.
   a. molecule
   b. Ionic compound
   c. Covalent network structure
   d. Empirical formula
   e. Ball-and-stick model
3. Ammonia, an ingredient in household cleaners, can be broken down to form one part nitrogen (N) and three parts hydrogen (H). This means that ammonia...
   a. is a colourless gas
   b. is not a compound
   c. cannot be an element
   d. has the formula N_3H
4. Represent each of the following molecules using its chemical formula, its structural formula and the ball-and-stick model:
   a. nitrogen
   b. carbon dioxide
   c. methane
   d. argon

Find the answers with the shortcodes:

(1.) l2e (2.) l2M (3.) liV (4.) l2L

Physical and Chemical Change

1.54 Physical and Chemical Change - Grade 10

(section shortcode: C10052)

Matter is all around us. The desks we sit at, the air we breathe and the water we drink are all examples of matter. But matter doesn't always stay the same. It can change in many different ways. In this chapter, we are going to take a closer look at **physical** and **chemical** changes that occur in matter.

1.55 Physical changes in matter

(section shortcode: C10053)

A **physical change** is one where the particles of the substances that are involved in the change are not broken up in any way. When water is heated for example, the temperature and energy of the water molecules increases and the liquid water evaporates to form water vapour. When this happens, some kind of change has taken place, but the molecular structure of the water has not changed. This is an example of a *physical change*.

$\mathrm{H_2O\,(l) \rightarrow H_2O\,(g)}$

Conduction (the transfer of energy through a material) is another example of a physical change. As energy is transferred from one material to another, the *energy* of each material is changed, but not its chemical makeup. Dissolving one substance in another is also a physical change.

Definition: Physical change
A change that can be seen or felt, but that doesn't involve the break up of the particles in the reaction. During a physical change, the *form* of matter may change, but not its *identity*. A change in temperature is an example of a physical change.

You can think of a physical change as a person who is standing still. When they start to move (start walking) then a change has occurred and this is similar to a physical change.

There are some important things to remember about physical changes in matter:

1. *Arrangement of particles*
When a physical change occurs, the particles (e.g. atoms, molecules) may re-arrange themselves without actually breaking up in any way. In the example of evaporation that we used earlier, the water molecules move further apart as their temperature (and therefore energy) increases. The same would be true if ice were to melt. In the solid phase, water molecules are packed close together in a very ordered way, but when the ice is heated, the molecules overcome the forces holding them together and they move apart. Once again, the particles have re-arranged themselves, but have not broken up.

$$H_2O\,(s) \rightarrow H_2O\,(l) \tag{8.1}$$

Figure 8.1 shows this more clearly. In each phase of water, the water molecule itself stays the same, but the way the molecules are arranged has changed. Note that in the solid phase, we simply show the water molecules as spheres. This makes it easier to see how tightly packed the molecules are. In reality the water molecules would all look the same.

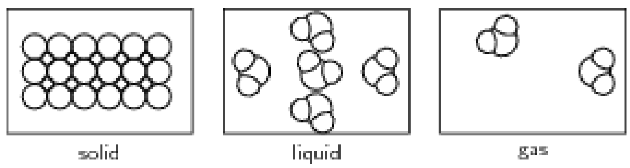


Figure 8.1: The arrangement of water molecules in the three phases of matter

2. *Conservation of mass*
In a physical change, the total mass, the number of atoms and the number of molecules will always stay the same. In other words you will always have the same number of molecules or atoms at the end of the change as you had at the beginning.
3. *Energy changes*
Energy changes may take place when there is a physical change in matter, but these energy changes are normally smaller than the energy changes that take place during a chemical change.
4. *Reversibility*
Physical changes in matter are usually easier to reverse than chemical changes. Water vapour for example, can be changed back to liquid water if the temperature is lowered. Liquid water can be changed into ice by simply decreasing the temperature.

Activity: Physical change Use plastic pellets or marbles to represent water in the solid state. What do you need to do to the pellets to represent the change from solid to liquid?

1.56 Chemical Changes in Matter

www (section shortcode: C10054)

When a **chemical change** takes place, new substances are formed in a chemical reaction. These new products may have very different properties from the substances that were there at the start of the reaction.

The breakdown of copper (II) chloride to form copper and chlorine is an example of chemical change. A simplified diagram of this reaction is shown in Figure 8.2. In this reaction, the initial substance is copper (II) chloride, but once the reaction is complete, the products are copper and chlorine.

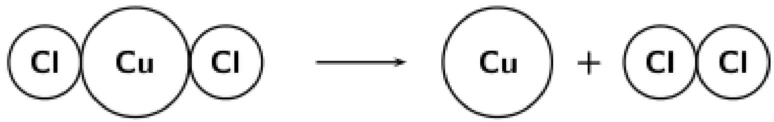


Figure 8.2: The decomposition of copper(II) chloride to form copper and chlorine. We write this as: $CuCl_2 \rightarrow Cu + Cl_2$

Definition: Chemical change
The formation of new substances in a chemical reaction. One type of matter is changed into something different.

There are some important things to remember about chemical changes:

1. *Arrangement of particles*
During a chemical change, the particles themselves are changed in some way. In the example of copper (II) chloride that was used earlier, the $CuCl_2$ molecules were split up into their component atoms. The number of particles will change because each $CuCl_2$ molecule breaks down into one copper atom (Cu) and one chlorine molecule ($CuCl_2$). However, what you should have noticed, is that the number of atoms of each element stays the same, as does the total mass of the atoms. This will be discussed in more detail in a later section.
2. *Energy changes*
The energy changes that take place during a chemical reaction are much greater than those that take place during a physical change in matter. During a chemical reaction, energy is used up in order to break bonds, and then energy is released when the new product is formed. This will be discussed in more detail in "Energy changes in chemical reactions" (Section 7.4: Energy changes in chemical reactions).
3. *Reversibility*
Chemical changes are far more difficult to reverse than physical changes.

We will consider two types of chemical reactions: **decomposition reactions** and **synthesis reactions**.

1.56.1 Decomposition reactions

A **decomposition reaction** occurs when a chemical compound is broken down into elements or smaller compounds. The generalised equation for a decomposition reaction is:

$AB \rightarrow A + B$

One example of such a reaction is the decomposition of mercury (II) oxide (Figure 8.3) to form mercury and oxygen according to the following equation:

$$2HgO \rightarrow 2Hg + O_2 \tag{8.2}$$

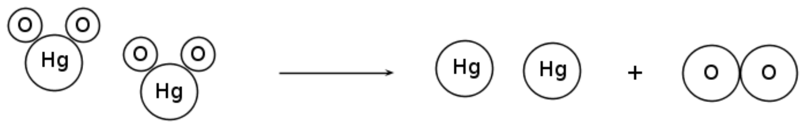


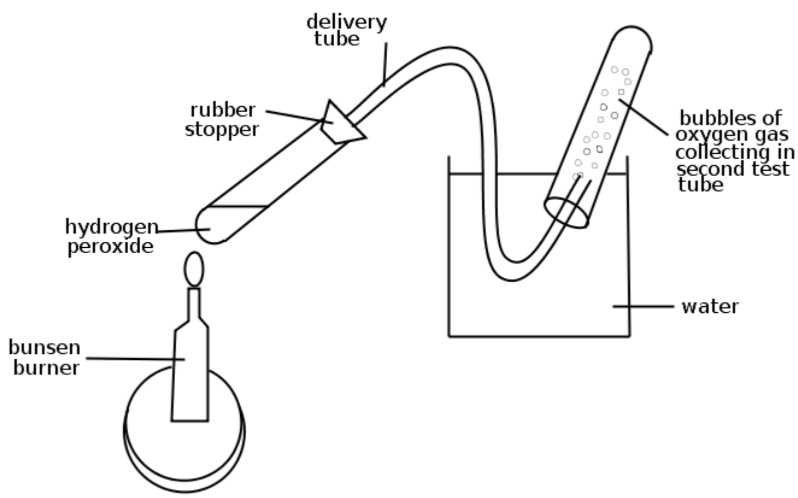
Figure 8.3: The decomposition of HgO to form Hg and O_2

The decomposition of hydrogen peroxide is another example.

Experiment : The decomposition of hydrogen peroxide

Aim:
To observe the decomposition of hydrogen peroxide when it is heated.

Apparatus:
Dilute hydrogen peroxide (about 3%); manganese dioxide; test tubes; a water bowl; stopper and delivery tube



Method:

1. Put a small amount (about 5 ml) of hydrogen peroxide in a test tube.
2. Set up the apparatus as shown in Figure 8.4
3. Very carefully add a small amount (about 0,5 g) of manganese dioxide to the test tube containing hydrogen peroxide.

Results:
You should observe a gas bubbling up into the second test tube. This reaction happens quite rapidly.

Conclusions:
When hydrogen peroxide is added to manganese dioxide it decomposes to form oxygen and water. The chemical decomposition reaction that takes place can be written as follows:

$$2H_2O_2 \rightarrow 2H_2O + O_2 \tag{8.3}$$

Note that the manganese dioxide is a catalyst and is not shown in the reaction. (A catalyst helps speed up a chemical reaction.)

NOTE: The previous experiment used the downward displacement of water to collect a gas. This is a very common way to collect a gas in chemistry. The oxygen that is evolved in this reaction moves along the delivery tube and then collects in the top of the test tube. It does this because it is lighter than water and so displaces the water downwards. If you use a test tube with an outlet attached, you could collect the oxygen into jars and store it for use in other experiments.

The above experiment can be very vigourous and produce a lot of oxygen very rapidly. For this reason you use dilute hydrogen peroxide and only a small amount of manganese dioxide.

1.56.2 Synthesis reactions

During a **synthesis reaction**, a new product is formed from elements or smaller compounds. The generalised equation for a synthesis reaction is as follows:

$$A + B \rightarrow AB \tag{8.4}$$

One example of a synthesis reaction is the burning of magnesium in oxygen to form magnesium oxide(Figure 8.5). The equation for the reaction is:

$$2Mg + O_2 \rightarrow 2MgO \tag{8.5}$$

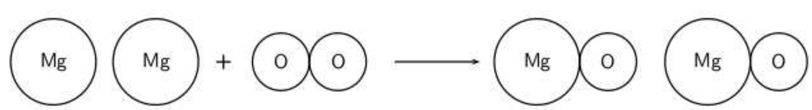
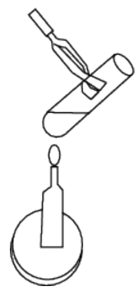


Figure 8.5: The synthesis of magnesium oxide (MgO) from magnesium and oxygen

Experiment: Chemical reactions involving iron and sulphur

Aim:
To demonstrate the synthesis of iron sulphide from iron and sulphur.

Apparatus:
5,6 g iron filings and 3,2 g powdered sulphur; porcelain dish; test tube; Bunsen burner



Method:

1. Measure the quantity of iron and sulphur that you need and mix them in a porcelain dish.
2. Take some of this mixture and place it in the test tube. The test tube should be about 1/3 full.
3. This reaction should ideally take place in a fume cupboard. Heat the test tube containing the mixture over the Bunsen burner. Increase the heat if no reaction takes place. Once the reaction begins, you will need to remove the test tube from the flame. Record your observations.
4. Wait for the product to cool before breaking the test tube with a hammer. Make sure that the test tube is rolled in paper before you do this, otherwise the glass will shatter everywhere and you may be hurt.
5. What does the product look like? Does it look anything like the original reactants? Does it have any of the properties of the reactants (e.g. the magnetism of iron)?

Results:

1. After you removed the test tube from the flame, the mixture glowed a bright red colour. The reaction is exothermic and *produces energy*.
2. The product, iron sulphide, is a dark colour and does not share any of the properties of the original reactants. It is an entirely new product.

Conclusions:
A synthesis reaction has taken place. The equation for the reaction is:

$$\mathrm{Fe} + \mathrm{S} \rightarrow \mathrm{FeS} \quad (8.6)$$

Investigation : Physical or chemical change?

Apparatus: Bunsen burner, 4 test tubes, a test tube rack and a test tube holder, small spatula, pipette, magnet, a birthday candle, $NaCl$ (table salt), $0,1\ M\ AgNO_3$, $6\ M\ HCl$, magnesium ribbon, iron filings, sulphur.

WARNING: $AgNO_3$ stains the skin. Be careful when working with it or use gloves.

Method:

1. Place a small amount of wax from a birthday candle into a test tube and heat it over the bunsen burner until it melts. Leave it to cool.
2. Add a small spatula of $NaCl$ to 5 ml water in a test tube and shake. Then use the pipette to add 10 drops of $AgNO_3$ to the sodium chloride solution. NOTE: Please be careful $AgNO_3$ causes bad stains!!
3. Take a 5 cm piece of magnesium ribbon and tear it into 1 cm pieces. Place two of these pieces into a test tube and add a few drops of $6\ M\ HCl$. NOTE: Be very careful when you handle this acid because it can cause major burns.
4. Take about 0,5 g iron filings and 0,5 g sulphur. Test each substance with a magnet. Mix the two samples in a test tube and run a magnet alongside the outside of the test tube.
5. Now heat the test tube that contains the iron and sulphur. What changes do you see? What happens now, if you run a magnet along the outside of the test tube?
6. In each of the above cases, record your observations.

Questions: Decide whether each of the following changes are physical or chemical and give a reason for your answer in each case. Record your answers in the table below:

Description	Physical or chemical change	Reason
melting candle wax		
dissolving $NaCl$		
mixing $NaCl$ with $AgNO_3$		
tearing magnesium ribbon		
adding HCl to magnesium ribbon		
mixing iron and sulphur		
heating iron and sulphur		

Table 8.1

1.57 Energy changes in chemical reactions

(section shortcode: C10055)

All reactions involve some change in energy. During a *physical* change in matter, such as the evaporation of liquid water to water vapour, the energy of the water molecules increases. However, the change in energy is much smaller than in chemical reactions.

When a chemical reaction occurs, some bonds will *break*, while new bonds may *form*. Energy changes in chemical reactions result from the breaking and forming of bonds. For bonds to *break*, energy must be *absorbed*. When new bonds *form*, energy will be *released* because the new product has a lower energy than the 'in between' stage of the reaction when the bonds in the reactants have just been broken.

In some reactions, the energy that must be *absorbed* to break the bonds in the reactants is less than the total energy that is *released* when new bonds are formed. This means that in the overall reaction, energy is *released*. This type of reaction is known as an **exothermic** reaction. In other reactions, the energy that must be *absorbed* to break the bonds in the reactants is more than the total energy that is *released* when new bonds are formed. This means that in the overall reaction, energy must be *absorbed* from the surroundings. This type of reaction is known as an **endothermic** reaction. Most decomposition reactions are endothermic and heating is needed for the reaction to occur. Most synthesis reactions are exothermic, meaning that energy is given off in the form of heat or light.

More simply, we can describe the energy changes that take place during a chemical reaction as:

Total energy absorbed to break bonds - Total energy released when new bonds form

So, for example, in the reaction...

$2\mathrm{Mg} + \mathrm{O}_2 \rightarrow 2\mathrm{MgO}$

Energy is needed to break the $\mathrm{O}-\mathrm{O}$ bonds in the oxygen molecule so that new $\mathrm{Mg}-\mathrm{O}$ bonds can be formed, and energy is released when the product (MgO) forms.

Despite all the energy changes that seem to take place during reactions, it is important to remember that energy cannot be created or destroyed. Energy that enters a system will have come from the surrounding environment and energy that leaves a system will again become part of that environment. This is known as the **conservation of energy** principle.

Definition: Conservation of energy principle
Energy cannot be created or destroyed. It can only be changed from one form to another.

Chemical reactions may produce some very visible and often violent changes. An explosion, for example, is a sudden increase in volume and release of energy when high temperatures are generated and gases are released. For example, $\mathrm{NH_4NO_3}$ can be heated to generate nitrous oxide. Under these conditions, it is highly sensitive and can detonate easily in an explosive exothermic reaction.

1.58 Conservation of atoms and mass in reactions

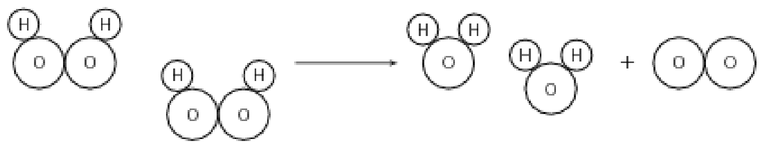
(section shortcode: C10056)

The total mass of all the substances taking part in a chemical reaction is conserved during a chemical reaction. This is known as the **law of conservation of mass**. The total number of **atoms** of each element also remains the same during a reaction, although these may be arranged differently in the products.

We will use two of our earlier examples of chemical reactions to demonstrate this:

1. The decomposition of hydrogen peroxide into water and oxygen

$2H_2O_2 \rightarrow 2H_2O + O_2$



Left hand side of the equation

$\text{Total atomic mass} = (4 \times 1) + (4 \times 16) = 68\,\text{u}$

$\text{Number of atoms of each element} = (4 \times \text{H}) + (4 \times \text{O})$

Right hand side of the equation

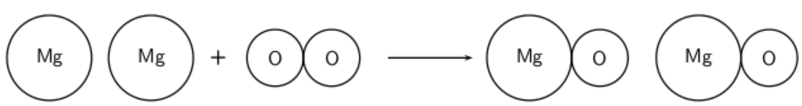
$\text{Total atomic mass} = (4 \times 1) + (4 \times 16) = 68\,\text{u}$

$\text{Number of atoms of each element} = (4 \times \text{H}) + (4 \times \text{O})$

Both the atomic mass and the number of atoms of each element are conserved in the reaction.

2. The synthesis of magnesium and oxygen to form magnesium oxide

$$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \tag{8.7}$$



Left hand side of the equation

$\text{Total atomic mass} = (2 \times 24,3) + (2 \times 16) = 80,6\,\text{u}$

$\text{Number of atoms of each element} = (2 \times \text{Mg}) + (2 \times \text{O})$

Right hand side of the equation

$\text{Total atomic mass} = (2 \times 24,3) + (2 \times 16) = 80,6\,\text{u}$

$\text{Number of atoms of each element} = (2 \times \text{Mg}) + (2 \times \text{O})$

Both the atomic mass and the number of atoms of each element are conserved in the reaction.

1.58.1 Activity : The conservation of atoms in chemical reactions

Materials:

1. Coloured marbles or small balls to represent atoms. Each colour will represent a different element.
2. Prestik

Method:

1. Choose a reaction from any that have been used in this chapter or any other *balanced* chemical reaction that you can think of. To help to explain this activity, we will use the decomposition reaction of calcium carbonate to produce carbon dioxide and calcium oxide. $CaCO_3 \rightarrow CO_2 + CaO$
2. Stick marbles together to represent the reactants and put these on one side of your table. In this example you may for example join one red marble (calcium), one green marble (carbon) and three yellow marbles (oxygen) together to form the molecule calcium carbonate ($CaCO_3$).
3. Leaving your reactants on the table, use marbles to make the product molecules and place these on the other side of the table.
4. Now count the number of atoms on each side of the table. What do you notice?
5. Observe whether there is any difference between the molecules in the reactants and the molecules in the products.

Discussion You should have noticed that the number of atoms in the reactants is the same as the number of atoms in the product. The number of atoms is conserved during the reaction. However, you will also see that the molecules in the reactants and products is not the same. The *arrangement of atoms* is not conserved during the reaction.

1.58.2 Experiment: Conservation of matter

Aim: To prove the law of conservation of matter experimentally.

Materials: Test tubes; glass beaker; lead (II) nitrate; sodium iodide; hydrochloric acid; bromothymol blue; Cal-C-Vita tablet, plastic bag; rubber band; mass meter

Method: Reaction 1

1. Carefully weigh out 5 g of lead (II) nitrate.
2. Dissolve the lead nitrate in 100 ml of water.
3. Weigh the lead nitrate solution.
4. Weigh out 4,5 g of sodium iodide and dissolve this in the lead (II) nitrate solution.
5. Weigh the beaker containing the lead nitrate and sodium iodide mixture.

Reaction 2

1. Measure out 20 ml of sodium hydroxide.
2. Add a few drops of bromothymol blue to the sodium hydroxide.
3. Weigh the sodium hydroxide.
4. Weigh 5 ml of hydrochloric acid.
5. Add 5 ml of hydrochloric acid to the sodium hydroxide. Repeat this step until you observe a colour change (this should occur around 20 ml).
6. Weigh the final solution.

Reaction 3

1. Measure out 100 ml of water into a beaker.
2. Weigh the beaker with water in it.
3. Place the Cal-C-Vita tablet into the plastic bag.
4. Weigh the Cal-C-Vita tablet and the plastic bag.
5. Place the plastic bag over the beaker, being careful to not let the tablet fall into the water
6. Seal the bag around the beaker using the rubber band. Drop the tablet into the water.
7. Observe what happens.
8. Weigh the bag and beaker containing the solution.

Results: Fill in the following table for reactants (starting materials) and products (ending materials) masses. For the second reaction, you will simply take the mass of 5 ml of hydrochloric acid and multiply it by how many amounts you put in, for example, if you put 4 amounts in, then you would have 20 ml and 4 times the mass of 5 ml.

	Reaction 1	Reaction 2	Reaction 3
Reactants			
Products			

Table 8.2

Add the masses for the reactants for each reaction. Do the same for the products. For each reaction compare the mass of the reactants to the mass of the products. What do you notice? Is the mass conserved?

In the experiment above you should have found that the mass at the start of the reaction is the same as the mass at the end of the reaction. You may have found that these masses differed slightly, but this is due to errors in measurements and in performing experiments (all scientists make some errors in performing experiments).

1.59 Law of constant composition

www (section shortcode: C10057)

In any given chemical compound, the elements always combine in the same proportion with each other. This is the **law of constant proportion**.

The **law of constant composition** says that, in any particular chemical compound, all samples of that compound will be made up of the same elements in the same proportion or ratio. For example, any water molecule is always made up of two hydrogen atoms and one oxygen atom in a 2:1 ratio. If we look at the relative masses of oxygen and hydrogen in a water molecule, we see that 94% of the mass of a water molecule is accounted for by oxygen and the remaining 6% is the mass of hydrogen. This mass proportion will be the same for any water molecule.

This does not mean that hydrogen and oxygen always combine in a 2:1 ratio to form H_2O. Multiple proportions are possible. For example, hydrogen and oxygen may combine in different proportions to form H_2O_2 rather than H_2O. In H_2O_2, the H:O ratio is 1:1 and the mass ratio of hydrogen to oxygen is 1:16. This will be the same for any molecule of hydrogen peroxide.

1.60 Volume relationships in gases

(section shortcode: C10058)

In a chemical reaction between gases, the relative volumes of the gases in the reaction are present in a ratio of small whole numbers if all the gases are at the same temperature and pressure. This relationship is also known as **Gay-Lussac's Law**.

For example, in the reaction between hydrogen and oxygen to produce water, two volumes of H_2 react with 1 volume of O_2 to produce 2 volumes of H_2O.

$2H_2 + O_2 \rightarrow 2H_2O$

In the reaction to produce ammonia, one volume of nitrogen gas reacts with three volumes of hydrogen gas to produce two volumes of ammonia gas.

$N_2 + 3H_2 \rightarrow 2NH_3$

This relationship will also be true for all other chemical reactions.

1.61 Summary

(section shortcode: C10059)

The following video provides a summary of the concepts covered in this chapter.

Physical and chemical change (Video: P10060)

1. Matter does not stay the same. It may undergo physical or chemical changes.
2. A **physical change** means that the form of matter may change, but not its identity. For example, when water evaporates, the energy and the arrangement of water molecules will change, but not the structure of the water molecules themselves.
3. During a physical change, the **arrangement of particles** may change but the mass, number of atoms and number of molecules will stay the same.
4. Physical changes involve small changes in **energy** and are easily reversible.
5. A chemical change occurs when one or more substances change into other materials. A chemical reaction involves the formation of new substances with **different properties**. For example, magnesium and oxygen react to form magnesium oxide (MgO)
6. A chemical change may involve a **decomposition** or **synthesis** reaction. During chemical change, the mass and number of atoms is conserved, but the number of molecules is not always the same.
7. Chemical reactions involve larger changes in energy. During a reaction, energy is needed to break bonds in the reactants and energy is released when new products form. If the energy released is greater than the energy absorbed, then the reaction is exothermic. If the energy released is less than the energy absorbed, then the reaction is endothermic. Chemical reactions are not easily reversible.
8. Decomposition reactions are usually **endothermic** and synthesis reactions are usually **exothermic**.
9. The **law of conservation of mass** states that the total mass of all the substances taking part in a chemical reaction is conserved and the number of atoms of each element in the reaction does not change when a new product is formed.

10. The **conservation of energy principle** states that energy cannot be created or destroyed, it can only change from one form to another.
11. The **law of constant composition** states that in any particular compound, all samples of that compound will be made up of the same elements in the same proportion or ratio.
12. **Gay-Lussac's Law** states that in a chemical reaction between gases, the relative volumes of the gases in the reaction are present in a ratio of small whole numbers if all the gases are at the same temperature and pressure.

1.61.1 End of chapter exercises

1. For each of the following definitions give one word or term:
   a. A change that can be seen or felt, where the particles involved are not broken up in any way
   b. The formation of new substances in a chemical reaction
   c. A reaction where a new product is formed from elements or smaller compounds
2. State the conservation of energy principle.
3. Explain how a chemical change differs from a physical change.
4. Complete the following table by saying whether each of the descriptions is an example of a physical or chemical change:

Description	Physical or chemical
hot and cold water mix together	
milk turns sour	
a car starts to rust	
food digests in the stomach	
alcohol disappears when it is placed on your skin	
warming food in a microwave	
separating sand and gravel	
fireworks exploding	

Table 8.3

5. For each of the following reactions, say whether it is an example of a synthesis or decomposition reaction:
   a. $(NH_4)_2 CO_3 \rightarrow NH_3 + CO_2 + H_2O$
   b. $N_2 (g) + 3H_2 (g) \rightarrow 2NH_3$
   c. $CaCO_3 (s) \rightarrow CaO + CO_2$
6. For the following equation: $CaCO_3 (s) \rightarrow CaO + CO_2$ show that the 'law of conservation of mass' applies.

Find the answers with the shortcodes:

(1.) l2z (2.) l2u (3.) l2J (4.) l3q (5.) l3l (6.) l3i

Representing Chemical Change

1.62 Introduction

(section shortcode: C10061)

As we have already mentioned, a number of changes can occur when elements react with one another. These changes may either be *physical* or *chemical*. One way of representing these changes is through **balanced chemical equations**. A chemical equation describes a chemical reaction by using symbols for the elements involved. For example, if we look at the reaction between iron (Fe) and sulphur (S) to form iron sulphide (FeS), we could represent these changes either in words or using chemical symbols:

$\mathrm{iron} + \mathrm{sulphur} \rightarrow \mathrm{iron\ sulphide}$

or

$\mathrm{Fe} + \mathrm{S} \rightarrow \mathrm{FeS}$

Another example would be:

$\mathrm{ammonia} + \mathrm{oxygen} \rightarrow \mathrm{nitric\ oxide} + \mathrm{water}$

or

$4\mathrm{NH}_3 + 5\mathrm{O}_2 \rightarrow 4\mathrm{NO} + 6\mathrm{H}_2\mathrm{O}$

Compounds on the left of the arrow are called the **reactants** and these are needed for the reaction to take place. In this equation, the reactants are ammonia and oxygen. The compounds on the right are called the **products** and these are what is formed from the reaction.

In order to be able to write a balanced chemical equation, there are a number of important things that need to be done:

1. Know the chemical symbols for the elements involved in the reaction
2. Be able to write the chemical formulae for different reactants and products
3. Balance chemical equations by understanding the laws that govern chemical change
4. Know the state symbols for the equation

We will look at each of these steps separately in the next sections.

1.63 Chemical symbols

(section shortcode: C10062)

It is very important to know the chemical symbols for common elements in the Periodic Table, so that you are able to write chemical equations and to recognise different compounds.

1.63.1 Revising common chemical symbols

- Write down the chemical symbols and names of all the elements that you know.
- Compare your list with another learner and add any symbols and names that you don't have.
- Spend some time, either in class or at home, learning the symbols for at least the first twenty elements in the periodic table. You should also learn the symbols for other common elements that are not in the first twenty.
- Write a short test for someone else in the class and then exchange tests with them so that you each have the chance to answer one.

1.64 Writing chemical formulae

(section shortcode: C10063)

A **chemical formula** is a concise way of giving information about the atoms that make up a particular chemical compound. A chemical formula shows each element by its symbol and also shows how many atoms of each element are found in that compound. The number of atoms (if greater than one) is shown as a subscript.

Examples: CH_4 (methane)

Number of atoms: $(1 \times \mathrm{carbon}) + (4 \times \mathrm{hydrogen}) = 5$ atoms in one methane molecule

H_2SO_4 (sulphuric acid)

Number of atoms: $(2 \times \mathrm{hydrogen}) + (1 \times \mathrm{sulphur}) + (4 \times \mathrm{oxygen}) = 7$ atoms in one molecule of sulphuric acid

A chemical formula may also give information about how the atoms are arranged in a molecule if it is written in a particular way. A molecule of ethane, for example, has the chemical formula C_2H_6. This formula tells us how many atoms of each element are in the molecule, but doesn't tell us anything about how these atoms are arranged. In fact, each carbon atom in the ethane molecule is bonded to three hydrogen atoms. Another way of writing the formula for ethane is CH_3CH_3. The number of atoms of each element has not changed, but this formula gives us more information about how the atoms are arranged in relation to each other.

The slightly tricky part of writing chemical formulae comes when you have to work out the ratio in which the elements combine. For example, you may know that sodium (Na) and chlorine (Cl) react to form sodium chloride, but how do you know that in each molecule of sodium chloride there is only *one* atom of sodium for every *one* atom of chlorine? It all comes down to the **valency** of an atom or group of atoms. Valency is the number of bonds that an element can form with another element. Working out the chemical formulae of chemical compounds using their valency, will be covered in Grade 11. For now, we will use formulae that you already know.

1.65 Balancing chemical equations

(section shortcode: C10064)

1.65.1 The law of conservation of mass

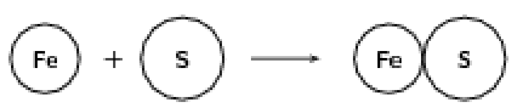
In order to balance a chemical equation, it is important to understand the law of conservation of mass.

Definition: The law of conservation of mass
The mass of a closed system of substances will remain constant, regardless of the processes acting inside the system. Matter can change form, but cannot be created or destroyed. For any chemical process in a closed system, the mass of the reactants must equal the mass of the products.

In a chemical equation then, the **mass** of the reactants must be equal to the mass of the products. In order to make sure that this is the case, the number of **atoms** of each element in the reactants must be equal to the number of atoms of those same elements in the products. Some examples are shown below:

Example 1:

$$\mathrm{Fe} + \mathrm{S} \rightarrow \mathrm{FeS} \qquad (9.1)$$



Reactants

$\mathrm{Atomic\ mass\ of\ reactants} = 55,8\,u + 32,1\,u = 87,9\,u$

Number of atoms of each element in the reactants: $(1 \times Fe)$ and $(1 \times \mathrm{S})$

Products

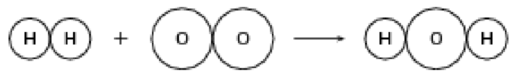
$\mathrm{Atomic\ mass\ of\ products} = 55,8\,u + 32,1\,u = 87,9\,u$

Number of atoms of each element in the products: $(1 \times Fe)$ and $(1 \times \mathrm{S})$

Since the number of atoms of each element is the same in the reactants and in the products, we say that the equation is **balanced**.

Example 2:

$$H_2 + O_2 \rightarrow H_2O \tag{9.2}$$



Reactants

Atomic mass of reactants $= (1+1)+(16+16) = 34\,u$

Number of atoms of each element in the reactants: $(2 \times \mathrm{H})$ and $(2 \times \mathrm{O})$

Product

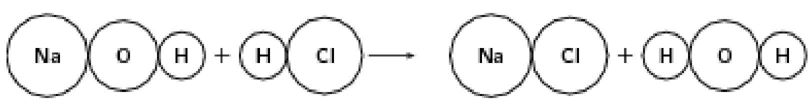
Atomic mass of product $= (1+1+16) = 18\,u$

Number of atoms of each element in the product: $(2 \times \mathrm{H})$ and $(1 \times \mathrm{O})$

Since the total atomic mass of the reactants and the products is not the same and since there are more oxygen atoms in the reactants than there are in the product, the equation is **not balanced**.

Example 3:

$$NaOH + HCl \rightarrow NaCl + H_2O \tag{9.3}$$



Reactants

Atomic mass of reactants $= (23+6+1)+(1+35,4) = 76,4\,u$

Number of atoms of each element in the reactants: $(1 \times \mathrm{Na}) + (1 \times \mathrm{O}) + (2 \times \mathrm{H}) + (1 \times \mathrm{Cl})$

Products

Atomic mass of products $= (23+35,4)+(1+1+16) = 76,4\,u$

Number of atoms of each element in the products: $(1 \times \mathrm{Na}) + (1 \times \mathrm{O}) + (2 \times \mathrm{H}) + (1 \times \mathrm{Cl})$

Since the number of atoms of each element is the same in the reactants and in the products, we say that the equation is **balanced**.

We now need to find a way to balance those equations that are not balanced so that the number of atoms of each element in the reactants is the same as that for the products. This can be done by changing the **coefficients** of the molecules until the atoms on each side of the arrow are balanced. You will see later that these coefficients tell us something about the **mole ratio** in which substances react. They also tell us about the volume relationship between gases in the reactants and products.

TIP: Coefficients : Remember that if you put a number in front of a molecule, that number applies to the *whole* molecule. For example, if you write $2H_2O$, this means that there are 2 molecules of water. In other words, there are 4 hydrogen atoms and 2 oxygen atoms. If we write $3HCl$, this means that there are 3 molecules of HCl. In other words there are 3 hydrogen atoms and 3 chlorine atoms in total. In the first example, 2 is the coefficient and in the second example, 3 is the coefficient.

1.65.2 Activity: Balancing chemical equations

You will need: coloured balls (or marbles), prestik, a sheet of paper and coloured pens.

We will try to balance the following equation:

$$\mathrm{Al} + \mathrm{O_2} \rightarrow \mathrm{Al_2O_3} \tag{9.4}$$

Take 1 ball of one colour. This represents a molecule of Al. Take two balls of another colour and stick them together. This represents a molecule of O_2. Place these molecules on your left. Now take two balls of one colour and three balls of another colour to form a molecule of Al_2O_3. Place these molecules on your right. On a piece of paper draw coloured circles to represent the balls. Draw a line down the center of the paper to represent the molecules on the left and on the right.

Count the number of balls on the left and the number on the right. Do you have the same number of each colour on both sides? If not the equation is not balanced. How many balls will you have to add to each side to make the number of balls the same? How would you add these balls?

You should find that you need 4 balls of one colour for Al and 3 pairs of balls of another colour (i.e. 6 balls in total) for O_2 on the left side. On the right side you should find that you need 2 clusters of balls for Al_2O_3. We say that the balanced equation is:

$$4\mathrm{Al} + 3\mathrm{O_2} \rightarrow 2\mathrm{Al_2O_3} \tag{9.5}$$

Repeat this process for the following reactions:

- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
- $2H_2 + O_2 \rightarrow 2H_2O$
- $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

1.65.3 Steps to balance a chemical equation

When balancing a chemical equation, there are a number of steps that need to be followed.

- STEP 1: Identify the reactants and the products in the reaction and write their chemical formulae.
- STEP 2: Write the equation by putting the reactants on the left of the arrow and the products on the right.
- STEP 3: Count the number of atoms of each element in the reactants and the number of atoms of each element in the products.
- STEP 4: If the equation is not balanced, change the coefficients of the molecules until the number of atoms of each element on either side of the equation balance.
- STEP 5: Check that the atoms are in fact balanced.
- STEP 6 (we will look at this a little later): Add any extra details to the equation e.g. phase.

Exercise 9.1: Balancing chemical equations 1 Balance the following equation:

$$Mg + HCl \rightarrow MgCl_2 + H_2 \quad (9.6)$$

Solution to Exercise

Step 1. Reactants: $Mg = 1\,atom; H = 1\,atom$ and $Cl = 1\,atom$
Products: $Mg = 1\,atom; H = 2\,atoms$ and $Cl = 2\,atoms$

Step 2. The equation is not balanced since there are 2 chlorine atoms in the product and only 1 in the reactants. If we add a coefficient of 2 to the HCl to increase the number of H and Cl atoms in the reactants, the equation will look like this:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \quad (9.7)$$

Step 3. If we count the atoms on each side of the equation, we find the following:
Reactants: $Mg = 1\,atom; H = 2\,atom$ and $Cl = 2\,atom$
Products: $Mg = 1\,atom; H = 2\,atom$ and $Cl = 2\,atom$
The equation is balanced. The final equation is:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \quad (9.8)$$

Exercise 9.2: Balancing chemical equations 2 Balance the following equation:

$$CH_4 + O_2 \rightarrow CO_2 + H_2O \quad (9.9)$$

Solution to Exercise

Step 1. Reactants: $C = 1; H = 4$ and $O = 2$
Products: $C = 1; H = 2$ and $O = 3$

Step 2. If we add a coefficient of 2 to H_2O, then the number of hydrogen atoms in the reactants will be 4, which is the same as for the reactants. The equation will be:

$$CH_4 + O_2 \rightarrow CO_2 + 2H_2O \quad (9.10)$$

Step 3. Reactants: $C = 1; H = 4$ and $O = 2$
Products: $C = 1; H = 4$ and $O = 4$

You will see that, although the number of *hydrogen* atoms now balances, there are more oxygen atoms in the products. You now need to repeat the previous step. If we put a coefficient of 2 in front of O_2, then we will increase the number of oxygen atoms in the reactants by 2. The new equation is:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \quad (9.11)$$

When we check the number of atoms again, we find that the number of atoms of each element in the reactants is the same as the number in the products. The equation is now balanced.

Exercise 9.3: Balancing chemical equations 3 Nitrogen gas reacts with hydrogen gas to form ammonia. Write a balanced chemical equation for this reaction.

Solution to Exercise

Step 1. The reactants are nitrogen (N_2) and hydrogen (H_2) and the product is ammonia (NH_3).

Step 2. The equation is as follows:

$$N_2 + H_2 \rightarrow NH_3 \quad (9.12)$$

Step 3. Reactants: $N = 2$ and $H = 2$
Products: $N = 1$ and $H = 3$

Step 4. In order to balance the number of nitrogen atoms, we could rewrite the equation as:

$$N_2 + H_2 \rightarrow 2NH_3 \quad (9.13)$$

Step 5. In the above equation, the nitrogen atoms now balance, but the hydrogen atoms don't (there are 2 hydrogen atoms in the reactants and 6 in the product). If we put a coefficient of 3 in front of the hydrogen (H_2), then the hydrogen atoms and the nitrogen atoms balance. The final equation is:

$$N_2 + 3H_2 \rightarrow 2NH_3 \quad (9.14)$$

Exercise 9.4: Balancing chemical equations 4 In our bodies, sugar ($C_6H_{12}O_6$) reacts with the oxygen we breathe in to produce carbon dioxide, water and energy. Write the balanced equation for this reaction.

Solution to Exercise

Step 1. Reactants: sugar ($C_6H_{12}O_6$) and oxygen (O_2)
Products: carbon dioxide (CO_2) and water (H_2O)

Step 2.

$$C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O \tag{9.15}$$

Step 3. Reactants: $C = 6$; $H = 12$ and $O = 8$
Products: $C = 1$; $H = 2$ and $O = 3$

Step 4. It is easier to start with carbon as it only appears once on each side.
If we add a 6 in front of CO_2, the equation looks like this:

$$C_6H_{12}O_6 + O_2 \rightarrow 6CO_2 + H_2O \tag{9.16}$$

Reactants: $C = 6$; $H = 12$ and $O = 8$
Products: $C = 6$; $H = 2$ and $O = 13$

Step 5. Let's try to get the number of hydrogens the same this time.

$$C_6H_{12}O_6 + O_2 \rightarrow 6CO_2 + 6H_2O \tag{9.17}$$

Reactants: $C = 6$; $H = 12$ and $O = 8$
Products: $C = 6$; $H = 12$ and $O = 18$

Step 6.

$$C_6H_{12}O_6 + 12O_2 \rightarrow 6CO_2 + 6H_2O \tag{9.18}$$

Reactants: $C = 6$; $H = 12$ and $O = 18$
Products: $C = 6$; $H = 12$ and $O = 18$

This simulation allows you to practice balancing simple equations.

(Simulation: lbD)

Balancing simple chemical equations

Balance the following equations:

1. Hydrogen fuel cells are extremely important in the development of alternative energy sources. Many of these cells work by reacting hydrogen and oxygen gases together to form water, a reaction which also produces electricity. Balance the following equation: $H_2\,(g) + O_2\,(g) \rightarrow H_2O\,(l)$

2. The synthesis of ammonia (NH_3), made famous by the German chemist Fritz Haber in the early 20th century, is one of the most important reactions in the chemical industry. Balance the following equation used to produce ammonia: $N_2\,(g) + H_2\,(g) \rightarrow NH_3\,(g)$
3. $Mg + P_4 \rightarrow Mg_3P_2$
4. $Ca + H_2O \rightarrow Ca\,(OH)_2 + H_2$
5. $CuCO_3 + H_2SO_4 \rightarrow CuSO_4 + H_2O + CO_2$
6. $CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + NaCl$
7. $C_{12}H_{22}O_{11} + O_2 \rightarrow H_2O + CO_2$
8. Barium chloride reacts with sulphuric acid to produce barium sulphate and hydrochloric acid.
9. Ethane (C_2H_6) reacts with oxygen to form carbon dioxide and steam.
10. Ammonium carbonate is often used as a smelling salt. Balance the following reaction for the decomposition of ammonium carbonate: $(NH_4)_2CO_3\,(s) \rightarrow NH_3\,(aq) + CO_2\,(g) + H_2O\,(l)$

Find the answers with the shortcodes:

(1.) IOg (2.) IO4 (3.) IO2 (4.) IOT (5.) IOb (6.) IOj
(7.) IOD (8.) IOW (9.) IOZ (10.) IOB

1.66 State symbols and other information

(section shortcode: C10065)

The state (phase) of the compounds can be expressed in the chemical equation. This is done by placing the correct label on the right hand side of the formula. There are only four labels that can be used:

1. (g) for gaseous compounds
2. (l) for liquids
3. (s) for solid compounds
4. (aq) for an aqueous (water) solution

Occasionally, a catalyst is added to the reaction. A catalyst is a substance that speeds up the reaction without undergoing any change to itself. In a chemical equation, this is shown by using the symbol of the catalyst above the arrow in the equation.

To show that heat is needed for a reaction, a Greek delta (Δ) is placed above the arrow in the same way as the catalyst.

TIP: You may remember from Physical and chemical change (Chapter 7) that energy cannot be created or destroyed during a chemical reaction but it may change form. In an exothermic reaction, ΔH is less than zero and in an endothermic reaction, ΔH is greater than zero. This value is often written at the end of a chemical equation.

Exercise 9.5: Balancing chemical equations 5 Solid zinc metal reacts with aqueous hydrochloric acid to form an aqueous solution of zinc chloride ($ZnCl_2$)and hydrogen gas. Write a balanced equation for this reaction.

Solution to Exercise

Step 1. The reactants are zinc (Zn) and hydrochloric acid (HCl). The products are zinc chloride ($ZnCl_2$) and hydrogen (H_2).

Step 2.

$$Zn + HCl \rightarrow ZnCl_2 + H_2 \tag{9.19}$$

Step 3. You will notice that the zinc atoms balance but the chlorine and hydrogen atoms don't. Since there are two chlorine atoms on the right and only one on the left, we will give HCl a coefficient of 2 so that there will be two chlorine atoms on each side of the equation.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2 \tag{9.20}$$

Step 4. When you look at the equation again, you will see that all the atoms are now balanced.

Step 5. In the initial description, you were told that zinc was a metal, hydrochloric acid and zinc chloride were in aqueous solutions and hydrogen was a gas.

$$Zn\,(s) + 2HCl\,(aq) \rightarrow ZnCl_2\,(aq) + H_2\,(g) \tag{9.21}$$

Exercise 9.6: Balancing chemical equations 5 (advanced) Balance the following equation:

$$(NH_4)\,_2SO_4 + NaOH \rightarrow NH_3 + H_2O + Na_2SO_4 \tag{9.22}$$

In this example, the first two steps are not necessary because the reactants and products have already been given.

Solution to Exercise

Step 1. With a complex equation, it is always best to start with atoms that appear only once on each side i.e. Na, N and S atoms. Since the S atoms already balance, we will start with Na and N atoms. There are two Na atoms on the right and one on the left. We will add a second Na atom by giving $NaOH$ a coefficient of two. There are two N atoms on the left and one on the right. To balance the N atoms, NH_3 will be given a coefficient of two. The equation now looks as follows:

$$(NH_4)_2SO_4 + 2NaOH \rightarrow 2NH_3 + H_2O + Na_2SO_4 \quad (9.23)$$

Step 2. N, Na and S atoms balance, but O and H atoms do not. There are six O atoms and ten H atoms on the left, and five O atoms and eight H atoms on the right. We need to add one O atom and two H atoms on the right to balance the equation. This is done by adding another H_2O molecule on the right hand side. We now need to check the equation again:

$$(NH_4)_2SO_4 + 2NaOH \rightarrow 2NH_3 + 2H_2O + Na_2SO_4 \quad (9.24)$$

The equation is now balanced.

The following video explains some of the concepts of balancing chemical equations.

Khan Academy video on balancing chemical equations www (Video: P10066)

1.66.1 Balancing more advanced chemical equations

Write balanced equations for each of the following reactions:

1. $Al_2O_3\,(s) + H_2SO_4\,(aq) \rightarrow Al_2\,(SO_4)_3\,(aq) + 3H_2O\,(l)$
2. $Mg\,(OH)_2\,(aq) + HNO_3\,(aq) \rightarrow Mg\,(NO_3)_2\,(aq) + 2H_2O\,(l)$
3. Lead (II) nitrate solution reacts with potassium iodide solution.
4. When heated, aluminium reacts with solid copper oxide to produce copper metal and aluminium oxide (Al_2O_3).
5. When calcium chloride solution is mixed with silver nitrate solution, a white precipitate (solid) of silver chloride appears. Calcium nitrate ($Ca\,(NO_3)_2$) is also produced in the solution.

www Find the answers with the shortcodes:

(1.) IOK

Balanced equations are very important in chemistry. It is only by working with the balanced equations that chemists can perform many different calculations that tell them what quantity of something reacts. In a later chapter we will learn how to work with some of these calculations. We can interpret balanced chemical equations in terms of the conservation of matter, the conservation of mass or the conservation of energy.

www (Presentation: P10067)

1.67 Summary

(section shortcode: C10068)

- A **chemical equation** uses symbols to describe a chemical reaction.
- In a chemical equation, **reactants** are written on the left hand side of the equation and the **products** on the right. The arrow is used to show the direction of the reaction.
- When representing chemical change, it is important to be able to write the **chemical formula** of a compound.
- In any chemical reaction, the **law of conservation of mass** applies. This means that the total atomic mass of the reactants must be the same as the total atomic mass of the products. This also means that the number of atoms of each element in the reactants must be the same as the number of atoms of each element in the product.
- If the number of atoms of each element in the reactants is the same as the number of atoms of each element in the product, then the equation is **balanced**.
- If the number of atoms of each element in the reactants is not the same as the number of atoms of each element in the product, then the equation is **not balanced**.
- In order to balance an equation, **coefficients** can be placed in front of the reactants and products until the number of atoms of each element is the same on both sides of the equation.

1.67.1 End of chapter exercises

1. Propane is a fuel that is commonly used as a heat source for engines and homes. Balance the following equation for the combustion of propane: $\mathrm{C_3H_8\,(l) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(l)}$
2. Aspartame, an artificial sweetener, has the formula $\mathrm{C_{14}H_{18}N_2O_2}$. Write the balanced equation for its combustion (reaction with $\mathrm{O_2}$) to form $\mathrm{CO_2}$ gas, liquid $\mathrm{H_2O}$, and $\mathrm{N_2}$ gas.
3. $Fe_2\,\mathrm{(SO_4)}_3 + \mathrm{K\,(SCN) \rightarrow K_3Fe\,(SCN)_6 + K_2SO_4}$
4. Chemical weapons were banned by the Geneva Protocol in 1925. According to this protocol, all chemicals that release suffocating and poisonous gases are not to be used as weapons. White phosphorus, a very reactive allotrope of phosphorus, was recently used during a military attack. Phosphorus burns vigorously in oxygen. Many people got severe burns and some died as a result. The equation for this spontaneous reaction is: $\mathrm{P_4\,(s) + O_2\,(g) \rightarrow P_2O_5\,(s)}$
   - **a.** Balance the chemical equation.
   - **b.** Prove that the law of conservation of mass is obeyed during this chemical reaction.
   - **c.** Name the product formed during this reaction.
   - **d.** Classify the reaction as endothermic or exothermic. Give a reason for your answer.
   - **e.** Classify the reaction as a synthesis or decomposition reaction. Give a reason for your answer.

   (DoE Exemplar Paper 2 2007)

5. Mixing bleach ($NaOCl$) and ammonia (two common household cleaners) is very dangerous. When these two substances are mixed they produce toxic chloaramine (NH_2Cl) fumes. Balance the following equations that occur when bleach and ammonia are mixed:
   a. $NaOCl\,(aq) + NH_3\,(aq) \rightarrow NaONH_3\,(aq) + Cl_2\,(g)$
   b. If there is more bleach than ammonia the following may occur: $NaOCl + NH_3 \rightarrow NaOH + NCl_3$
      Nitrogen trichloride (NCl_3) is highly explosive.
   c. If there is more ammonia than bleach the following may occur: $NH_3 + NaOCl \rightarrow NaOH + NH_2Cl$
      These two products then react with ammonia as follows:
      $NH_3 + NH_2Cl + NaOH \rightarrow N_2H_4 + NaCl + H_2O$
      One last reaction occurs to stabilise the hydrazine and chloramine: $NH_2Cl + N_2H_4 \rightarrow NH_4Cl + N_2$
      This reaction is highly exothermic and will explode.
6. Balance the following chemical equation: $N_2O_5 \rightarrow NO_2 + O_2$
7. Sulphur can be produced by the Claus process. This two-step process involves reacting hydrogen sulphide with oxygen and then reacting the sulphur dioxide that is produced with more hydrogen sulphide. The equations for these two reactions are:

$$\begin{aligned} H_2S + O_2 &\rightarrow SO_2 + H_2O \\ H_2S + SO_2 &\rightarrow S + H_2O \end{aligned} \tag{9.25}$$

   Balance these two equations.

Find the answers with the shortcodes:

(1.) IOk (2.) IO0 (3.) IO8 (4.) IO9 (5.) I2S (6.) I2h
(7.) ITq

Reactions in Aqueous Solutions

1.68 Introduction

(section shortcode: C10069)

Many reactions in chemistry and all biological reactions (reactions in living systems) take place in water. We say that these reactions take place in aqueous solution. Water has many unique properties and is plentiful on Earth. For these reasons reactions in aqueous solutions occur frequently. In this chapter we will look at some of these reactions in detail. Almost all the reactions that occur in aqueous solutions involve ions. We will look at three main types of reactions that occur in aqueous solutions, namely precipitation reactions, acid-base reactions and redox reactions. Before we can learn about the types of reactions, we need to first look at ions in aqueous solutions and electrical conductivity.

1.69 Ions in aqueous solution

(section shortcode: C10070)

Water is seldom pure. Because of the structure of the water molecule, substances can dissolve easily in it. This is very important because if water wasn't able to do this, life would not be able to survive. In rivers and the oceans for example, dissolved oxygen means that organisms (such as fish) are still able to respire (breathe). For plants, dissolved nutrients are also available. In the human body, water is able to carry dissolved substances from one part of the body to another.

Many of the substances that dissolve are *ionic* and when they dissolve they form ions in solution. We are going to look at how water is able to dissolve ionic compounds, how these ions maintain a balance in the human body, how they affect water hardness and how they cause acid rain.

1.69.1 Dissociation in water

Water is a **polar molecule** (Figure 10.1). This means that one part of the molecule has a slightly positive charge (positive pole) and the other part has a slightly negative charge (negative pole).

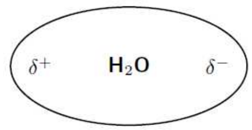


Figure 10.1: Water is a polar molecule

It is the polar nature of water that allows ionic compounds to dissolve in it. In the case of sodium chloride (NaCl) for example, the positive sodium ions (Na^+) will be attracted to the negative pole of the water molecule, while the negative chloride ions (Cl^-) will be attracted to the positive pole of the water molecule. In the process, the ionic bonds between the sodium and chloride ions are weakened and the water molecules are able to work their way between the individual ions, surrounding them and slowly dissolving the compound. This process is called **dissociation**. A simplified representation of this is shown in Figure 10.2. We say that dissolution of a substance has occurred when a substance dissociates or dissolves.

Definition: Dissociation
Dissociation in chemistry and biochemistry is a general process in which ionic compounds separate or split into smaller molecules or ions, usually in a reversible manner.

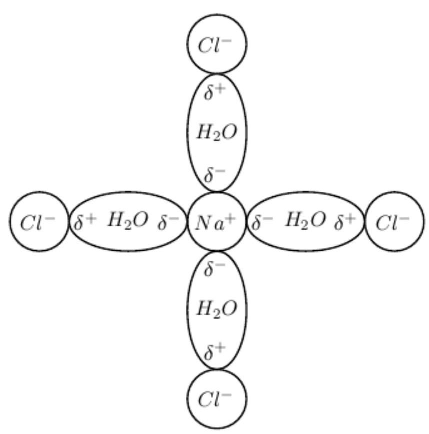


Figure 10.2: Sodium chloride dissolves in water

The dissolution of sodium chloride can be represented by the following equation:

$$\mathrm{NaCl}\,(\mathrm{s}) \rightarrow \mathrm{Na}^+(\mathrm{aq}) + \mathrm{Cl}^-(\mathrm{aq})$$

The symbols **s** (solid), **l** (liquid), **g** (gas) and **aq** (material is dissolved in water) are written after the chemical formula to show the state or phase of the material. The dissolution of potassium sulphate into potassium and sulphate ions is shown below as another example:

$K_2SO_4(s) \rightarrow 2K^+(aq) + SO_4^{2-}(aq)$

Remember that **molecular** substances (e.g. covalent compounds) may also dissolve, but most will not form ions. One example is sugar.

$C_6H_{12}O_6(s) \rightleftharpoons C_6H_{12}O_6(aq)$

There are exceptions to this and some molecular substances *will* form ions when they dissolve. Hydrogen chloride for example can ionise to form hydrogen and chloride ions.

$HCl\,(g) \rightarrow H^+(aq) + Cl^-(aq)$

NOTE: The ability of ionic compounds to dissolve in water is extremely important in the human body! The body is made up of *cells*, each of which is surrounded by a *membrane*. Dissolved ions are found inside and outside of body cells in different concentrations. Some of these ions are positive (e.g. Mg^{2+}) and some are negative (e.g. Cl^-). If there is a difference in the charge that is inside and outside the cell, then there is a *potential difference* across the cell membrane. This is called the **membrane potential** of the cell. The membrane potential acts like a battery and affects the movement of all charged substances across the membrane. Membrane potentials play a role in muscle functioning, digestion, excretion and in maintaining blood pH to name just a few. The movement of ions across the membrane can also be converted into an electric signal that can be transferred along *neurons* (nerve cells), which control body processes. If ionic substances were not able to dissociate in water, then none of these processes would be possible! It is also important to realise that our bodies can *lose* ions such as Na^+, K^+, Ca^{2+}, Mg^{2+}, and Cl^-, for example when we sweat during exercise. Sports drinks such as Lucozade and Powerade are designed to replace these lost ions so that the body's normal functioning is not affected.

Exercise 10.1: Dissociation in water Write a balanced equation to show how silver nitrate ($AgNO_3$) dissociates in water.

Solution to Exercise

Step 1. The cation is: Ag^+ and the anion is: NO_3^-

Step 2. Since we know both the anion and the cation that silver nitrate dissociates into we can write the following equation:

$$AgNO_3(s) \rightarrow Ag^+(aq) + NO_3^-(aq) \tag{10.1}$$

Ions in solution

1. For each of the following, say whether the substance is ionic or molecular.
   - **a.** potassium nitrate (KNO_3)
   - **b.** ethanol (C_2H_5OH)
   - **c.** sucrose (a type of sugar) ($C_{12}H_{22}O_{11}$)
   - **d.** sodium bromide ($NaBr$)
2. Write a balanced equation to show how each of the following ionic compounds dissociate in water.
   - **a.** sodium sulphate (Na_2SO_4)
   - **b.** potassium bromide (KBr)
   - **c.** potassium permanganate ($KMnO_4$)
   - **d.** sodium phosphate (Na_3PO_4)

Find the answers with the shortcodes:

(1.) l33 (2.) l3g

1.69.2 Ions and water hardness

This section is not examinable and is included as an example of ions in aqueous solution.

Definition: Water hardness
Water hardness is a measure of the mineral content of water. Minerals are substances such as calcite, quartz and mica that occur naturally as a result of geological processes.

Hard water is water that has a high mineral content. Water that has a low mineral content is known as **soft water**. If water has a high mineral content, it usually contains high levels of metal ions, mainly calcium (Ca) and magnesium (Mg). The calcium enters the water from either $CaCO_3$ (limestone or chalk) or from mineral deposits of $CaSO_4$. The main source of magnesium is a sedimentary rock called dolomite, $CaMg\,(CO_3)_2$. Hard water may also contain other metals as well as bicarbonates and sulphates.

NOTE: The simplest way to check whether water is hard or soft is to use the lather/froth test. If the water is very soft, soap will lather more easily when it is rubbed against the skin. With hard water this won't happen. Toothpaste will also not froth well in hard water.

A **water softener** works on the principle of **ion exchange**. Hard water passes through a media bed, usually made of resin beads that are supersaturated with sodium. As the water passes through the beads, the hardness minerals (e.g. calcium and magnesium) attach themselves to the beads. The sodium that was originally on the beads is released into the water. When the resin becomes saturated with calcium and magnesium, it must be recharged. A salt solution is passed through the resin. The sodium replaces the calcium and magnesium and these ions are released into the waste water and discharged.

1.70 Acid rain

(section shortcode: C10071)

This section is not examinable and is included as an example of ions in aqueous solution.

The acidity of rainwater comes from the natural presence of three substances (CO_2, NO, and SO_2) in the lowest layer of the atmosphere. These gases are able to dissolve in water and therefore make rain more acidic than it would otherwise be. Of these gases, carbon dioxide (CO_2) has the highest concentration and therefore contributes the most to the natural acidity of rainwater. We will look at each of these gases in turn.

Definition: Acid rain
Acid rain refers to the deposition of acidic components in rain, snow and dew. Acid rain occurs when sulphur dioxide and nitrogen oxides are emitted into the atmosphere, undergo chemical transformations and are absorbed by water droplets in clouds. The droplets then fall to earth as rain, snow, mist, dry dust, hail, or sleet. This increases the acidity of the soil and affects the chemical balance of lakes and streams.

1. **Carbon dioxide** Carbon dioxide reacts with water in the atmosphere to form **carbonic acid** (H_2CO_3).

$$CO_2 + H_2O \rightarrow H_2CO_3 \tag{10.2}$$

The carbonic acid dissociates to form hydrogen and hydrogen carbonate ions. It is the presence of hydrogen ions that lowers the pH of the solution making the rain acidic.

$$H_2CO_3 \rightarrow H^+ + HCO_3^- \tag{10.3}$$

2. **Nitric oxide** Nitric oxide (NO) also contributes to the natural acidity of rainwater and is formed during lightning storms when nitrogen and oxygen react. In air, NO is oxidised to form nitrogen dioxide (NO_2). It is the nitrogen dioxide which then reacts with water in the atmosphere to form **nitric acid** (HNO_3).

$$3NO_2\,(g) + H_2O\,(l) \rightarrow 2HNO_3\,(aq) + NO\,(g) \tag{10.4}$$

The nitric acid dissociates in water to produce hydrogen ions and nitrate ions. This again lowers the pH of the solution making it acidic.

$$HNO_3 \rightarrow H^+ + NO_3^- \tag{10.5}$$

3. **Sulphur dioxide** Sulphur dioxide in the atmosphere first reacts with oxygen to form sulphur trioxide, before reacting with water to form **sulphuric acid**.

$$2SO_2 + O_2 \rightarrow 2SO_3 \tag{10.6}$$

$$SO_3 + H_2O \rightarrow H_2SO_4 \tag{10.7}$$

Sulphuric acid dissociates in a similar way to the previous reactions.

$$H_2SO_4 \rightarrow HSO_4^- + H^+ \tag{10.8}$$

Although these reactions do take place naturally, human activities can greatly increase the concentration of these gases in the atmosphere, so that rain becomes far more acidic than it would otherwise be. The burning of fossil fuels in industries, vehicles etc is one of the biggest culprits. If the acidity of the rain drops to below 5, it is referred to as **acid rain**.

Acid rain can have a very damaging effect on the environment. In rivers, dams and lakes, increased acidity can mean that some species of animals and plants will not survive. Acid rain can also degrade soil minerals, producing metal ions that are washed into water systems. Some of these ions may be toxic e.g. Al^{3+}. From an economic perspective, altered soil pH can drastically affect agricultural productivity.

Acid rain can also affect buildings and monuments, many of which are made from marble and limestone. A chemical reaction takes place between $CaCO_3$ (limestone) and sulphuric acid to produce aqueous ions which can be easily washed away. The same reaction can occur in the lithosphere where limestone rocks are present e.g. limestone caves can be eroded by acidic rainwater.

$$H_2SO_4 + CaCO_3 \rightarrow CaSO_4 \cdot H_2O + CO_2 \tag{10.9}$$

1.70.1 Investigation : Acid rain

You are going to test the effect of 'acid rain' on a number of substances.

Materials needed:
samples of chalk, marble, zinc, iron, lead, dilute sulphuric acid, test tubes, beaker, glass dropper

Method:

1. Place a small sample of each of the following substances in a separate test tube: chalk, marble, zinc, iron and lead
2. To each test tube, add a few drops of dilute sulphuric acid.
3. Observe what happens and record your results.

Discussion questions:

- In which of the test tubes did reactions take place? What happened to the sample substances?
- What do your results tell you about the effect that acid rain could have on each of the following: buildings, soils, rocks and geology, water ecosystems?
- What precautions could be taken to reduce the potential impact of acid rain?

1.71 Electrolytes, ionisation and conductivity

www (section shortcode: C10072)

Conductivity in aqueous solutions, is a measure of the ability of water to conduct an electric current. The more **ions** there are in the solution, the higher its conductivity.

Definition: Conductivity
Conductivity is a measure of a solution's ability to conduct an electric current.

1.71.1 Electrolytes

An **electrolyte** is a material that *increases* the conductivity of water when dissolved in it. Electrolytes can be further divided into **strong electrolytes** and **weak electrolytes**.

Definition: Electrolyte
An electrolyte is a substance that contains free ions and behaves as an electrically conductive medium. Because they generally consist of ions in solution, electrolytes are also known as ionic solutions.

1. **Strong electrolytes** A strong electrolyte is a material that ionises completely when it is dissolved in water:

$$\mathrm{AB\ (s,\ l,\ g)} \rightarrow \mathrm{A^{+}(aq)} + \mathrm{B^{-}(aq)} \tag{10.10}$$

This is a **chemical change** because the original compound has been split into its component ions and bonds have been broken. In a strong electrolyte, we say that the *extent of ionisation* is high. In other words, the original material dissociates completely so that there is a high concentration of ions in the solution. An example is a solution of potassium nitrate:

$$\mathrm{KNO_3(s)} \rightarrow \mathrm{K^{+}(aq)} + \mathrm{NO_3^{-}(aq)} \tag{10.11}$$

2. **Weak electrolytes** A weak electrolyte is a material that goes into solution and will be surrounded by water molecules when it is added to water. However, not *all* of the molecules will dissociate into ions. The *extent of ionisation* of a weak electrolyte is low and therefore the concentration of ions in the solution is also low.

$$AB\,(s,l,g) \rightarrow AB\,(aq) \rightleftharpoons A^{+}\,\mathrm{(aq)} + B^{-}\,\mathrm{(aq)} \tag{10.12}$$

The following example shows that in the final solution of a weak electrolyte, some of the original compound *plus* some dissolved ions are present.

$$\mathrm{C_2H_3O_2H}\,(l) \rightarrow \mathrm{C_2H_3O_2H} \rightleftharpoons \mathrm{C_2H_3O_2^{-}\,(aq)} + \mathrm{H^{+}\,(aq)} \tag{10.13}$$

1.71.2 Non-electrolytes

A **non-electrolyte** is a material that does not increase the conductivity of water when dissolved in it. The substance goes into solution and becomes surrounded by water molecules, so that the molecules of the chemical become separated from each other. However, although the substance does dissolve, it is not changed in any way and no chemical bonds are broken. The change is a **physical change**. In the oxygen example below, the reaction is shown to be reversible because oxygen is only partially soluble in water and comes out of solution very easily.

$$C_2H_5OH\,(l) \rightarrow C_2H_5OH\,(aq) \tag{10.14}$$

$$O_2\,(g) \rightleftharpoons O_2\,(aq) \tag{10.15}$$

1.71.3 Factors that affect the conductivity of water

The conductivity of water is therefore affected by the following factors:

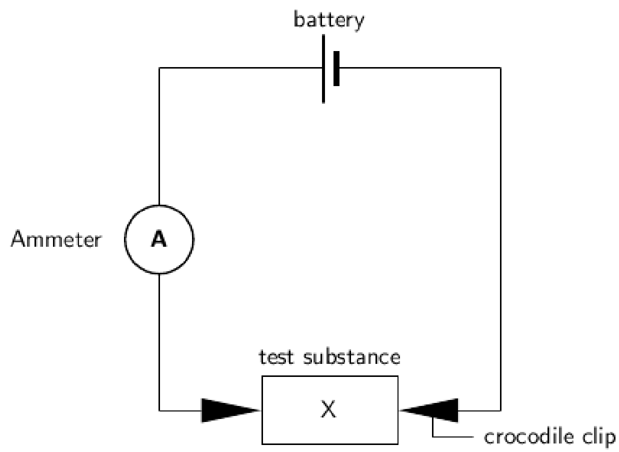
- The **type of substance** that dissolves in water. Whether a material is a strong electrolyte (e.g. potassium nitrate, KNO_3), a weak electrolyte (e.g. acetate, CH_3COOH) or a non-electrolyte (e.g. sugar, alcohol, oil) will affect the conductivity of water because the concentration of ions in solution will be different in each case.
- The **concentration of ions** in solution. The higher the concentration of ions in solution, the higher its conductivity will be.
- **Temperature.** The warmer the solution, the higher the solubility of the material being dissolved and therefore the higher the conductivity as well.

Experiment : Electrical conductivity

Aim:
To investigate the electrical conductivities of different substances and solutions.

Apparatus:
Solid salt ($NaCl$) crystals; different liquids such as distilled water, tap water, seawater, benzene and alcohol; solutions of salts e.g. $NaCl$, KBr; a solution of an acid (e.g. HCl) and a solution of a base (e.g. $NaOH$); torch cells; ammeter; conducting wire, crocodile clips and 2 carbon rods.

Method:
Set up the experiment by connecting the circuit as shown in the diagram below. In the diagram, 'X' represents the substance or solution that you will be testing. When you are using the solid crystals, the crocodile clips can be attached directly to each end of the crystal. When you are using solutions, two carbon rods are placed into the liquid and the clips are attached to each of the rods. In each case, complete the circuit and allow the current to flow for about 30 seconds. Observe whether the ammeter shows a reading.



Results:
Record your observations in a table similar to the one below:

Test substance	Ammeter reading

Table 10.1

What do you notice? Can you explain these observations?

Remember that for electricity to flow, there needs to be a movement of charged particles e.g. ions. With the solid NaCl crystals, there was no flow of electricity recorded on the ammeter. Although the solid is made up of ions, they are held together very tightly within the crystal lattice and therefore no current will flow. Distilled water, benzene and alcohol also don't conduct a current because they are *covalent compounds* and therefore do not contain ions.

The ammeter should have recorded a current when the salt solutions and the acid and base solutions were connected in the circuit. In solution, salts *dissociate* into their ions, so that these are free to move in the solution. Acids and bases behave in a similar way and dissociate to form hydronium and oxonium ions. Look at the following examples:

$$\mathrm{KBr} \rightarrow \mathrm{K}^{+} + \mathrm{Br}^{-} \tag{10.16}$$

$$\mathrm{NaCl} \rightarrow \mathrm{Na}^{+} + \mathrm{Cl}^{-} \tag{10.17}$$

$$\mathrm{HCl} + \mathrm{H_2O} \rightarrow \mathrm{H_3O}^{+} + \mathrm{Cl}^{-} \tag{10.18}$$

$$\mathrm{NaOH} \rightarrow \mathrm{Na}^{+} + \mathrm{OH}^{-} \tag{10.19}$$

Conclusions:
Solutions that contain free-moving ions are able to conduct electricity because of the movement of charged particles. Solutions that do not contain free-moving ions do not conduct electricity.

NOTE: Conductivity in streams and rivers is affected by the geology of the area where the water is flowing through. Streams that run through areas with granite bedrock tend to have lower conductivity because granite is made of materials that do not ionise when washed into the water. On the other hand, streams that run through areas with clay soils tend to have higher conductivity because the materials ionise when they are washed into the water. Pollution can also affect conductivity. A failing sewage system or an inflow of fertiliser runoff would raise the conductivity because of the presence of chloride, phosphate, and nitrate ions, while an oil spill (non-ionic) would lower the conductivity. It is very important that conductivity is kept within a certain acceptable range so that the organisms living in these water systems are able to survive.

1.72 Types of reactions

(section shortcode: C10073)

We will look at three types of reactions that occur in aqueous solutions. These are precipitation reactions, acid-base reactions and redox reactions. Precipitation and acid-base reactions are sometimes called ion exchange reactions. Redox reactions are electron transfer reactions. It is important to remember the difference between these two types of reactions. In ion exchange reactions ions are exchanged, in electron transfer reactions electrons are transferred. These terms will be explained further in the following sections.

Ion exchange reactions can be represented by:

$$\mathrm{AB\,(aq) + CD\,(aq) \rightarrow AD + CB} \tag{10.20}$$

Either AD or CB may be a solid or a gas. When a solid forms this is known as a precipitation reaction. If a gas is formed then this may be called a gas forming reaction. Acid-base reactions are a special class of ion exchange reactions and we will look at them seperately.

The formation of a precipitate or a gas helps to make the reaction happen. We say that the reaction is driven by the formation of a precipitate or a gas. All chemical reactions will only take place if there is something to make them happen. For some reactions this happens easily and for others it is harder to make the reaction occur.

Definition: Ion exchange reaction
A type of reaction where the positive ions exchange their respective negative ions due to a driving force.

NOTE: Ion exchange reactions are used in ion exchange chromatography. Ion exchange chromatography is used to purify water and as a means of softening water. Often when chemists talk about ion exchange, they mean ion exchange chromatography.

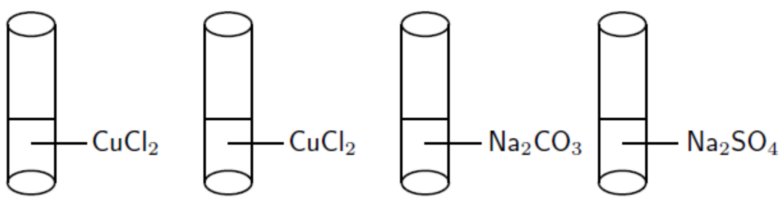
1.72.1 Precipitation reactions

Sometimes, ions in solution may react with each other to form a new substance that is *insoluble*. This is called a **precipitate**.

Definition: Precipitate
A precipitate is the solid that forms in a solution during a chemical reaction.

Demonstration : The reaction of ions in solution

Apparatus and materials:
4 test tubes; copper(II) chloride solution; sodium carbonate solution; sodium sulphate solution



Method:

1. Prepare 2 test tubes with approximately 5 ml of dilute Cu(II) chloride solution in each
2. Prepare 1 test tube with 5 ml sodium carbonate solution
3. Prepare 1 test tube with 5 ml sodium sulphate solution
4. Carefully pour the sodium carbonate solution into one of the test tubes containing copper(II) chloride and observe what happens
5. Carefully pour the sodium sulphate solution into the second test tube containing copper(II) chloride and observe what happens

Results:

1. A light blue precipitate forms when sodium carbonate reacts with copper(II) chloride
2. No precipitate forms when sodium sulphate reacts with copper(II) chloride

It is important to understand what happened in the previous demonstration. We will look at what happens in each reaction, step by step.

1. **Reaction 1:** Sodium carbonate reacts with copper(II) chloride.
   When these compounds react, a number of ions are present in solution: Cu^{2+}, Cl^{-}, Na^{+} and CO_3^{2-}.

Because there are lots of ions in solution, they will collide with each other and may recombine in different ways. The product that forms may be insoluble, in which case a precipitate will form, or the product will be soluble, in which case the ions will go back into solution. Let's see how the ions in this example could have combined with each other:

$$Cu^{2+} + CO_3^{2-} \rightarrow CuCO_3 \quad (10.21)$$

$$Cu^{2+} + 2Cl^{-} \rightarrow CuCl_2 \quad (10.22)$$

$$Na^{+} + Cl^{-} \rightarrow NaCl \quad (10.23)$$

$$2Na^{+} + CO_3^{2-} \rightarrow Na_2CO_3 \quad (10.24)$$

You can automatically exclude the reactions where sodium carbonate and copper(II) chloride are the products because these were the initial reactants. You also know that sodium chloride ($NaCl$) is soluble in water, so the remaining product (copper carbonate) must be the one that is insoluble. It is also possible to look up which salts are soluble and which are insoluble. If you do this, you will find that most carbonates are insoluble, therefore the precipitate that forms in this reaction must be $CuCO_3$. The reaction that has taken place between the ions in solution is as follows:

$$2Na^{+} + CO_3^{2-} + Cu^{2+} + 2Cl^{-} \rightarrow CuCO_3 + 2Na^{+} + 2Cl^{-} \quad (10.25)$$

2. **Reaction 2:** Sodium sulphate reacts with copper(II) chloride.
The ions that are present in solution are Cu^{2+}, Cl^{-}, Na^{+} and SO_4^{2-}. The ions collide with each other and may recombine in different ways. The possible combinations of the ions are as follows:

$$Cu^{2+} + SO_4^{2-} \rightarrow CuSO_4 \quad (10.26)$$

$$Cu^{2+} + 2Cl^{-} \rightarrow CuCl_2 \quad (10.27)$$

$$Na^{+} + Cl^{-} \rightarrow NaCl \quad (10.28)$$

$$Na^{+} + SO_4^{2-} \rightarrow Na_2SO_4 \quad (10.29)$$

If we look up which of these salts are soluble and which are insoluble, we see that most chlorides and most sulphates are soluble. This is why no precipitate forms in this second reaction. Even when the ions recombine, they immediately separate and go back into solution. The reaction that has taken place between the ions in solution is as follows:

$$2Na^{+} + SO_4^{2-} + Cu^{2+} + 2Cl^{-} \rightarrow 2Na^{+} + SO_4^{2-} + Cu^{2+} + 2Cl^{-} \quad (10.30)$$

Table 10.2 shows some of the general rules about the solubility of different salts based on a number of investigations:

Salt	Solubility
Nitrates	All are soluble
Potassium, sodium and ammonium salts	All are soluble
Chlorides, bromides and iodides	All are soluble except silver, lead(II) and mercury(II) salts (e.g. silver chloride)
Sulphates	All are soluble except lead(II) sulphate, barium sulphate and calcium sulphate
Carbonates	All are insoluble except those of potassium, sodium and ammonium
Compounds with fluorine	Almost all are soluble except those of magnesium, calcium, strontium (II), barium (II) and lead (II)
Perchlorates and acetates	All are soluble
Chlorates	All are soluble except potassium chlorate
Metal hydroxides and oxides	Most are insoluble

Table 10.2: General rules for the solubility of salts

Salts of carbonates, phosphates, oxalates, chromates and sulphides are generally insoluble.

1.72.2 Testing for common anions in solution

It is also possible to carry out tests to determine which ions are present in a solution. You should try to do each of these tests in class.

Test for a chloride

Prepare a solution of the unknown salt using distilled water and add a small amount of **silver nitrate** solution. If a white precipitate forms, the salt is either a chloride or a carbonate.

$$\mathrm{Cl^-} + Ag^+ + \mathrm{NO_3^-} \rightarrow \mathrm{AgCl} + \mathrm{NO_3^-} \tag{10.31}$$

(AgCl is white precipitate)

$$\mathrm{CO_3^{2-}} + 2Ag^+ + 2\mathrm{NO_3^-} \rightarrow Ag_2\mathrm{CO_3} + 2\mathrm{NO_3^-} \tag{10.32}$$

($\mathrm{Ag_2CO_3}$ is white precipitate)

The next step is to treat the precipitate with a small amount of **concentrated nitric acid**. If the precipitate remains unchanged, then the salt is a chloride. If carbon dioxide is formed, and the precipitate disappears, the salt is a carbonate.

$\mathrm{AgCl} + \mathrm{HNO_3} \rightarrow$ (no reaction; precipitate is unchanged)

$\mathrm{Ag_2CO_3} + 2\mathrm{HNO_3} \rightarrow 2\mathrm{AgNO_3} + \mathrm{H_2O} + \mathrm{CO_2}$ (precipitate disappears)

Test for a sulphate

Add a small amount of barium chloride solution to a solution of the test salt. If a white precipitate forms, the salt is either a sulphate or a carbonate.

$SO_4^{2-} + Ba^{2+} + Cl^- \rightarrow BaSO_4 + Cl^-$ ($BaSO_4$ is a white precipitate)

$CO_3^{2-} + Ba^{2+} + Cl^- \rightarrow BaCO_3 + Cl^-$ ($BaCO_3$ is a white precipitate)

If the precipitate is treated with nitric acid, it is possible to distinguish whether the salt is a sulphate or a carbonate (as in the test for a chloride).

$BaSO_4 + HNO_3 \rightarrow$ (no reaction; precipitate is unchanged)

$BaCO_3 + 2HNO_3 \rightarrow Ba(NO_3)_2 + H_2O + CO_2$ (precipitate disappears)

Test for a carbonate

If a sample of the dry salt is treated with a small amount of acid, the production of carbon dioxide is a positive test for a carbonate.

$\mathrm{Acid} + CO_3^{2-} \rightarrow CO_2$

If the gas is passed through limewater and the solution becomes milky, the gas is carbon dioxide.

$Ca\left(OH\right)_2 + CO_2 \rightarrow CaCO_3 + H_2O$ (It is the insoluble $CaCO_3$ precipitate that makes the limewater go milky)

Test for bromides and iodides

As was the case with the chlorides, the bromides and iodides also form precipitates when they are reacted with silver nitrate. Silver chloride is a white precipitate, but the silver bromide and silver iodide precipitates are both pale yellow. To determine whether the precipitate is a bromide or an iodide, we use chlorine water and carbon tetrachloride (CCl_4).

Chlorine water frees bromine gas from the bromide and colours the carbon tetrachloride a reddish brown.

Chlorine water frees iodine gas from an iodide and colours the carbon tetrachloride purple.

Precipitation reactions and ions in solution

1. Silver nitrate ($AgNO_3$) reacts with potassium chloride (KCl) and a white precipitate is formed.
   a. Write a balanced equation for the reaction that takes place.
   b. What is the name of the insoluble salt that forms?
   c. Which of the salts in this reaction are soluble?
2. Barium chloride reacts with sulphuric acid to produce barium sulphate and hydrochloric acid.
   a. Write a balanced equation for the reaction that takes place.
   b. Does a precipitate form during the reaction?
   c. Describe a test that could be used to test for the presence of barium sulphate in the products.

3. A test tube contains a clear, colourless salt solution. A few drops of silver nitrate solution are added to the solution and a pale yellow precipitate forms. Which one of the following salts was dissolved in the original solution?
   a. NaI
   b. KCl
   c. K_2CO_3
   d. Na_2SO_4
   (IEB Paper 2, 2005)

Find the answers with the shortcodes:

(1.) l3c (2.) l3O (3.) l3x

1.72.3 Other reactions in aqueous solutions

There are many types of reactions that can occur in aqueous solutions. In this section we will look at two of them: acid-base reactions and redox reactions. These reactions will be covered in more detail in Grade 11.

Acid-base reactions

Acid base reactions take place between acids and bases. In general, the products will be water and a salt (i.e. an ionic compound). An example of this type of reaction is:

$$\mathrm{NaOH\,(aq) + HCl\,(aq) \rightarrow NaCl\,(aq) + H_2O\ \ (l)} \tag{10.33}$$

This is an special case of an ion exchange reaction since the sodium in the sodium hydroxide swaps places with the hydrogen in the hydrogen chloride forming sodium chloride. At the same time the hydroxide and the hydrogen combine to form water.

Redox reactions

Redox reactions involve the exchange of electrons. One ion loses electrons and becomes more positive, while the other ion gains electrons and becomes more negative. To decide if a redox reaction has occurred we look at the charge of the atoms, ions or molecules involved. If one of them has become more positive and the other one has become more negative then a redox reaction has occurred. For example, sodium metal is oxidised to form sodium oxide (and sometimes sodium peroxide as well). The balanced equation for this is:

$$\mathrm{4Na + O_2 \rightarrow 2Na_2O} \tag{10.34}$$

In the above reaction sodium and oxygen are both neutral and so have no charge. In the products however, the sodium atom has a charge of $+1$ and the oxygen atom has a charge of -2. This tells us that the sodium has lost electrons and the oxygen has gained electrons. Since one species has become more positive and one more negative we can conclude that a redox reaction has occurred. We could also say that electrons have been transferred from one species to the other. (In this case the electrons were transferred from the sodium to the oxygen).

Demonstration: Oxidation of sodium metal

You will need a bunsen burner, a small piece of sodium metal and a metal spatula. Light the bunsen burner. Place the sodium metal on the spatula. Place the sodium in the flame. When the reaction finishes, you should observe a white powder on the spatula. This is a mixture of sodium oxide (Na_2O) and sodium peroxide (Na_2O_2).

WARNING: Sodium metal is very reactive. Sodium metal reacts vigourously with water and should never be placed in water. Be very careful when handling sodium metal.

1.72.4 Experiment: Reaction types

Aim:
To use experiments to determine what type of reaction occurs.

Apparatus:
Soluble salts (e.g. potassium nitrate, ammonium chloride, sodium carbonate, silver nitrate, sodium bromide); hydrochloric acid (HCl); sodium hydroxide($NaOH$); bromothymol blue; zinc metal; copper (II) sulphate; beakers; test-tubes

Method:

- For each of the salts, dissolve a small amount in water and observe what happens.
- Try dissolving pairs of salts (e.g. potassium nitrate and sodium carbonate) in water and observe what happens.
- Dissolve some sodium carbonate in hydrochloric acid and observe what happens.
- Carefully measure out 20cm^3 of sodium hydroxide into a beaker.
- Add some bromothymol blue to the sodium hydroxide
- Carefully add a few drops of hydrochloric acid to the sodium hydroxide and swirl. Repeat until you notice the colour change.
- Place the zinc metal into the copper sulphate solution and observe what happens.

Results:
Answer the following questions:

- What did you observe when you dissolved each of the salts in water?
- What did you observe when you dissolved pairs of salts in the water?
- What did you observe when you dissolved sodium carbonate in hydrochloric acid?
- Why do you think we used bromothymol blue when mixing the hydrochloric acid and the sodium hydroxide? Think about the kind of reaction that occurred.
- What did you observe when you placed the zinc metal into the copper sulphate?
- Classify each reaction as either precipitation, gas forming, acid-base or redox.
- What makes each reaction happen (i.e. what is the driving force)? Is it the formation of a precipitate or something else?
- What criteria would you use to determine what kind of reaction occurs?
- Try to write balanced chemical equations for each reaction

Conclusion:
We can see how we can classify reactions by performing experiments.

In the experiment above, you should have seen how each reaction type differs from the others. For example, a gas forming reaction leads to bubbles in the solution, a precipitation reaction leads to a precipitate forming, an acid-base reaction can be seen by adding a suitable indicator and a redox reaction can be seen by one metal disappearing and a deposit forming in the solution.

1.73 Summary

(section shortcode: C10074)

- The **polar** nature of water means that **ionic compounds** dissociate easily in aqueous solution into their component ions.
- **Ions** in solution play a number of roles. In the human body for example, ions help to regulate the internal environment (e.g. controlling muscle function, regulating blood pH). Ions in solution also determine water hardness and pH.
- **Water hardness** is a measure of the mineral content of water. Hard water has a high mineral concentration and generally also a high concentration of metal ions e.g. calcium and magnesium. The opposite is true for soft water.
- **Conductivity** is a measure of a solution's ability to conduct an electric current.
- An **electrolyte** is a substance that contains free ions and is therefore able to conduct an electric current. Electrolytes can be divided into **strong** and **weak** electrolytes, based on the extent to which the substance ionises in solution.
- A **non-electrolyte** cannot conduct an electric current because it dooes not contain free ions.
- The **type of substance**, the **concentration of ions** and the **temperature** of the solution affect its conductivity.
- There are three main types of reactions that occur in aqueous solutions. These are precipitation reactions, acid-base reactions and redox reactions.
- Precipitation and acid-base reactions are sometimes known as ion exchange reactions. Ion exchange reactions also include gas forming reactions.
- A **precipitate** is formed when ions in solution react with each other to form an insoluble product. Solubility 'rules' help to identify the precipitate that has been formed.
- A number of tests can be used to identify whether certain **anions** are present in a solution.
- An acid-base reaction is one in which an acid reacts with a base to form a salt and water.
- A redox reaction is one in which electrons are transferred from one substance to another.

1.74 End of chapter exercises

(section shortcode: C10075)

1. Give one word for each of the following descriptions:
   - **a.** the change in phase of water from a gas to a liquid
   - **b.** a charged atom
   - **c.** a term used to describe the mineral content of water
   - **d.** a gas that forms sulphuric acid when it reacts with water

2. Match the information in column A with the information in column B by writing only the letter (A to I) next to the question number (1 to 7)

Column A	**Column B**
1. A polar molecule	A. H_2SO_4
2. molecular solution	B. $CaCO_3$
3. Mineral that increases water hardness	C. $NaOH$
4. Substance that increases the hydrogen ion concentration	D. salt water
5. A strong electrolyte	E. calcium
6. A white precipitate	F. carbon dioxide
7. A non-conductor of electricity	G. potassium nitrate
	H. sugar water
	I. O_2

Table 10.3

3. For each of the following questions, choose the one correct answer from the list provided.

a. Which one of the following substances does not conduct electricity in the solid phase but is an electrical conductor when molten?

i. Cu
ii. $PbBr_2$
iii. H_2O
iv. I_2

(IEB Paper 2, 2003)

b. The following substances are dissolved in water. Which one of the solutions is basic?

i. sodium nitrate
ii. calcium sulphate
iii. ammonium chloride
iv. potassium carbonate

(IEB Paper 2, 2005)

4. Explain the difference between a weak electrolyte and a strong electrolyte. Give a generalised equation for each.

5. What factors affect the conductivity of water? How do each of these affect the conductivity?

6. For each of the following substances state whether they are molecular or ionic. If they are ionic, give a balanced reaction for the dissociation in water.

a. Methane (CH_4)
b. potassium bromide
c. carbon dioxide
d. hexane (C_6H_{14})
e. lithium fluoride (LiF)
f. magnesium chloride

7. Three test tubes (X, Y and Z) each contain a solution of an unknown potassium salt. The following observations were made during a practical investigation to identify the solutions in the test tubes: A: A white precipitate formed when silver nitrate ($AgNO_3$) was added to test tube Z. B: A white precipitate formed in test tubes X and Y when barium chloride ($BaCl_2$) was added. C: The precipitate in test tube X dissolved in hydrochloric acid (HCl) and a gas was released. D: The precipitate in test tube Y was insoluble in hydrochloric acid.
   a. Use the above information to identify the solutions in each of the test tubes X, Y and Z.
   b. Write a chemical equation for the reaction that took place in test tube X before hydrochloric acid was added.

   (DoE Exemplar Paper 2 2007)

Find the answers with the shortcodes:

(1.) l3a (2.) l3C (3a.) l31 (3b.) l3r (4.) lTl (5.) lTi
(6.) lbk (7.) l3Y

Quantitative Aspects of Chemical Change

1.75 Quantitative Aspects of Chemical Change

(section shortcode: C10076)

An equation for a chemical reaction can provide us with a lot of useful information. It tells us what the reactants and the products are in the reaction, and it also tells us the ratio in which the reactants combine to form products. Look at the equation below:

$\mathrm{Fe} + \mathrm{S} \rightarrow \mathrm{FeS}$

In this reaction, every atom of iron (Fe) will react with a single atom of sulphur (S) to form one molecule of iron sulphide (FeS). However, what the equation doesn't tell us, is the **quantities** or the **amount** of each substance that is involved. You may for example be given a small sample of iron for the reaction. How will you know how many atoms of iron are in this sample? And how many atoms of sulphur will you need for the reaction to use up all the iron you have? Is there a way of knowing what mass of iron sulphide will be produced at the end of the reaction? These are all very important questions, especially when the reaction is an industrial one, where it is important to know the quantities of reactants that are needed, and the quantity of product that will be formed. This chapter will look at how to quantify the changes that take place in chemical reactions.

1.76 The Mole

(section shortcode: C10077)

Sometimes it is important to know exactly how many particles (e.g. atoms or molecules) are in a sample of a substance, or what quantity of a substance is needed for a chemical reaction to take place.

You will remember from Relative atomic mass (Section 3.7.2: Relative atomic mass) that the **relative atomic mass** of an element, describes the mass of an atom of that element relative to the mass of an atom of carbon-12. So the mass of an atom of carbon (relative atomic mass is $12\,\mathrm{u}$) for example, is twelve times greater than the mass of an atom of hydrogen, which has a relative atomic mass of $1\,\mathrm{u}$. How can this information be used to help us to know what mass of each element will be needed if we want to end up with the same number of *atoms* of carbon and hydrogen?

Let's say for example, that we have a sample of $12\,\mathrm{g}$ carbon. What mass of *hydrogen* will contain the same number of atoms as $12\,\mathrm{g}$ carbon? We know that each atom of carbon weighs twelve times more than an atom of hydrogen. Surely then, we will only need $1\,\mathrm{g}$ of hydrogen for the number of atoms in the two samples to be the same? You will notice that the number of particles (in this case, *atoms*) in the two substances is the same when the ratio of their sample masses (12 g carbon: 1g hydrogen = 12:1) is the same as the ratio of their relative atomic masses (12 u: 1 u = 12:1).

To take this a step further, if you were to weigh out samples of a number of elements so that the mass of the sample was the same as the relative atomic mass of that element, you would find that the number of particles in each sample is $6,022 \times 10^{23}$. These results are shown in Table 11.1 below for a number of different elements. So, $24,31\,\text{g}$ of magnesium $(\text{relative atomic mass} = 24,31\,\text{u})$ for example, has the same number of atoms as $40,08\,\text{g}$ of calcium $(\text{relative atomic mass} = 40,08\,\text{u})$.

Element	**Relative atomic mass (u)**	**Sample mass (g)**	**Atoms in sample**
Hydrogen (H)	1	1	$6,022 \times 10^{23}$
Carbon (C)	12	12	$6,022 \times 10^{22}$
Magnesium (Mg)	24.31	24.31	$6,022 \times 10^{23}$
Sulphur (S)	32.07	32.07	$6,022 \times 10^{23}$
Calcium (Ca)	40.08	40.08	$6,022 \times 10^{23}$

Table 11.1: Table showing the relationship between the sample mass, the relative atomic mass and the number of atoms in a sample, for a number of elements.

This result is so important that scientists decided to use a special unit of measurement to define this quantity: the **mole** or 'mol'. A **mole** is defined as being an amount of a substance which contains the same number of particles as there are atoms in 12 g of carbon. In the examples that were used earlier, $24,31\,\text{g}$ magnesium is *one mole* of magnesium, while $40,08\,\text{g}$ of calcium is *one mole* of calcium. A mole of any substance always contains the same number of particles.

Definition: Mole
The mole (abbreviation 'n') is the SI (Standard International) unit for 'amount of substance'. It is defined as an amount of substance that contains the same number of particles (atoms, molecules or other particle units) as there are atoms in $12\,\text{g}$ carbon.

In one mole of any substance, there are $6,022 \times 10^{23}$ particles.

Definition: Avogadro's number
The number of particles in a mole, equal to $6,022 \times 10^{23}$. It is also sometimes referred to as the number of atoms in $12\,\text{g}$ of carbon-12.

If we were to write out Avogadro's number then it would look like: $602\,200\,000\,000\,000\,000\,000\,000$. This is a very large number. If we had this number of cold drink cans, then we could cover the surface of the earth to a depth of over $300\,\text{km}$! If you could count atoms at a rate of 10 million per second, then it would take you 2 billion years to count the atoms in one mole!

We can build up to the idea of Avogadro's number. For example, if you have 12 eggs then you have a dozen eggs. After this number we get a gross of eggs, which is 144 eggs. Finally if we wanted one mole of eggs this would be $6,022 \times 10^{23}$. That is a lot of eggs!

NOTE: The original hypothesis that was proposed by Amadeo Avogadro was that *'equal volumes of gases, at the same temperature and pressure, contain the same number of molecules'*. His ideas were not accepted by the scientific community and it was only four years after his death, that his original hypothesis was accepted and that it became known as 'Avogadro's Law'. In honour of his contribution to science, the number of particles in one mole was named *Avogadro's number*.

1.76.1 Moles and mass

1. Complete the following table:

Element	Relative atomic mass (u)	Sample mass (g)	Number of moles in the sample
Hydrogen	1.01	1.01	
Magnesium	24.31	24.31	
Carbon	12.01	24.02	
Chlorine	35.45	70.9	
Nitrogen		42.08	

Table 11.2

2. How many atoms are there in...
   a. 1 mole of a substance
   b. 2 moles of calcium
   c. 5 moles of phosphorus
   d. $24,31\ \mathrm{g}$ of magnesium
   e. $24,02\ \mathrm{g}$ of carbon

Find the answers with the shortcodes:

(1.) lgl (2.) lgi

1.77 Molar Mass

(section shortcode: C10078)

Definition: Molar mass
Molar mass (M) is the mass of 1 mole of a chemical substance. The unit for molar mass is **grams per mole** or $\mathrm{g \cdot mol^{-1}}$.

Refer to Table 11.1. You will remember that when the mass, in grams, of an element is equal to its relative atomic mass, the sample contains one mole of that element. This mass is called the **molar mass** of that element.

You may sometimes see the molar mass written as M_m. We will use M in this book, but you should be aware of the alternate notation.

It is worth remembering the following: On the periodic table, the relative atomic mass that is shown can be interpreted in two ways.

1. The mass of a *single, average atom* of that element relative to the mass of an atom of carbon.
2. The mass of *one mole of the element*. This second use is the molar mass of the element.

Element	**Relative atomic mass (u)**	**Molar mass ($\mathrm{g \cdot mol^{-1}}$)**	**Mass of one mole of the element (g)**
Magnesium	24,31	24,31	24,31
Lithium	6,94	6,94	6,94
Oxygen	16	16	16
Nitrogen	14,01	14,01	14,01
Iron	55,85	55,85	55,85

Table 11.3: The relationship between relative atomic mass, molar mass and the mass of one mole for a number of elements.

Exercise 11.1: Calculating the number of moles from mass Calculate the number of moles of iron (Fe) in a $11,7\,\mathrm{g}$ sample.

Solution to Exercise

Step 1. If we look at the periodic table, we see that the molar mass of iron is $55,85\,\mathrm{g \cdot mol^{-1}}$. This means that 1 mole of iron will have a mass of $55,85\,\mathrm{g}$.

Step 2. If 1 mole of iron has a mass of $55,85\,\text{g}$, then: the number of moles of iron in $111,7\,\text{g}$ must be:

$$\frac{111,7\text{g}}{55,85\text{g}\cdot\text{mol}^{-1}} = 2\,\text{mol} \tag{11.1}$$

There are 2 moles of iron in the sample.

Exercise 11.2: Calculating mass from moles You have a sample that contains 5 moles of zinc.

1. What is the mass of the zinc in the sample?
2. How many atoms of zinc are in the sample?

Solution to Exercise

Step 1. Molar mass of zinc is $65,38\,\text{g}\cdot\text{mol}^{-1}$, meaning that 1 mole of zinc has a mass of $65,38\,\text{g}$.

Step 2. If 1 mole of zinc has a mass of $65,38\,\text{g}$, then 5 moles of zinc has a mass of: $65,38\,\text{g}\times 5\,\text{mol} = 326,9\,\text{g}$ (answer to a)

Step 3.

$$5\times 6,022\times 10^{23} = 30,115\times 10^{23} \tag{11.2}$$

(answer to b)

1.77.1 Moles and molar mass

1. Give the molar mass of each of the following elements:
   a. hydrogen
   b. nitrogen
   c. bromine
2. Calculate the number of moles in each of the following samples:
   a. $21,62\,\text{g}$ of boron (B)
   b. $54,94\,\text{g}$ of manganese (Mn)
   c. $100,3\,\text{g}$ of mercury (Hg)
   d. $50\,\text{g}$ of barium (Ba)
   e. $40\,\text{g}$ of lead (Pb)

Find the answers with the shortcodes:

(1.) lg3 (2.) lgO

1.78 An equation to calculate moles and mass in chemical reactions

(section shortcode: C10079)

The calculations that have been used so far, can be made much simpler by using the following equation:

$$\mathbf{n}\,(\text{number of moles}) = \frac{\mathbf{m}\,(\text{mass of substance in g})}{\mathbf{M}\left(\text{molar mass of substance in g}\cdot\text{mol}^{-1}\right)} \tag{11.3}$$

TIP: Remember that when you use the equation $n = \frac{m}{M}$, the mass is always in grams (g) and molar mass is in grams per mol ($\text{g}\cdot\text{mol}^{-1}$).

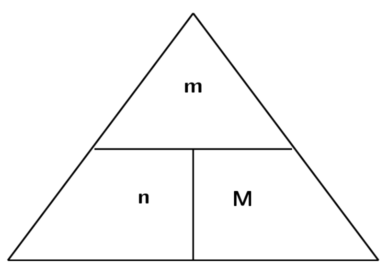
The equation can also be used to calculate mass and molar mass, using the following equations:

$$m = n \times M \tag{11.4}$$

and

$$M = \frac{m}{n} \tag{11.5}$$

The following diagram may help to remember the relationship between these three variables. You need to imagine that the horizontal line is like a 'division' sign and that the vertical line is like a 'multiplication' sign. So, for example, if you want to calculate 'M', then the remaining two letters in the triangle are 'm' and 'n' and 'm' is above 'n' with a division sign between them. In your calculation then, 'm' will be the numerator and 'n' will be the denominator.



Exercise 11.3: Calculating moles from mass Calculate the number of moles of copper there are in a sample that weighs $127\,\text{g}$.

Solution to Exercise

Step 1.

$$n = \frac{m}{M} \tag{11.6}$$

Step 2.

$$n = \frac{127}{63,55} = 2 \quad (11.7)$$

There are 2 moles of copper in the sample.

Exercise 11.4: Calculating mass from moles You are given a 5 mol sample of sodium. What mass of sodium is in the sample?

Solution to Exercise

Step 1.

$$m = n \times M \quad (11.8)$$

Step 2. M$_{Na} = 22,99\,\mathrm{g}\cdot\mathrm{mol}^{-1}$
Therefore,

$$m = 5 \times 22,99 = 114,95\,\mathrm{g} \quad (11.9)$$

The sample of sodium has a mass of $114,95\,\mathrm{g}$.

Exercise 11.5: Calculating atoms from mass Calculate the number of atoms there are in a sample of aluminium that weighs $80,94\,\mathrm{g}$.

Solution to Exercise

Step 1.

$$n = \frac{m}{M} = \frac{80,94}{26,98} = 3\,\mathrm{moles} \quad (11.10)$$

Step 2. Number of atoms in 3 mol aluminium $= 3 \times 6,022 \times 10^{23}$
There are $18,069 \times 10^{23}$ aluminium atoms in a sample of $80,94\,\mathrm{g}$.

1.78.1 Some simple calculations

1. Calculate the number of moles in each of the following samples:
   - **a.** $5,6\,\mathrm{g}$ of calcium
   - **b.** $0,02\,\mathrm{g}$ of manganese
   - **c.** $40\,\mathrm{g}$ of aluminium

2. A lead sinker has a mass of $5\,\mathrm{g}$.
   a. Calculate the number of moles of lead the sinker contains.
   b. How many lead atoms are in the sinker?
3. Calculate the mass of each of the following samples:
   a. $2,5\,\mathrm{mol}$ magnesium
   b. $12\,\mathrm{mol}$ lithium
   c. $4,5 \times 10^{25}$ atoms of silicon

www Find the answers with the shortcodes:

(1.) lgc (2.) lga (3.) lgx

1.79 Molecules and compounds

(section shortcode: C10080)

So far, we have only discussed moles, mass and molar mass in relation to *elements*. But what happens if we are dealing with a molecule or some other chemical compound? Do the same concepts and rules apply? The answer is 'yes'. However, you need to remember that all your calculations will apply to the *whole molecule*. So, when you calculate the molar mass of a molecule, you will need to add the molar mass of each atom in that compound. Also, the number of moles will also apply to the whole molecule. For example, if you have one mole of nitric acid ($\mathrm{HNO_3}$), it means you have $6,022 \times 10^{23}$ **molecules** of nitric acid in the sample. This also means that there are $6,022 \times 10^{23}$ **atoms** of hydrogen, $6,022 \times 10^{23}$ **atoms** of nitrogen and ($3 \times 6,022 \times 10^{23}$) **atoms** of oxygen in the sample.

In a balanced chemical equation, the number that is written in front of the element or compound, shows the **mole ratio** in which the reactants combine to form a product. If there are no numbers in front of the element symbol, this means the number is '1'.

e.g. $\mathrm{N_2 + 3H_2 \rightarrow 2NH_3}$

In this reaction, 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia.

Exercise 11.6: Calculating molar mass Calculate the molar mass of $\mathrm{H_2SO_4}$.

Solution to Exercise

Step 1. Hydrogen $= 1,008\,\mathrm{g \cdot mol^{-1}}$; Sulphur = $= 32,07\,\mathrm{g \cdot mol^{-1}}$; Oxygen $= 16\,\mathrm{g \cdot mol^{-1}}$

Step 2.

$$M_{(H_2SO_4)} = (2 \times 1,008) + (32,07) + (4 \times 16) = 98,09\,\mathrm{g \cdot mol^{-1}} \quad (11.11)$$

Exercise 11.7: Calculating moles from mass Calculate the number of moles there are in $1\,\mathrm{kg}$ of $\mathrm{MgCl_2}$.

Solution to Exercise

Step 1.

$$n = \frac{m}{M} \tag{11.12}$$

Step 2. **a.** Convert mass into grams

$$m = 1\mathrm{kg} \times 1000 = 1000\,\mathrm{g} \tag{11.13}$$

b. Calculate the molar mass of $\mathrm{MgCl_2}$.

$$M_{(\mathrm{MgCl_2})} = 24,31 + (2 \times 35,45) = 95,21\,\mathrm{g} \cdot \mathrm{mol}^{-1} \tag{11.14}$$

Step 3.

$$n = \frac{1000}{95,21} = 10,5\,\mathrm{mol} \tag{11.15}$$

There are $10,5\,\mathrm{moles}$ of magnesium chloride in a $1\,\mathrm{kg}$ sample.

Exercise 11.8: Calculating the mass of reactants and products Barium chloride and sulphuric acid react according to the following equation to produce barium sulphate and hydrochloric acid.

$\mathrm{BaCl_2} + \mathrm{H_2SO_4} \rightarrow \mathrm{BaSO_4} + 2\mathrm{HCl}$

If you have $2\,\mathrm{g}$ of $\mathrm{BaCl_2}$...

1. What quantity (in g) of $\mathrm{H_2SO_4}$ will you need for the reaction so that all the barium chloride is used up?
2. What mass of HCl is produced during the reaction?

Solution to Exercise

Step 1.

$$n = \frac{m}{M} = \frac{2}{208,24} = 0,0096\,\mathrm{mol} \tag{11.16}$$

Step 2. According to the balanced equation, 1 mole of $\mathrm{BaCl_2}$ will react with 1 mole of $\mathrm{H_2SO_4}$. Therefore, if $0,0096\,\mathrm{mol}$ of $\mathrm{BaCl_2}$ react, then there must be the same number of moles of $\mathrm{H_2SO_4}$ that react because their mole ratio is 1:1.

Step 3.

$$m = n \times M = 0,0096 \times 98,086 = 0,94\,\mathrm{g} \tag{11.17}$$

(answer to 1)

Step 4. According to the balanced equation, 2 moles of HCl are produced for every 1 mole of the two reactants. Therefore the number of moles of HCl produced is $(2 \times 0,0096)$, which equals $0,0096\,\mathrm{moles}$.

Step 5.

$$m = n \times M = 0,0192 \times 35,73 = 0,69\,\mathrm{g} \qquad (11.18)$$

(answer to 2)

1.79.1 Group work : Understanding moles, molecules and Avogadro's number

Divide into groups of three and spend about 20 minutes answering the following questions together:

1. What are the units of the mole? Hint: Check the definition of the mole.
2. You have a $56\,\mathrm{g}$ sample of iron sulphide (FeS)
   a. How many **moles** of FeS are there in the sample?
   b. How many **molecules** of FeS are there in the sample?
   c. What is the difference between a mole and a molecule?
3. The exact size of **Avogadro's number** is sometimes difficult to imagine.
   a. Write down Avogadro's number without using scientific notation.
   b. How long would it take to count to Avogadro's number? You can assume that you can count two numbers in each second.

Khan academy video on the mole - 1 www (Video: P10081)

1.79.2 More advanced calculations

1. Calculate the molar mass of the following chemical compounds:
   a. KOH
   b. $\mathrm{FeCl_3}$
   c. $\mathrm{Mg\,(OH)_2}$
2. How many moles are present in:
   a. $10\,\mathrm{g}$ of $\mathrm{Na_2SO_4}$
   b. $34\,\mathrm{g}$ of $\mathrm{Ca\,(OH)_2}$
   c. $2,45 \times 10^{23}$ molecules of $\mathrm{CH_4}$?
3. For a sample of $0,2\,\mathrm{moles}$ of potassium bromide (KBr), calculate...
   a. the number of moles of $\mathrm{K^+}$ ions
   b. the number of moles of $\mathrm{Br^-}$ ions

4. You have a sample containing $3\ \text{moles}$ of calcium chloride.
   a. What is the chemical formula of calcium chloride?
   b. How many calcium atoms are in the sample?
5. Calculate the mass of:
   a. $3\ \text{moles}$ of NH_4OH
   b. $4,2\ \text{moles}$ of $Ca(NO_3)_2$
6. $96,2\ \text{g}$ sulphur reacts with an unknown quantity of zinc according to the following equation: $Zn + S \rightarrow ZnS$
   a. What mass of zinc will you need for the reaction, if all the sulphur is to be used up?
   b. What mass of zinc sulphide will this reaction produce?
7. Calcium chloride reacts with carbonic acid to produce calcium carbonate and hydrochloric acid according to the following equation: $CaCl_2 + H_2CO_3 \rightarrow CaCO_3 + 2HCl$ If you want to produce $10\ \text{g}$ of calcium carbonate through this chemical reaction, what quantity (in g) of calcium chloride will you need at the start of the reaction?

Find the answers with the shortcodes:

(1.) lgC (2.) lgr (3.) lg1 (4.) lgY (5.) lgg (6.) lg4
(7.) lg2

1.80 The Composition of Substances

(section shortcode: C10082)

The **empirical formula** of a chemical compound is a simple expression of the relative number of each type of atom in that compound. In contrast, the **molecular formula** of a chemical compound gives the actual number of atoms of each element found in a molecule of that compound.

Definition: Empirical formula
The empirical formula of a chemical compound gives the relative number of each type of atom in that compound.

Definition: Molecular formula
The molecular formula of a chemical compound gives the exact number of atoms of each element in one molecule of that compound.

The compound *ethanoic acid* for example, has the molecular formula CH_3COOH or simply $C_2H_4O_2$. In one molecule of this acid, there are two carbon atoms, four hydrogen atoms and two oxygen atoms. The ratio of atoms in the compound is 2:4:2, which can be simplified to 1:2:1. Therefore, the empirical formula for this compound is CH_2O. The empirical formula contains the smallest whole number ratio of the elements that make up a compound.

Knowing either the empirical or molecular formula of a compound, can help to determine its composition in more detail. The opposite is also true. Knowing the *composition* of a substance can help you to determine its formula. There are four different types of composition problems that you might come across:

1. Problems where you will be given the formula of the substance and asked to calculate the percentage by mass of each element in the substance.
2. Problems where you will be given the percentage composition and asked to calculate the formula.
3. Problems where you will be given the products of a chemical reaction and asked to calculate the formula of one of the reactants. These are often referred to as combustion analysis problems.
4. Problems where you will be asked to find number of moles of waters of crystallisation.

Exercise 11.9: Calculating the percentage by mass of elements in a compound Calculate the percentage that each element contributes to the overall mass of sulphuric acid (H_2SO_4).

Solution to Exercise

Step 1. $\mathrm{Hydrogen} = 1,008 \times 2 = 2,016\,\mathrm{u}$
$\mathrm{Sulphur} = 32,07\,\mathrm{u}$
$\mathrm{Oxygen} = 4 \times 16 = 64\,\mathrm{u}$

Step 2. Use the calculations in the previous step to calculate the molecular mass of sulphuric acid.

$$\mathrm{Mass} = 2,016 + 32,07 + 64 = 98,09\,\mathrm{u} \tag{11.19}$$

Step 3. Use the equation:
$\mathrm{Percentage\ by\ mass} = \frac{\mathrm{atomic\ mass}}{\mathrm{molecular\ mass\ of\ H_2SO_4}} \times 100\%$
Hydrogen

$$\frac{2,016}{98,09} \times 100\% = 2,06\% \tag{11.20}$$

Sulphur

$$\frac{32,07}{98,09} \times 100\% = 32,69\% \tag{11.21}$$

Oxygen

$$\frac{64}{98,09} \times 100\% = 65,25\% \tag{11.22}$$

(You should check at the end that these percentages add up to 100%!)
In other words, in one molecule of sulphuric acid, hydrogen makes up 2,06% of the mass of the compound, sulphur makes up 32,69% and oxygen makes up 65,25%.

Exercise 11.10: Determining the empirical formula of a compound
A compound contains 52.2% carbon (C), 13.0% hydrogen (H) and 34.8% oxygen (O). Determine its empirical formula.

Solution to Exercise

Step 1. Carbon $= 52,2\ \mathrm{g}$, hydrogen $= 13,0\ \mathrm{g}$ and oxygen $= 34,8\ \mathrm{g}$

Step 2.

$$n = \frac{m}{M} \tag{11.23}$$

Therefore,

$$n\,(\mathrm{Carbon}) = \frac{52,2}{12,01} = 4,35\mathrm{mol} \tag{11.24}$$

$$n\,(\mathrm{Hydrogen}) = \frac{13,0}{1,008} = 12,90\mathrm{mol} \tag{11.25}$$

$$n\,(\mathrm{Oxygen}) = \frac{34,8}{16} = 2,18\mathrm{mol} \tag{11.26}$$

Step 3. In this case, the smallest number of moles is 2.18. Therefore...
Carbon

$$\frac{4,35}{2,18} = 2 \tag{11.27}$$

Hydrogen

$$\frac{12,90}{2,18} = 6 \tag{11.28}$$

Oxygen

$$\frac{2,18}{2,18} = 1 \tag{11.29}$$

Therefore the empirical formula of this substance is: $\mathrm{C_2H_6O}$. Do you recognise this compound?

Exercise 11.11: Determining the formula of a compound $207\ \mathrm{g}$ of lead combines with oxygen to form $239\ \mathrm{g}$ of a lead oxide. Use this information to work out the formula of the lead oxide (Relative atomic masses: $Pb = 207\ \mathrm{u}$ and $\mathrm{O} = 16\ \mathrm{u}$).

Solution to Exercise

Step 1.

$$239 - 207 = 32\,\mathrm{g} \tag{11.30}$$

Step 2.

$$n = \frac{m}{M} \tag{11.31}$$

Lead

$$\frac{207}{207} = 1\,\mathrm{mol} \tag{11.32}$$

Oxygen

$$\frac{32}{16} = 2\,\mathrm{mol} \tag{11.33}$$

Step 3. The mole ratio of $Pb:\mathrm{O}$ in the product is 1:2, which means that for every atom of lead, there will be two atoms of oxygen. The formula of the compound is $\mathrm{PbO_2}$.

Exercise 11.12: Empirical and molecular formula Vinegar, which is used in our homes, is a dilute form of acetic acid. A sample of acetic acid has the following percentage composition: 39,9% carbon, 6,7% hyrogen and 53,4% oxygen.

1. Determine the empirical formula of acetic acid.
2. Determine the molecular formula of acetic acid if the molar mass of acetic acid is $60\,\mathrm{g \cdot mol^{-1}}$.

Solution to Exercise

Step 1. In $100\,\mathrm{g}$ of acetic acid, there is $39,9\,\mathrm{g\,C}$, $6,7\,\mathrm{g\,H}$ and $53,4\,\mathrm{g\,O}$

Step 2. $n = \frac{m}{M}$

$$\begin{aligned} n_C &= \tfrac{39,9}{12} = 3,33\,\mathrm{mol} \\ n_H &= \tfrac{6,7}{1} = 6,7\,\mathrm{mol} \\ n_O &= \tfrac{53,4}{16} = 3,34\,\mathrm{mol} \end{aligned} \tag{11.34}$$

Step 3. Empirical formula is $\mathrm{CH_2O}$

Step 4. The molar mass of acetic acid using the empirical formula is $30\,\mathrm{g \cdot mol^{-1}}$. Therefore the actual number of moles of each element must be double what it is in the empirical formula.
The molecular formula is therefore $\mathrm{C_2H_4O_2}$ or $\mathrm{CH_3COOH}$

Exercise 11.13: Waters of crystallisation Aluminium trichloride ($AlCl_3$) is an ionic substance that forms crystals in the solid phase. Water molecules may be trapped inside the crystal lattice. We represent this as: $AlCl_3 \cdot nH_2O$. A learner heated some aluminium trichloride crystals until all the water had evaporated and found that the mass after heating was $2,8\,\text{g}$. The mass before heating was $5\,\text{g}$. What is the number of moles of water molecules in the aluminium trichloride?

Solution to Exercise

Step 1. We first need to find n, the number of water molecules that are present in the crystal. To do this we first note that the mass of water lost is $5 - 2,8 = 2,2$.

Step 2. The next step is to work out the mass ratio of aluminium trichloride to water and the mole ratio. The mass ratio is:

$$2,8 : 2,2 \tag{11.35}$$

To work out the mole ratio we divide the mass ratio by the molecular mass of each species:

$$\frac{2,8}{133} : \frac{2,2}{18} = 0,021 : 0,12 \tag{11.36}$$

Next we do the following:

$$0,021 \frac{1}{0,021} = 1 \tag{11.37}$$

and

$$\frac{0,12}{0,021} = 6 \tag{11.38}$$

So the mole ratio of aluminium trichloride to water is:

$$1 : 6 \tag{11.39}$$

Step 3. And now we know that there are 6 moles of water molecules in the crystal.

Khan academy video on molecular and empirical formulae - 1 (Video: P10083)

Khan academy video on mass composition - 1 (Video: P10084)

1.80.1 Moles and empirical formulae

1. Calcium chloride is produced as the product of a chemical reaction.
   a. What is the formula of calcium chloride?
   b. What percentage does each of the elements contribute to the mass of a molecule of calcium chloride?
   c. If the sample contains $5\,\mathrm{g}$ of calcium chloride, what is the mass of calcium in the sample?
   d. How many moles of calcium chloride are in the sample?
2. $13\,\mathrm{g}$ of zinc combines with $6,4\,\mathrm{g}$ of sulphur. What is the empirical formula of zinc sulphide?
   a. What mass of zinc sulphide will be produced?
   b. What percentage does each of the elements in zinc sulphide contribute to its mass?
   c. Determine the formula of zinc sulphide.
3. A calcium mineral consisted of 29,4% calcium, 23,5% sulphur and 47,1% oxygen by mass. Calculate the empirical formula of the mineral.
4. A chlorinated hydrocarbon compound was analysed and found to consist of 24,24% carbon, 4,04% hydrogen and 71,72% chlorine. From another experiment the molecular mass was found to be $99\mathrm{g}\cdot\mathrm{mol}^{-1}$. Deduce the empirical and molecular formula.

Find the answers with the shortcodes:

(1.) lgT (2.) lgb (3.) lgj (4.) lgD

1.81 Molar Volumes of Gases

(section shortcode: C10085)

It is possible to calculate the volume of one mole of gas at STP using what we know about gases.

1. *Write down the ideal gas equation* $\mathrm{pV} = \mathrm{nRT}$, therefore $V = \frac{nRT}{p}$
2. *Record the values that you know, making sure that they are in SI units* You know that the gas is under STP conditions. These are as follows: $p = 101,3\,\mathrm{kPa} = 101300\,Pa$ $n = 1\,\mathrm{mol}$ $R = 8,31\,\mathrm{J}\cdot\mathrm{K}^{-1}\cdot\mathrm{mol}^{-1}$ $T = 273\,\mathrm{K}$
3. *Substitute these values into the original equation.*

$$V = \frac{nRT}{p} \tag{11.40}$$

$$V = \frac{1\mathrm{mol} \times 8,31\mathrm{J}\cdot\mathrm{K}^{-1}\cdot\mathrm{mol}^{-1} \times 273\mathrm{K}}{101300 Pa} \tag{11.41}$$

4. *Calculate the volume of* $1\,\mathrm{mole}$ *of gas under these conditions* The volume of $1\,\mathrm{mole}$ of gas at STP is $22,4 \times 10^{-3}\,\mathrm{m}^3 = 22,4\mathrm{dm}^3$.

TIP: The standard units used for this equation are P in Pa, V in m^3 and T in K.
Remember also that $1\,000\,\mathrm{cm}^3 = 1\mathrm{dm}^3$ and $1\,000\,\mathrm{dm}^3 = 1\,\mathrm{m}^3$.

Exercise 11.14: Ideal Gas A sample of gas occupies a volume of $20\,\mathrm{dm}^3$, has a temperature of $200\,\mathrm{K}$ and has a pressure of $105\,\mathrm{Pa}$. Calculate the number of moles of gas that are present in the sample.

Solution to Exercise

Step 1. The only value that is not in SI units is volume. $V = 0,02\,\mathrm{m}^3$.

Step 2. We know that $\mathrm{pV} = \mathrm{nRT}$
Therefore,

$$n = \frac{pV}{RT} \tag{11.42}$$

Step 3.

$$n = \frac{105 \times 0,02}{8,31 \times 280} = \frac{2,1}{2326,8} = 0,0009\,\mathrm{moles} \tag{11.43}$$

1.81.1 Using the combined gas law

1. An enclosed gas(i.e. one in a sealed container) has a volume of $300\,\mathrm{cm}^3$ and a temperature of $300\,\mathrm{K}$. The pressure of the gas is $50\,\mathrm{kPa}$. Calculate the number of moles of gas that are present in the container.
2. What pressure will $3\,\mathrm{mol}$ of gaseous nitrogen exert if it is pumped into a container that has a volume of $25\,\mathrm{dm}^3$ at a temperature of $29\,^0\,\mathrm{C}$?
3. The volume of air inside a tyre is 19 litres and the temperature is $290\,\mathrm{K}$. You check the pressure of your tyres and find that the pressure is $190\,\mathrm{kPa}$. How many moles of air are present in the tyre?
4. Compressed carbon dioxide is contained within a gas cylinder at a pressure of 700 kPa. The temperature of the gas in the cylinder is $310\,\mathrm{K}$ and the number of moles of gas is $13\,\mathrm{mols}$ of carbon dioxide. What is the volume of the gas inside the cylinder?

Find the answers with the shortcodes:

(1.) lgW (2.) lgZ (3.) lgB (4.) lgK

1.82 Molar concentrations of liquids

(section shortcode: C10086)

A typical solution is made by dissolving some solid substance in a liquid. The amount of substance that is dissolved in a given volume of liquid is known as the **concentration** of the liquid. Mathematically, concentration (C) is defined as moles of solute (n) per unit volume (V) of solution.

$$C = \frac{n}{V} \quad (11.44)$$

For this equation, the units for volume are dm^3. Therefore, the unit of concentration is $\mathrm{mol} \cdot \mathrm{dm}^{-3}$. When concentration is expressed in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$ it is known as the **molarity** (M) of the solution. Molarity is the most common expression for concentration.

TIP: Do not confuse molarity (M) with molar mass (M). Look carefully at the question in which the M appears to determine whether it is concentration or molar mass.

Definition: Concentration
Concentration is a measure of the amount of solute that is dissolved in a given volume of liquid. It is measured in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$. Another term that is used for concentration is **molarity (M)**

Exercise 11.15: Concentration Calculations 1 If $3,5\,\mathrm{g}$ of sodium hydroxide (NaOH) is dissolved in $2,5\,\mathrm{dm}^3$ of water, what is the concentration of the solution in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$?

Solution to Exercise

Step 1.

$$n = \frac{m}{M} = \frac{3,5}{40} = 0,0875\,\mathrm{mol} \quad (11.45)$$

Step 2.

$$C = \frac{n}{V} = \frac{0,0875}{2,5} = 0,035 \quad (11.46)$$

The concentration of the solution is $0,035\,\mathrm{mol} \cdot \mathrm{dm}^{-3}$ or $0,035\,\mathrm{M}$

Exercise 11.16: Concentration Calculations 2 You have a $1\,\mathrm{dm}^3$ container in which to prepare a solution of potassium permanganate ($\mathrm{KMnO_4}$). What mass of $\mathrm{KMnO_4}$ is needed to make a solution with a concentration of $0,2\,\mathrm{M}$?

Solution to Exercise

Step 1.

$$C = \frac{n}{V} \tag{11.47}$$

therefore

$$n = C \times V = 0,2 \times 1 = 0,2\,\mathrm{mol} \tag{11.48}$$

Step 2.

$$m = n \times M = 0,2 \times 158,04 = 31,61\,\mathrm{g} \tag{11.49}$$

The mass of $\mathrm{KMnO_4}$ that is needed is $31,61\,\mathrm{g}$.

Exercise 11.17: Concentration Calculations 3 How much sodium chloride (in g) will one need to prepare $500\,\mathrm{cm}^3$ of solution with a concentration of $0,01\,\mathrm{M}$?

Solution to Exercise

Step 1.

$$V = \frac{500}{1000} = 0,5\,\mathrm{dm}^3 \tag{11.50}$$

Step 2.

$$n = C \times V = 0,01 \times 0,5 = 0,005\,\mathrm{mol} \tag{11.51}$$

Step 3.

$$m = n \times M = 0,005 \times 58,45 = 0,29\,\mathrm{g} \tag{11.52}$$

The mass of sodium chloride needed is $0,29\,\mathrm{g}$

1.82.1 Molarity and the concentration of solutions

1. $5,95\,\mathrm{g}$ of potassium bromide was dissolved in $400\,\mathrm{cm}^3$ of water. Calculate its molarity.
2. $100\,\mathrm{g}$ of sodium chloride (NaCl) is dissolved in $450\,\mathrm{cm}^3$ of water.
   - **a.** How many moles of NaCl are present in solution?
   - **b.** What is the volume of water (in dm^3)?
   - **c.** Calculate the concentration of the solution.
   - **d.** What mass of sodium chloride would need to be added for the concentration to become $5,7\,\mathrm{mol{\cdot}dm^{-3}}$?

3. What is the molarity of the solution formed by dissolving $80\,\mathrm{g}$ of sodium hydroxide (NaOH) in $500\,\mathrm{cm}^3$ of water?

4. What mass (g) of hydrogen chloride (HCl) is needed to make up $1000\,\mathrm{cm}^3$ of a solution of concentration $1\,\mathrm{mol}\cdot\mathrm{dm}^{-3}$?

5. How many moles of $\mathrm{H_2SO_4}$ are there in $250\,\mathrm{cm}^3$ of a $0,8\,\mathrm{M}$ sulphuric acid solution? What mass of acid is in this solution?

Find the answers with the shortcodes:

(1.) lgk (2.) lg0 (3.) lg8 (4.) lg9 (5.) lgX

1.83 Stoichiometric calculations

(section shortcode: C10087)

Stoichiometry is the calculation of the quantities of reactants and products in chemical reactions. It is also the numerical relationship between reactants and products. In representing chemical change showed how to write balanced chemical equations. By knowing the ratios of substances in a reaction, it is possible to use stoichiometry to calculate the amount of either reactants or products that are involved in the reaction. The examples shown below will make this concept clearer.

Exercise 11.18: Stoichiometric calculation 1 What volume of oxygen at S.T.P. is needed for the complete combustion of $2\,\mathrm{dm}^3$ of propane $(\mathrm{C_3H_8})$? (Hint: $\mathrm{CO_2}$ and $\mathrm{H_2O}$ are the products in this reaction (and in all combustion reactions))

Solution to Exercise

Step 1. $\mathrm{C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)}$

Step 2. From the balanced equation, the ratio of oxygen to propane in the reactants is 5:1.

Step 3. 1 volume of propane needs 5 volumes of oxygen, therefore $2\,\mathrm{dm}^3$ of propane will need $10\,\mathrm{dm}^3$ of oxygen for the reaction to proceed to completion.

Exercise 11.19: Stoichiometric calculation 2 What mass of iron (II) sulphide is formed when $5,6\,\mathrm{g}$ of iron is completely reacted with sulphur?

Solution to Exercise

Step 1. $\mathrm{Fe\,(s) + S\,(s) \rightarrow FeS\,(s)}$

Step 2.

$$n = \frac{m}{M} = \frac{5,6}{55,85} = 0,1\mathrm{mol} \tag{11.53}$$

Step 3. From the equation $1\,\mathrm{mole}$ of Fe gives $1\,\mathrm{mole}$ of FeS. Therefore, $0,1\,\mathrm{moles}$ of iron in the reactants will give $0,1\,\mathrm{moles}$ of iron sulphide in the product.

Step 4.

$$m = n \times M = 0,1 \times 87,911 = 8,79\,\mathrm{g} \tag{11.54}$$

The mass of iron (II) sulphide that is produced during this reaction is $8,79\,\mathrm{g}$.

When we are given a known mass of a reactant and are asked to work out how much product is formed, we are working out the theoretical yield of the reaction. In the laboratory chemists never get this amount of product. In each step of a reaction a small amount of product and reactants is 'lost' either because a reactant did not completely react or some of the product was left behind in the original container. Think about this. When you make your lunch or supper, you might be a bit hungry, so you eat some of the food that you are preparing. So instead of getting the full amount of food out (theoretical yield) that you started preparing, you lose some along the way.

Exercise 11.20: Industrial reaction to produce fertiliser Sulphuric acid ($\mathrm{H_2SO_4}$) reacts with ammonia ($\mathrm{NH_3}$) to produce the fertiliser ammonium sulphate (($\mathrm{NH_4)_2SO_4}$) according to the following equation:

$\mathrm{H_2SO_4\,(aq) + 2NH_3}\,(g) \rightarrow \mathrm{(NH_4)_2SO_4\,(aq)}$

What is the maximum mass of ammonium sulphate that can be obtained from $2,0\,\mathrm{kg}$ of sulphuric acid?

Solution to Exercise

Step 1.

$$n\,(\mathrm{H_2SO_4}) = \frac{m}{M} = \frac{2000\,\mathrm{g}}{98,078\,\mathrm{g \cdot mols^{-1}}} = 20,39\,\mathrm{mols} \tag{11.55}$$

Step 2. From the balanced equation, the mole ratio of $\mathrm{H_2SO_4}$ in the reactants to $\mathrm{(NH_4)\,_2SO_4}$ in the product is 1:1. Therefore, $20,39\,\mathrm{mols}$ of $\mathrm{H_2SO_4}$ of $\mathrm{(NH_4)\,_2SO_4}$.
The maximum mass of ammonium sulphate that can be produced is calculated as follows:

$$m = n \times M = 20,41\mathrm{mol} \times 132\,\mathrm{g \cdot mol^{-1}} = 2694\,\mathrm{g} \tag{11.56}$$

The maximum amount of ammonium sulphate that can be produced is $2,694\,\mathrm{kg}$.

Khan academy video on stoichiometry - 1 (Video: P10088)

1.83.1 Stoichiometry

1. Diborane, B_2H_6, was once considered for use as a rocket fuel. The combustion reaction for diborane is: $B_2H_6\,(g) + 3O_2\,(g) \rightarrow 2HBO_2\,(g) + 2H_2O\,(l)$ If we react $2,37\,\text{grams}$ of diborane, how many grams of water would we expect to produce?
2. Sodium azide is a commonly used compound in airbags. When triggered, it has the following reaction: $2NaN_3\,(s) \rightarrow 2Na\,(s) + 3N_2\,(g)$ If $23,4\,\text{grams}$ of sodium azide is used, how many moles of nitrogen gas would we expect to produce?
3. Photosynthesis is a chemical reaction that is vital to the existence of life on Earth. During photosynthesis, plants and bacteria convert carbon dioxide gas, liquid water, and light into glucose ($C_6H_{12}O_6$) and oxygen gas.
   a. Write down the equation for the photosynthesis reaction.
   b. Balance the equation.
   c. If $3\,\text{moles}$ of carbon dioxide are used up in the photosynthesis reaction, what mass of glucose will be produced?

(Presentation: P10089)

Find the answers with the shortcodes:

(1.) lgI (2.) lg5 (3.) lgN

1.84 Summary

(section shortcode: C10090)

- It is important to be able to quantify the changes that take place during a chemical reaction.
- The **mole (n)** is a SI unit that is used to describe an amount of substance that contains the same number of particles as there are atoms in $12\,g$ of carbon.
- The number of particles in a mole is called the **Avogadro constant** and its value is $6,022 \times 10^{23}$. These particles could be atoms, molecules or other particle units, depending on the substance.
- The **molar mass (M)** is the mass of one mole of a substance and is measured in grams per mole or $g{\cdot}mol^{-1}$. The numerical value of an element's molar mass is the same as its relative atomic mass. For a compound, the molar mass has the same numerical value as the molecular mass of that compound.

- The relationship between moles (n), mass in grams (m) and molar mass (M) is defined by the following equation:

$$n = \frac{m}{M} \tag{11.57}$$

- In a balanced chemical equation, the number in front of the chemical symbols describes the **mole ratio** of the reactants and products.
- The **empirical formula** of a compound is an expression of the relative number of each type of atom in the compound.
- The **molecular formula** of a compound describes the actual number of atoms of each element in a molecule of the compound.
- The formula of a substance can be used to calculate the **percentage by mass** that each element contributes to the compound.
- The **percentage composition** of a substance can be used to deduce its chemical formula.
- One mole of gas occupies a volume of $22,4\,\mathrm{dm}^3$.
- The **concentration** of a solution can be calculated using the following equation,

$$C = \frac{n}{V} \tag{11.58}$$

  where C is the concentration (in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$), n is the number of moles of solute dissolved in the solution and V is the volume of the solution (in dm^{-3}).
- **Molarity** is a measure of the concentration of a solution, and its units are $\mathrm{mol} \cdot \mathrm{dm}^{-3}$.
- **Stoichiometry** is the calculation of the quantities of reactants and products in chemical reactions. It is also the numerical relationship between reactants and products.
- The theoretical yield of a reaction is the maximum amount of product that we expect to get out of a reaction

1.85 End of chapter exercises

(section shortcode: C10091)

1. Write only the word/term for each of the following descriptions:
   a. the mass of one mole of a substance
   b. the number of particles in one mole of a substance
2. Multiple choice: Choose the one correct answer from those given.
   a. $5\,\mathrm{g}$ of magnesium chloride is formed as the product of a chemical reaction. Select the **true** statement from the answers below:
      a. 0.08 moles of magnesium chloride are formed in the reaction
      b. the number of atoms of Cl in the product is $0,6022 \times 10^{23}$
      c. the number of atoms of Mg is 0,05
      d. the atomic ratio of Mg atoms to Cl atoms in the product is 1:1
   b. 2 moles of oxygen gas react with hydrogen. What is the mass of oxygen in the reactants?
      a. 32 g
      b. 0,125 g
      c. 64 g
      d. 0,063 g

- **c.** In the compound potassium sulphate (K_2SO_4), oxygen makes up x% of the mass of the compound. x = ...
  - **a.** 36.8
  - **b.** 9,2
  - **c.** 4
  - **d.** 18,3
- **d.** The molarity of a $150\,cm^3$ solution, containing 5 g of $NaCl$ is...
  - **a.** $0,09\,M$
  - **b.** $5,7\ \times 10^{-4}\,M$
  - **c.** $0,57\,M$
  - **d.** $0,03\,M$

3. Calculate the number of moles in:
- **a.** 5 g of methane (CH_4)
- **b.** 3,4 g of hydrochloric acid
- **c.** 6,2 g of potassium permanganate ($KMnO_4$)
- **d.** 4 g of neon
- **e.** 9,6 kg of titanium tetrachloride ($TiCl_4$)

4. Calculate the mass of:
- **a.** 0,2 mols of potassium hydroxide (KOH)
- **b.** 0,47 mols of nitrogen dioxide
- **c.** 5,2 mols of helium
- **d.** 0,05 mols of copper (II) chloride ($CuCl_2$)
- **e.** $31,31 \times 10^{23}$ molecules of carbon monoxide (CO)

5. Calculate the percentage that each element contributes to the overall mass of:
- **a.** Chloro-benzene (C_6H_5Cl)
- **b.** Lithium hydroxide ($LiOH$)

6. CFC's (chlorofluorocarbons) are one of the gases that contribute to the depletion of the ozone layer. A chemist analysed a CFC and found that it contained 58,64% chlorine, 31,43% fluorine and 9,93% carbon. What is the empirical formula?

7. 14 g of nitrogen combines with oxygen to form 46 g of a nitrogen oxide. Use this information to work out the formula of the oxide.

8. Iodine can exist as one of three oxides (I_2O_4; I_2O_5; I_4O_9). A chemist has produced one of these oxides and wishes to know which one they have. If he started with 508 g of iodine and formed 652 g of the oxide, which form has he produced?

9. A fluorinated hydrocarbon (a hydrocarbon is a chemical compound containing hydrogen and carbon.) was analysed and found to contain 8,57% H, 51,05% C and 40,38% F.
- **a.** What is its empirical formula?
- **b.** What is the molecular formula if the molar mass is $94,1\,g \cdot mol^{-1}$?

10. Copper sulphate crystals often include water. A chemist is trying to determine the number of moles of water in the copper sulphate crystals. She weighs out 3 g of copper sulphate and heats this. After heating, she finds that the mass is 1,9 g. What is the number of moles of water in the crystals? (Copper sulphate is represented by $CuSO_4 \cdot xH_2O$).

11. $300\,\text{cm}^3$ of a $0,1\,\text{mol}\cdot\text{dm}^{-3}$ solution of sulphuric acid is added to $200\,\text{cm}^3$ of a $0,5\,\text{mol}\cdot\text{dm}^{-3}$ solution of sodium hydroxide.

a. Write down a balanced equation for the reaction which takes place when these two solutions are mixed.

b. Calculate the number of moles of sulphuric acid which were added to the sodium hydroxide solution.

c. Is the number of moles of sulphuric acid enough to fully neutralise the sodium hydroxide solution? Support your answer by showing all relevant calculations. (IEB Paper 2 2004)

12. A learner is asked to make $200\,\text{cm}^3$ of sodium hydroxide (NaOH) solution of concentration $0,5\,\text{mol}\cdot\text{dm}^{-3}$.

a. Determine the mass of sodium hydroxide pellets he needs to use to do this.

b. Using an accurate balance the learner accurately measures the correct mass of the NaOH pellets. To the pellets he now adds exactly $200\,\text{cm}^3$ of pure water. Will his solution have the correct concentration? Explain your answer.

c. The learner then takes $300\,\text{cm}^3$ of a $0,1\,\text{mol}\cdot\text{dm}^{-3}$ solution of sulphuric acid (H_2SO_4) and adds it to $200\,\text{cm}^3$ of a $0,5\,\text{mol}\cdot\text{dm}^{-3}$ solution of NaOH at 25^0C.

d. Write down a balanced equation for the reaction which takes place when these two solutions are mixed.

e. Calculate the number of moles of H_2SO_4 which were added to the NaOH solution.

f. Is the number of moles of H_2SO_4 calculated in the previous question enough to fully neutralise the NaOH solution? Support your answer by showing all the relevant calculations. (IEB Paper 2, 2004)

www Find the answers with the shortcodes:

(1.) IgR (2.a) Ign (2.b) IgQ (2.c) IgU (2.d) Igy (3.) IT3
(4.) ITO (5.) ITc (6.) ITx (7.) ITa (8.) ITC (9.) IT1
(10.) ITr (11.) IgP (12.) Igm

The Hydrosphere

1.86 Introduction

(section shortcode: C10092)

As far as we know, the Earth we live on is the only planet that is able to support life. Amongst other factors, the Earth is just the right distance from the sun to have temperatures that are suitable for life to exist. Also, the Earth's atmosphere has exactly the right type of gases in the right amounts for life to survive. Our planet also has **water** on its surface, which is something very unique. In fact, Earth is often called the 'Blue Planet' because most of it is covered in water. This water is made up of *freshwater* in rivers and lakes, the *saltwater* of the oceans and estuaries, *groundwater* and *water vapour*. Together, all these water bodies are called the **hydrosphere**.

NOTE: The total mass of the hydrosphere is approximately $1,4 \times 10^{18}\,\mathrm{tonnes!}$ (The volume of one tonne of water is approximately 1 cubic metre.)

1.87 Interactions of the hydrosphere

(section shortcode: C10093)

It is important to realise that the hydrosphere is not an isolated system, but rather interacts with other global systems, including the *atmosphere*, *lithosphere* and *biosphere*. These interactions are sometimes known collectively as the water cycle.

- *Atmosphere* When water is heated (e.g. by energy from the sun), it evaporates and forms water vapour. When water vapour cools again, it condenses to form liquid water which eventually returns to the surface by precipitation e.g. rain or snow. This cycle of water moving through the atmosphere and the energy changes that accompany it, is what drives weather patterns on earth.

- *Lithosphere* In the lithosphere (the ocean and continental crust at the Earth's surface), water is an important *weathering* agent, which means that it helps to break rock down into rock fragments and then soil. These fragments may then be transported by water to another place, where they are deposited. These two processes (weathering and the transporting of fragments) are collectively called *erosion*. Erosion helps to shape the earth's surface. For example, you can see this in rivers. In the upper streams, rocks are eroded and sediments are transported down the river and deposited on the wide flood plains lower down. On a bigger scale, river valleys in mountains have been carved out by the action of water, and cliffs and caves on rocky beach coastlines are also the result of weathering and erosion by water. The processes of weathering and erosion also increase the content of dissolved minerals in the water. These dissolved minerals are important for the plants and animals that live in the water.
- *Biosphere* In the biosphere, land plants absorb water through their roots and then transport this through their vascular (transport) system to stems and leaves. This water is needed in *photosynthesis*, the food production process in plants. Transpiration (evaporation of water from the leaf surface) then returns water back to the atmosphere.

1.88 Exploring the Hydrosphere

www (section shortcode: C10094)

The large amount of water on our planet is something quite unique. In fact, about 71% of the earth is covered by water. Of this, almost 97% is found in the oceans as saltwater, about 2.2% occurs as a solid in ice sheets, while the remaining amount (less than 1%) is available as freshwater. So from a human perspective, despite the vast amount of water on the planet, only a very small amount is actually available for human consumption (e.g. drinking water). In Reactions in aqueous solutions (Chapter 9) we looked at some of the reactions that occur in aqueous solution and saw some of the chemistry of water, in this section we are going to spend some time exploring a part of the hydrosphere in order to start appreciating what a complex and beautiful part of the world it is. After completing the following investigation, you should start to see just how important it is to know about the chemistry of water.

1.88.1 Investigation : Investigating the hydrosphere

1. For this exercise, you can choose any part of the hydrosphere that you would like to explore. This may be a rock pool, a lake, river, wetland or even just a small pond. The guidelines below will apply best to a river investigation, but you can ask similar questions and gather similar data in other areas. When choosing your study site, consider how accessible it is (how easy is it to get to?) and the problems you may experience (e.g. tides, rain).
2. Your teacher will provide you with the equipment you need to collect the following data. You should have at least one study site where you will collect data, but you might decide to have more if you want to compare your results in different areas. This works best in a river, where you can choose sites down its length.
   a. *Chemical data* Measure and record data such as temperature, pH, conductivity and dissolved oxygen at each of your sites. You may not know exactly what these measurements mean right now, but it will become clearer later.
   b. *Hydrological data* Measure the water velocity of the river and observe how the volume of water in the river changes as you move down its length. You can also collect a water sample in a clear bottle, hold it to the light and see whether the water is clear or whether it has particles in it.
   c. *Biological data* What types of animals and plants are found in or near this part of the hydrosphere? Are they specially adapted to their environment?

Record your data in a table like the one shown below:

	Site 1	Site 2	Site 3
Temperature			
pH			
Conductivity			
Dissolved oxygen			
Animals and plants			

Table 12.1

3. **Interpreting the data** Once you have collected and recorded your data, think about the following questions:

- How does the data you have collected vary at different sites?
- Can you explain these differences?
- What effect do you think *temperature*, *dissolved oxygen* and *pH* have on animals and plants that are living in the hydrosphere?
- Water is seldom 'pure'. It usually has lots of things dissolved (e.g. Mg^{2+}, Ca^{2+} and NO_3^- ions) or suspended (e.g. soil particles, debris) in it. Where do these substances come from?
- Are there any human activities near this part of the hydrosphere? What effect could these activities have on the hydrosphere?

1.89 The Importance of the Hydrosphere

www (section shortcode: C10095)

It is so easy sometimes to take our hydrosphere for granted and we seldom take the time to really think about the role that this part of the planet plays in keeping us alive. Below are just some of the very important functions of water in the hydrosphere:

- *Water is a part of living cells* Each cell in a living organism is made up of almost 75% water, and this allows the cell to function normally. In fact, most of the chemical reactions that occur in life, involve substances that are dissolved in water. Without water, cells would not be able to carry out their normal functions and life could not exist.
- *Water provides a habitat* The hydrosphere provides an important place for many animals and plants to live. Many gases (e.g. CO_2, O_2), nutrients e.g. nitrate (NO_3^-), nitrite (NO_2^-) and ammonium (NH_4^+) ions, as well as other ions (e.g. Mg^{2+} and Ca^{2+}) are dissolved in water. The presence of these substances is critical for life to exist in water.
- *Regulating climate* One of water's unique characteristics is its high *specific heat*. This means that water takes a long time to heat up and also a long time to cool down. This is important in helping to regulate temperatures on earth so that they stay within a range that is acceptable for life to exist. *Ocean currents* also help to disperse heat.
- *Human needs* Humans use water in a number of ways. Drinking water is obviously very important, but water is also used domestically (e.g. washing and cleaning) and in industry. Water can also be used to generate electricity through hydropower.

These are just a few of the very important functions that water plays on our planet. Many of the functions of water relate to its chemistry and to the way in which it is able to dissolve substances in it.

1.90 Threats to the Hydrosphere

(section shortcode: C10096)

It should be clear by now that the hydrosphere plays an extremely important role in the survival of life on Earth and that the unique properties of water allow various important chemical processes to take place which would otherwise not be possible. Unfortunately for us however, there are a number of factors that threaten our hydrosphere and most of these threats are because of human activities. We are going to focus on two of these issues: **overuse** and **pollution** and look at ways in which these problems can possibly be overcome.

1. **Pollution**
   Pollution of the hydrosphere is also a major problem. When we think of pollution, we sometimes only think of things like plastic, bottles, oil and so on. But any chemical that is present in the hydrosphere in an amount that is not what it should be is a pollutant. Animals and plants that live in the Earth's water bodies are specially adapted to surviving within a certain range of conditions. If these conditions are changed (e.g. through pollution), these organisms may not be able to survive. Pollution then, can affect entire aquatic ecosystems. The most common forms of pollution in the hydrosphere are *waste products* from humans and from industries, *nutrient pollution* e.g. fertiliser runoff which causes eutrophication (an excess of nutrients in the water leading to excessive plant growth) and toxic trace elements such as aluminium, mercury and copper to name a few. Most of these elements come from mines or from industries.
2. **Overuse of water**
   We mentioned earlier that only a very small percentage of the hydrosphere's water is available as freshwater. However, despite this, humans continue to use more and more water to the point where water *consumption* is fast approaching the amount of water that is *available*. The situation is a serious one, particularly in countries such as South Africa which are naturally dry and where water resources are limited. It is estimated that between 2020 and 2040, water supplies in South Africa will no longer be able to meet the growing demand for water in this country. This is partly due to population growth, but also because of the increasing needs of industries as they expand and develop. For each of us, this should be a very scary thought. Try to imagine a day without water... difficult isn't it? Water is so much a part of our lives, that we are hardly aware of the huge part that it plays in our daily lives.

1.90.1 Discussion : Creative water conservation

As populations grow, so do the demands that are placed on dwindling water resources. While many people argue that building dams helps to solve this water-shortage problem, there is evidence that dams are only a temporary solution and that they often end up doing far more ecological damage than good. The only sustainable solution is to reduce the *demand* for water, so that water supplies are sufficient to meet this. The more important question then is how to do this.

Discussion:
Divide the class into groups, so that there are about five people in each. Each group is going to represent a different sector within society. Your teacher will tell you which sector you belong to from the following: Farming, industry, city management or civil society (i.e. you will represent the ordinary 'man on the street'). In your groups, discuss the following questions as they relate to the group of people you represent: (Remember to take notes during your discussions, and nominate a spokesperson to give feedback to the rest of the class on behalf of your group)

- What steps could be taken by your group to conserve water?
- Why do you think these steps are *not* being taken?
- What incentives do you think could be introduced to encourage this group to conserve water more efficiently?

1.90.2 Investigation: Building of dams

In the previous discussion, we mentioned that there is evidence that dams are only a temporary solution to the water crisis. In this investigation you will look at why dams are a potentially bad solution to the problem.

For this investigation you will choose a dam that has been built in your area, or an area close to you. Make a note of which rivers are in the area. Try to answer the following questions:

- If possible talk to people who have lived in the area for a long time and try to get their opinion on how life changed since the dam was built. If it is not possible to talk to people in the area, then look for relevant literature on the area.
- Try to find out if any environmental impact assessments (this is where people study the environmnent and see what effect the proposed project has on the environment) were done before the dam was built. Why do you think this is important? Why do you think companies do not do these assessments?
- Look at how the ecology has changed. What was the ecology of the river? What is the current ecology? Do you think it has changed in a good way or a bad way?

Write a report or give a presentation in class on your findings from this investigation. Critically examine your findings and draw your own conclusion as to whether or not dams are only a short term solution to the growing water crisis.

It is important to realise that our hydrosphere exists in a delicate balance with other systems and that disturbing this balance can have serious consequences for life on this planet.

1.90.3 Group Project : School Action Project

There is a lot that can be done within a school to save water. As a class, discuss what actions could be taken by your class to make people more aware of how important it is to conserve water. Also consider what ways your school can save water. If possible, try to put some of these ideas into action and see if they really do conserve water.

1.91 How pure is our water?

www (section shortcode: C10097)

When you drink a glass of water you are not just drinking water, but many other substances that are dissolved into the water. Some of these come from the process of making the water safe for humans to drink, while others come from the environment. Even if you took water from a mountain stream (which is often considered pure and bottled for people to consume), the water would still have impurities in it. Water pollution increases the amount of impurities in the water and sometimes makes the water unsafe for drinking. In this section we will look at a few of the substances that make water impure and how we can make pure water. We will also look at the pH of water.

In Reactions in aqueous solutions (Chapter 9) we saw how compounds can dissolve in water. Most of these compounds (e.g. Na^{+}, Cl^{-}, Ca^{2+}, Mg^{2+}, etc.) are safe for humans to consume in the small amounts that are naturally present in water. It is only when the amounts of these ions rise above the safe levels that the water is considered to be polluted.

You may have noticed sometimes that when you pour a glass a water straight from the tap, it has a sharp smell. This smell is the same smell that you notice around swimming pools and is due to chlorine in the water. Chlorine is the most common compound added to water to make it safe for humans to use. Chlorine helps to remove bacteria and other biological contaminants in the water. Other methods to purify water include filtration (passing the water through a very fine mesh) and flocculation (a process of adding chemicals to the water to help remove small particles).

pH of water is also important. Water that is to basic (pH greater than 7) or to acidic (pH less than 7) may present problems when humans consume the water. If you have ever noticed after swimming that your eyes are red or your skin is itchy, then the pH of the swimming pool was probably to basic or to acidic. This shows you just how sensitive we are to the smallest changes in our environment. The pH of water depends on what ions are dissolved in the water. Adding chlorine to water often lowers the pH. You will learn more about pH in grade 11.

1.91.1 Experiment: Water purity

Aim:
To test the purity and pH of water samples

Apparatus:
pH test strips (you can find these at pet shops, they are used to test pH of fish tanks), microscope (or magnifying glass), filter paper, funnel, silver nitrate, concentrated nitric acid, barium chloride, acid, chlorine water (a solution of chlorine in water), carbon tetrachloride, some test-tubes or beakers, water samples from different sources (e.g. a river, a dam, the sea, tap water, etc.).

Method:

1. Look at each water sample and note if the water is clear or cloudy.
2. Examine each water sample under a microscope and note what you see.
3. Test the pH of each of the water samples.
4. Pour some of the water from each sample through filter paper.
5. Refer to Testing for common anions in solutions (Section 9.5.2: Testing for common anions in solution) for the details of common anion tests. Test for chloride, sulphate, carbonate, bromide and iodide in each of the water samples.

Results:
Write down what you saw when you just looked at the water samples. Write down what you saw when you looked at the water samples under a microscope. Where there any dissolved particles? Or other things in the water? Was there a difference in what you saw with just looking and with looking with a a microscope? Write down the pH of each water sample. Look at the filter paper from each sample. Is there sand or other particles on it? Which anions did you find in each sample?

Discussion:
Write a report on what you observed. Draw some conclusions on the purity of the water and how you can tell if water is pure or not.

Conclusion:
You should have seen that water is not pure, but rather has many substances dissolved in it.

1.91.2 Project: water purification

Prepare a presentation on how water is purified. This can take the form of a poster, or a presentation or a project. Things that you should look at are:

- Water for drinking (potable water)
- Distilled water and its uses
- Deionised water and its uses
- What methods are used to prepare water for various uses
- What regulations govern drinking water
- Why water needs to be purified
- How safe are the purification methods

1.92 Summary

(section shortcode: C10098)

- The **hydrosphere** includes all the water that is on Earth. Sources of water include freshwater (e.g. rivers, lakes), saltwater (e.g. oceans), groundwater (e.g. boreholes) and water vapour. Ice (e.g. glaciers) is also part of the hydrosphere.
- The hydrosphere interacts with other **global systems**, including the atmosphere, lithosphere and biosphere.
- The hydrosphere has a number of important **functions**. Water is a part of all living cells, it provides a habitat for many living organisms, it helps to regulate climate and it is used by humans for domestic, industrial and other use.
- Despite the importance of the hydrosphere, a number of factors threaten it. These include **overuse** of water, and **pollution**.
- Water is not pure, but has many substances dissolved in it.

1.93 End of chapter exercises

(section shortcode: C10099)

1. What is the hydrosphere? How does it interact with other global systems?
2. Why is the hydrosphere important?

Find the answers with the shortcodes:

(1.) lgp (2.) lgd

Part II

Physics

Vectors

1.94 Introduction

(section shortcode: P10000)

This chapter focuses on vectors. We will learn what a vector is and how it differs from everyday numbers. We will also learn how to add, subtract and multiply them and where they appear in Physics.

Are vectors Physics? No, vectors themselves are not Physics. Physics is just a description of the world around us. To describe something we need to use a language. The most common language used to describe Physics is Mathematics. Vectors form a very important part of the mathematical description of Physics, so much so that it is absolutely essential to master the use of vectors.

1.95 Scalars and Vectors

(section shortcode: P10001)

In Mathematics, you learned that a number is something that represents a quantity. For example if you have 5 books, 6 apples and 1 bicycle, the 5, 6, and 1 represent how many of each item you have.

These kinds of numbers are known as *scalars*.

Definition: Scalar
A scalar is a quantity that has only magnitude (size).

An extension to a scalar is a vector, which is a scalar with a direction. For example, if you travel 1 km down Main Road to school, the quantity **1 km down Main Road** is a vector. The "**1 km**" is the quantity (or scalar) and the "**down Main Road**" gives a direction.

In Physics we use the word *magnitude* to refer to the scalar part of the vector.

Definition: Vectors
A vector is a quantity that has both magnitude and direction.

A vector should tell you **how much** and **which way**.

For example, a man is driving his car east along a freeway at 100 km·hr^{-1}. What we have given here is a vector – the velocity. The car is moving at 100 $k \cdot h^{-1}$ (this is the magnitude) and we know where it is going – east (this is the direction). Thus, we know the speed and direction of the car. These two quantities, a magnitude and a direction, form a vector we call velocity.

1.96 Notation

(section shortcode: P10002)

Vectors are different to scalars and therefore have their own notation.

1.96.1 Mathematical Representation

There are many ways of writing the symbol for a vector. Vectors are denoted by symbols with an arrow pointing to the right above it. For example, $\vec{a}$, $\vec{v}$ and $\vec{F}$ represent the vectors acceleration, velocity and force, meaning they have both a magnitude and a direction.

Sometimes just the magnitude of a vector is needed. In this case, the arrow is omitted. In other words, F denotes the magnitude of the vector $\vec{F}$. $|\vec{F}|$ is another way of representing the magnitude of a vector.

1.96.2 Graphical Representation

Vectors are drawn as arrows. An arrow has both a magnitude (how long it is) and a direction (the direction in which it points). The starting point of a vector is known as the *tail* and the end point is known as the *head*.

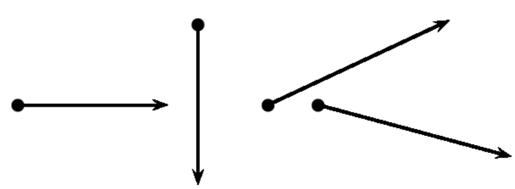


Figure 13.1: Examples of vectors

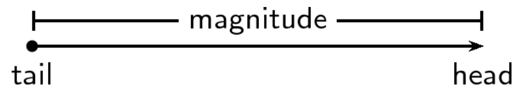


Figure 13.2: Parts of a vector

1.97 Directions

(section shortcode: P10003)

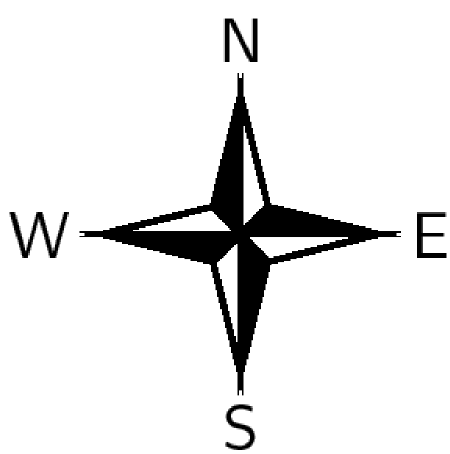
There are many acceptable methods of writing vectors. As long as the vector has a magnitude and a direction, it is most likely acceptable. These different methods come from the different methods of expressing a direction for a vector.

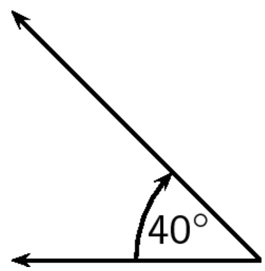
1.97.1 Relative Directions

The simplest method of expressing direction is with relative directions: to the left, to the right, forward, backward, up and down.

1.97.2 Compass Directions

Another common method of expressing directions is to use the points of a compass: North, South, East, and West. If a vector does not point exactly in one of the compass directions, then we use an angle. For example, we can have a vector pointing 40° North of West. Start with the vector pointing along the West direction: Then rotate the vector towards the north until there is a 40° angle between the vector and the West. The direction of this vector can also be described as: W 40° N (West 40° North); or N 50° W (North 50° West)

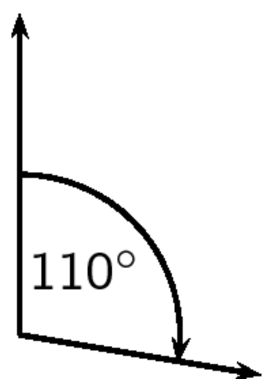




1.97.3 Bearing

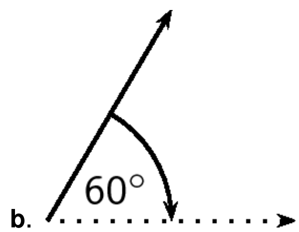
The final method of expressing direction is to use a *bearing*. A bearing is a direction relative to a fixed point.

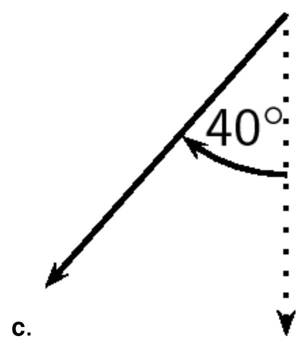
Given just an angle, the convention is to define the angle with respect to the North. So, a vector with a direction of 110° has been rotated clockwise 110° relative to the North. A bearing is always written as a three digit number, for example 275° or 080° (for 80°).



Scalars and Vectors

1. Classify the following quantities as scalars or vectors:
   - **a.** 12 km
   - **b.** 1 m south
   - **c.** 2 $m \cdot s^{-1}$, 45°
   - **d.** 075°, 2 cm
   - **e.** 100 $k \cdot h^{-1}$, 0°
2. Use two different notations to write down the direction of the vector in each of the following diagrams:





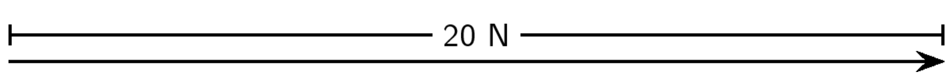
Find the answers with the shortcodes:

(1.) l4s (2.) l4H

1.98 Drawing Vectors

(section shortcode: P10004)

In order to draw a vector accurately we must specify a scale and include a reference direction in the diagram. A scale allows us to translate the length of the arrow into the vector's magnitude. For instance if one chose a scale of 1 cm = 2 N (1 cm represents 2 N), a force of 20 N towards the East would be represented as an arrow 10 cm long. A reference direction may be a line representing a horizontal surface or the points of a compass.



Method: Drawing Vectors

1. Decide upon a scale and write it down.
2. Determine the length of the arrow representing the vector, by using the scale.
3. Draw the vector as an arrow. Make sure that you fill in the arrow head.
4. Fill in the magnitude of the vector.

Exercise 13.1: Drawing vectors Represent the following vector quantities:

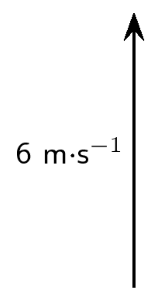
1. 6 $m \cdot s^{-1}$ north
2. 16 m east

Solution to Exercise

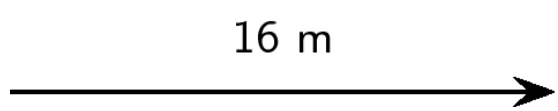
Step 1. **a.** 1 cm = 2 $m \cdot s^{-1}$
b. 1 cm = 4 m

Step 2. **a.** If 1 cm = 2 $m \cdot s^{-1}$, then 6 $m \cdot s^{-1}$ = 3 cm
b. If 1 cm = 4 m, then 16 m = 4 cm

Step 3. **a.** Scale used: 1 cm = 2 $m \cdot s^{-1}$ Direction = North



b. Scale used: 1 cm = 4 m Direction = East



1.98.1 Drawing Vectors

Draw each of the following vectors to scale. Indicate the scale that you have used:

1. 12 km south

2. 1,5 m N 45° W
3. 1 m·s^{-1}, 20° East of North
4. 50 km·hr^{-1}, 085°
5. 5 mm, 225°

Find the answers with the shortcodes:

(1.) l46

1.99 Mathematical Properties of Vectors

(section shortcode: P10005)

Vectors are mathematical objects and we need to understand the mathematical properties of vectors, like adding and subtracting.

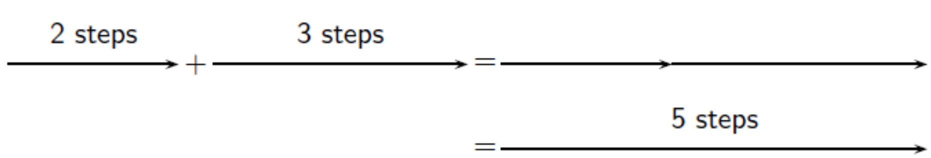
For all the examples in this section, we will use displacement as our vector quantity. Displacement was discussed in Grade 10.

Displacement is defined as the distance together with direction of the straight line joining a final point to an initial point.

Remember that displacement is just one example of a vector. We could just as well have decided to use forces or velocities to illustrate the properties of vectors.

1.99.1 Adding Vectors

When vectors are added, we need to add both a magnitude **and** a direction. For example, take 2 steps in the forward direction, stop and then take another 3 steps in the forward direction. The first 2 steps is a displacement vector and the second 3 steps is also a displacement vector. If we did not stop after the first 2 steps, we would have taken 5 steps in the forward direction in total. Therefore, if we add the displacement vectors for 2 steps and 3 steps, we should get a total of 5 steps in the forward direction. Graphically, this can be seen by first following the first vector two steps forward and then following the second one three steps forward (ie. in the same direction):



We add the second vector at the end of the first vector, since this is where we now are after the first vector has acted. The vector from the tail of the first vector (the starting point) to the head of the last (the end point) is then the sum of the vectors. This is the *head-to-tail* method of vector addition.

As you can convince yourself, the order in which you add vectors does not matter. In the example above, if you decided to first go 3 steps forward and then another 2 steps forward, the end result would still be 5 steps forward.

The final answer when adding vectors is called the **resultant**. The resultant displacement in this case will be 5 steps forward.

Definition: Resultant of Vectors
The resultant of a number of vectors is the single vector whose effect is the same as the individual vectors acting together.

In other words, the individual vectors can be replaced by the resultant – the overall effect is the same. If vectors $\vec{a}$ and $\vec{b}$ have a resultant $\vec{R}$, this can be represented mathematically as,

$$\vec{R} = \vec{a} + \vec{b}\,. \tag{13.1}$$

Let us consider some more examples of vector addition using displacements. The arrows tell you how far to move and in what direction. Arrows to the right correspond to steps forward, while arrows to the left correspond to steps backward. Look at all of the examples below and check them.

$$\xrightarrow{\text{1 step}} + \xrightarrow{\text{1 step}} = \xrightarrow{\text{2 steps}} = \xrightarrow{\text{2 steps}}$$

This example says 1 step forward and then another step forward is the same as an arrow twice as long – two steps forward.

$$\xleftarrow{\text{1 step}} + \xleftarrow{\text{1 step}} = \xleftarrow{\text{2 steps}} = \xleftarrow{\text{2 steps}}$$

This examples says 1 step backward and then another step backward is the same as an arrow twice as long – two steps backward.

It is sometimes possible that you end up back where you started. In this case the net result of what you have done is that you have gone nowhere (your start and end points are at the same place). In this case, your resultant displacement is a vector with length zero units. We use the symbol $\vec{0}$ to denote such a vector:

$$\xrightarrow{\text{1 step}} + \xleftarrow{\text{1 step}} = \overset{\text{1 step}}{\underset{\text{1 step}}{\rightleftarrows}} = \vec{0}$$

$$\xleftarrow{\text{1 step}} + \xrightarrow{\text{1 step}} = \overset{\text{1 step}}{\underset{\text{1 step}}{\leftrightarrows}} = \vec{0}$$

Check the following examples in the same way. Arrows up the page can be seen as steps left and arrows down the page as steps right.

Try a couple to convince yourself!

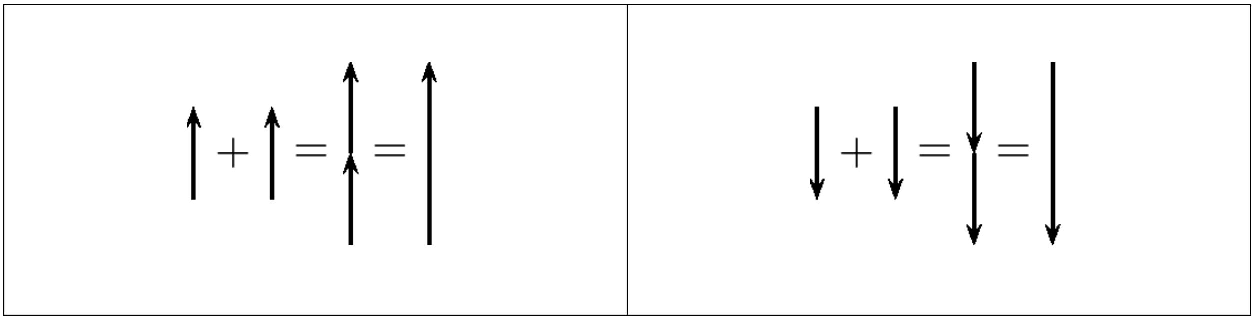


Table 13.1

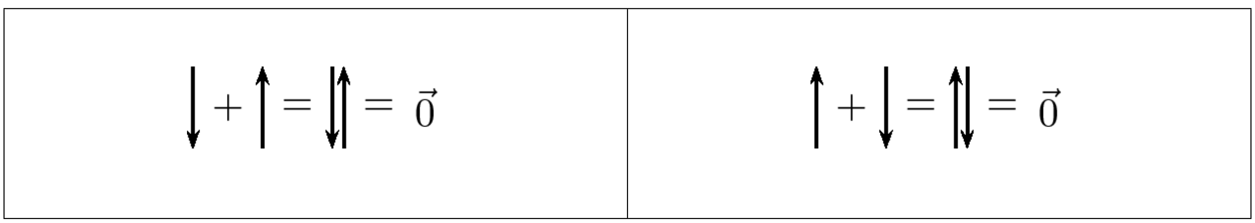


Table 13.2

It is important to realise that the directions are not special– ‘forward and backwards’ or ‘left and right’ are treated in the same way. The same is true of any set of parallel directions:

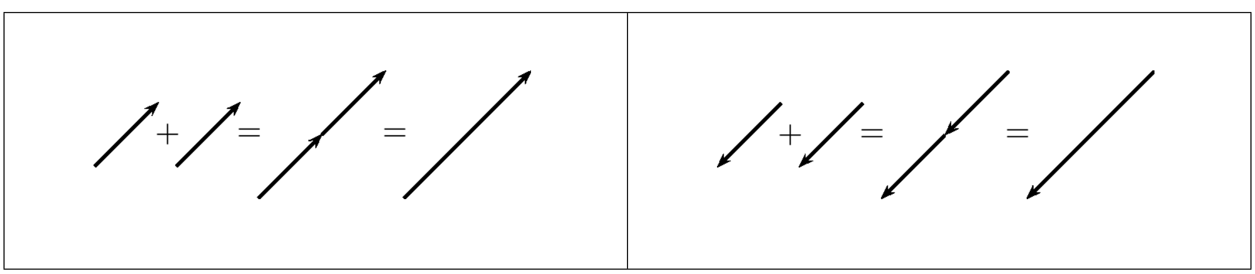


Table 13.3

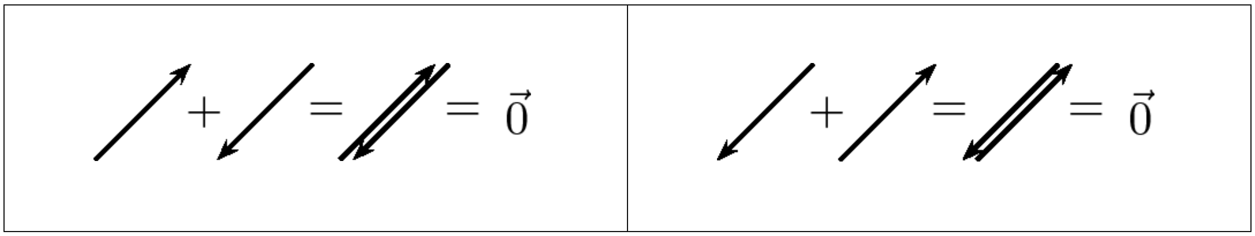


Table 13.4

In the above examples the separate displacements were parallel to one another. However the same head-to-tail technique of vector addition can be applied to vectors in any direction.

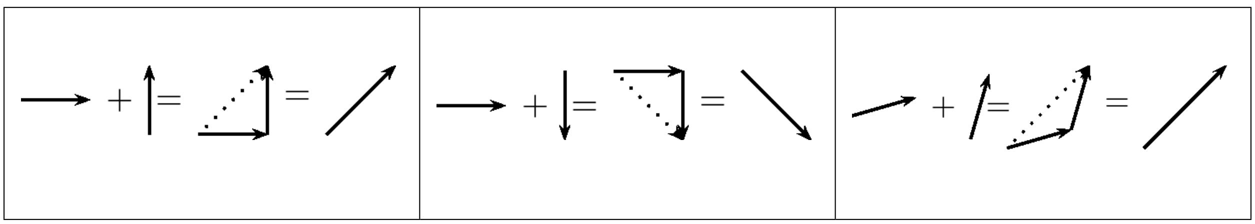


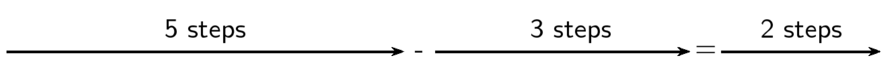
Table 13.5

Now you have discovered one use for vectors; describing resultant displacement – how far and in what direction you have travelled after a series of movements.

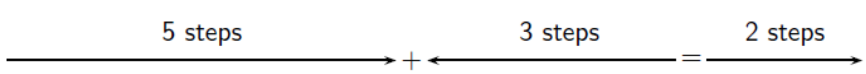
Although vector addition here has been demonstrated with displacements, all vectors behave in exactly the same way. Thus, if given a number of forces acting on a body you can use the same method to determine the resultant force acting on the body. We will return to vector addition in more detail later.

1.99.2 Subtracting Vectors

What does it mean to subtract a vector? Well this is really simple; if we have 5 apples and we subtract 3 apples, we have only 2 apples left. Now lets work in steps; if we take 5 steps forward and then subtract 3 steps forward we are left with only two steps forward:



What have we done? You originally took 5 steps forward but then you took 3 steps back. That backward displacement would be represented by an arrow pointing to the left (backwards) with length 3. The net result of adding these two vectors is 2 steps forward:



Thus, subtracting a vector from another is the same as adding a vector in the opposite direction (i.e. subtracting 3 steps forwards is the same as adding 3 steps backwards).

TIP: Subtracting a vector from another is the same as adding a vector in the opposite direction.

In the problem, motion in the forward direction has been represented by an arrow to the right. Arrows to the right are positive and arrows to the left are negative. More generally, vectors in opposite directions differ in sign (i.e. if we define up as positive, then vectors acting down are negative). Thus, changing the sign of a vector simply reverses its direction:

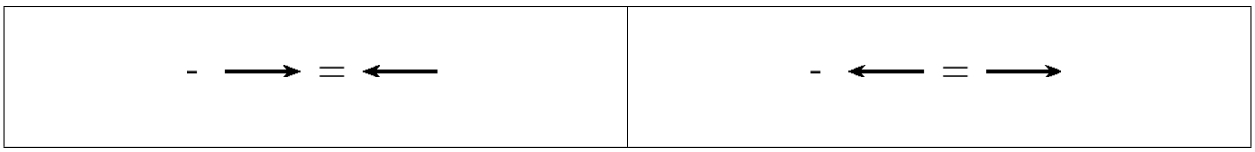


Table 13.6

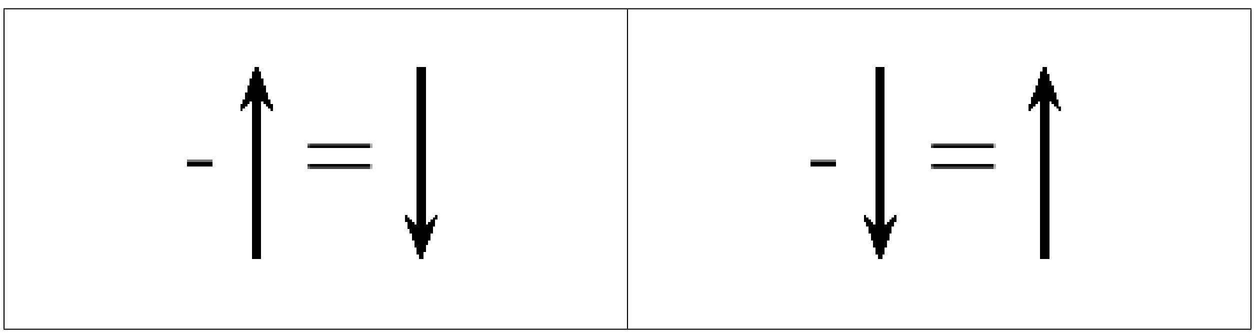


Table 13.7

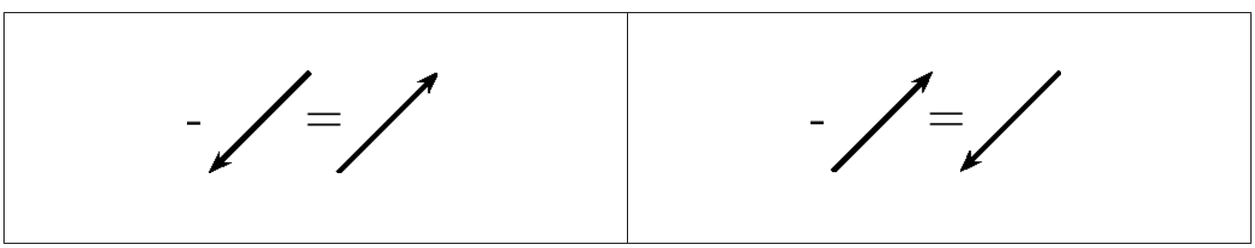


Table 13.8

In mathematical form, subtracting $\vec{a}$ from $\vec{b}$ gives a new vector $\vec{c}$:

$$\begin{aligned} \vec{c} &= \vec{b} - \vec{a} \\ &= \vec{b} + \left(-\vec{a}\right) \end{aligned} \tag{13.2}$$

This clearly shows that subtracting vector $\vec{a}$ from $\vec{b}$ is the same as adding $\left(-\vec{a}\right)$ to $\vec{b}$. Look at the following examples of vector subtraction.

$$\longrightarrow - \longrightarrow = \longrightarrow + \longleftarrow = \vec{0}$$

$$\longrightarrow - \longleftarrow = \longrightarrow + \longrightarrow = \longrightarrow\!\!\!\!\longrightarrow$$

1.99.3 Scalar Multiplication

What happens when you multiply a vector by a scalar (an ordinary number)?

Going back to normal multiplication we know that 2×2 is just 2 groups of 2 added together to give 4. We can adopt a similar approach to understand how vector multiplication works.

$$2 \times \longrightarrow = \longrightarrow + \longrightarrow = \longrightarrow\!\!\!\!\longrightarrow$$

1.100 Techniques of Vector Addition

(section shortcode: P10006)

Now that you have learned about the mathematical properties of vectors, we return to vector addition in more detail. There are a number of techniques of vector addition. These techniques fall into two main categories - graphical and algebraic techniques.

1.100.1 Graphical Techniques

Graphical techniques involve drawing accurate scale diagrams to denote individual vectors and their resultants. We next discuss the two primary graphical techniques, the head-to-tail technique and the parallelogram method.

The Head-to-Tail Method

In describing the mathematical properties of vectors we used displacements and the head-to-tail graphical method of vector addition as an illustration. The head-to-tail method of graphically adding vectors is a standard method that must be understood.

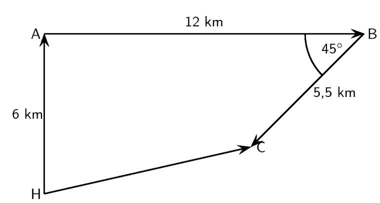
Method: Head-to-Tail Method of Vector Addition

1. Draw a rough sketch of the situation.
2. Choose a scale and include a reference direction.
3. Choose any of the vectors and draw it as an arrow in the correct direction and of the correct length – remember to put an arrowhead on the end to denote its direction.
4. Take the next vector and draw it as an arrow starting from the arrowhead of the first vector in the correct direction and of the correct length.
5. Continue until you have drawn each vector – each time starting from the head of the previous vector. In this way, the vectors to be added are drawn one after the other head-to-tail.
6. The resultant is then the vector drawn from the tail of the first vector to the head of the last. Its magnitude can be determined from the length of its arrow using the scale. Its direction too can be determined from the scale diagram.

Exercise 13.2: Head-to-Tail Addition I A ship leaves harbour H and sails 6 km north to port A. From here the ship travels 12 km east to port B, before sailing 5,5 km south-west to port C. Determine the ship's resultant displacement using the head-to-tail technique of vector addition.

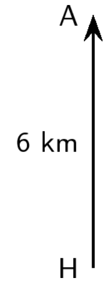
Solution to Exercise

Step 1. Its easy to understand the problem if we first draw a quick sketch. The rough sketch should include all of the information given in the problem. All of the magnitudes of the displacements are shown and a compass has been included as a reference direction. In a rough sketch one is interested in the approximate shape of the vector diagram.

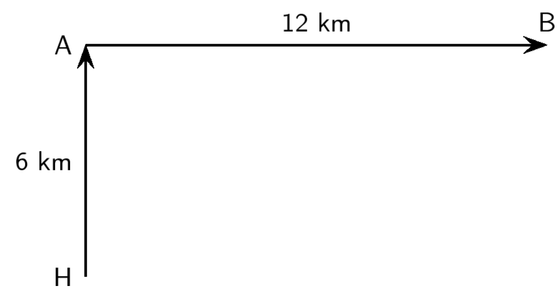


Step 2. The choice of scale depends on the actual question – you should choose a scale such that your vector diagram fits the page.
It is clear from the rough sketch that choosing a scale where 1 cm represents 2 km (scale: 1 cm = 2 km) would be a good choice in this problem. The diagram will then take up a good fraction of an A4 page. We now start the accurate construction.

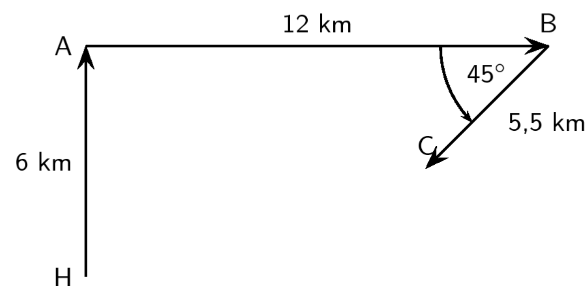
Step 3. Starting at the harbour H we draw the first vector 3 cm long in the direction north.



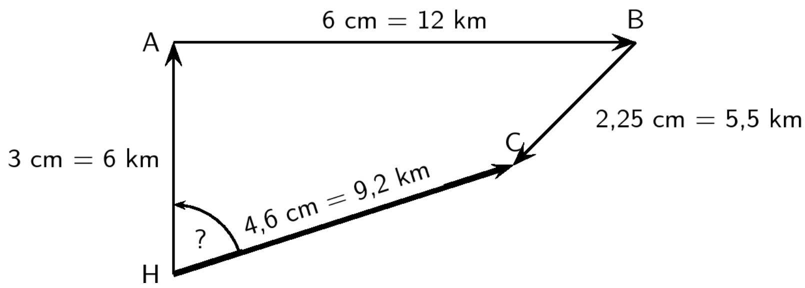
Step 4. Since the ship is now at port A we draw the second vector 6 cm long starting from point A in the direction east.



Step 5. Since the ship is now at port B we draw the third vector 2,25 cm long starting from this point in the direction south-west. A protractor is required to measure the angle of 45°.



Step 6. As a final step we draw the resultant displacement from the starting point (the harbour H) to the end point (port C). We use a ruler to measure the length of this arrow and a protractor to determine its direction.



Step 7. We now use the scale to convert the length of the resultant in the scale diagram to the actual displacement in the problem. Since we have chosen a scale of 1 cm = 2 km in this problem the resultant has a magnitude of 9,2 km. The direction can be specified in terms of the angle measured either as 072,3° east of north or on a bearing of 072,3°.

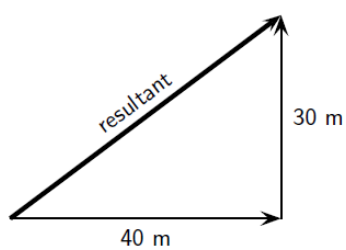
Step 8. The resultant displacement of the ship is 9,2 km on a bearing of 072,3°.

Exercise 13.3: Head-to-Tail Graphical Addition II A man walks 40 m East, then 30 m North.

1. What was the total distance he walked?
2. What is his resultant displacement?

Solution to Exercise

Step 1.

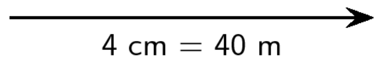


Step 2. In the first part of his journey he traveled 40 m and in the second part he traveled 30 m. This gives us a total distance traveled of 40 m + 30 m = 70 m.

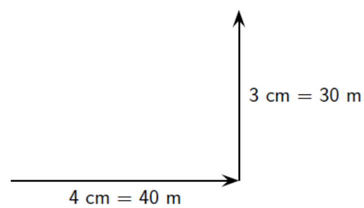
Step 3. The man's resultant displacement is the **vector** from where he started to where he ended. It is the vector sum of his two separate displacements. We will use the head-to-tail method of accurate construction to find this vector.

Step 4. A scale of 1 cm represents 10 m (1 cm = 10 m) is a good choice here. Now we can begin the process of construction.

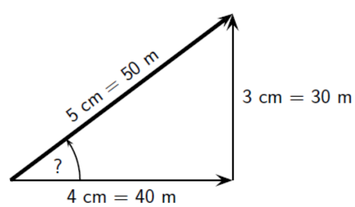
Step 5. We draw the first displacement as an arrow 4 cm long in an eastwards direction.



Step 6. Starting from the head of the first vector we draw the second vector as an arrow 3 cm long in a northerly direction.



Step 7. Now we connect the starting point to the end point and measure the length and direction of this arrow (the resultant).



Step 8. To find the direction you measure the angle between the resultant and the 40 m vector. You should get about 37°.

Step 9. Finally we use the scale to convert the length of the resultant in the scale diagram to the actual magnitude of the resultant displacement. According to the chosen scale 1 cm = 10 m. Therefore 5 cm represents 50 m. The resultant displacement is then 50 m 37° north of east.

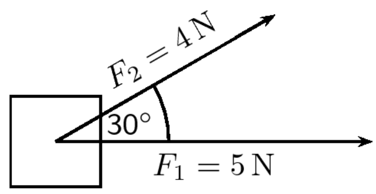
The Parallelogram Method

The *parallelogram method* is another graphical technique of finding the resultant of two vectors.

Method: The Parallelogram Method

1. Make a rough sketch of the vector diagram.
2. Choose a scale and a reference direction.
3. Choose either of the vectors to be added and draw it as an arrow of the correct length in the correct direction.
4. Draw the second vector as an arrow of the correct length in the correct direction from the tail of the first vector.
5. Complete the parallelogram formed by these two vectors.
6. The resultant is then the diagonal of the parallelogram. The magnitude can be determined from the length of its arrow using the scale. The direction too can be determined from the scale diagram.

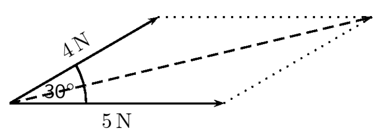
Exercise 13.4: Parallelogram Method of Vector Addition I A force of $F_1 = 5\mathrm{N}$ is applied to a block in a horizontal direction. A second force $F_2 = 4\mathrm{N}$ is applied to the object at an angle of 30° above the horizontal.



Determine the resultant force acting on the block using the parallelogram method of accurate construction.

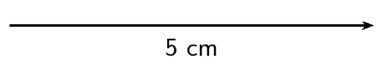
Solution to Exercise

Step 1.

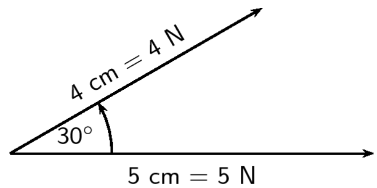


Step 2. In this problem a scale of 1 cm = 1 N would be appropriate, since then the vector diagram would take up a reasonable fraction of the page. We can now begin the accurate scale diagram.

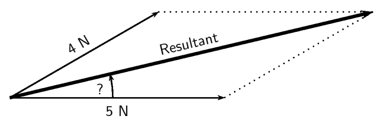
Step 3. Let us draw F_1 first. According to the scale it has length 5 cm.



Step 4. Next we draw F_2. According to the scale it has length 4 cm. We make use of a protractor to draw this vector at 30° to the horizontal.



Step 5. Next we complete the parallelogram and draw the diagonal.



The resultant has a measured length of 8,7 cm.

Step 6. We use a protractor to measure the angle between the horizontal and the resultant. We get 13,3°.

Step 7. Finally we use the scale to convert the measured length into the actual magnitude. Since 1 cm = 1 N, 8,7 cm represents 8,7 N. Therefore the resultant force is 8,7 N at 13,3° above the horizontal.

The parallelogram method is restricted to the addition of just two vectors. However, it is arguably the most intuitive way of adding two forces acting on a point.

1.100.2 Algebraic Addition and Subtraction of Vectors

Vectors in a Straight Line

Whenever you are faced with adding vectors acting in a straight line (i.e. some directed left and some right, or some acting up and others down) you can use a very simple algebraic technique:

Method: Addition/Subtraction of Vectors in a Straight Line

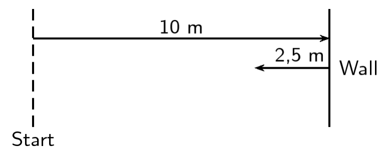
1. Choose a positive direction. As an example, for situations involving displacements in the directions west and east, you might choose west as your positive direction. In that case, displacements east are negative.
2. Next simply add (or subtract) the magnitude of the vectors using the appropriate signs.
3. As a final step the direction of the resultant should be included in words (positive answers are in the positive direction, while negative resultants are in the negative direction).

Let us consider a few examples.

Exercise 13.5: Adding vectors algebraically I A tennis ball is rolled towards a wall which is 10 m away from the ball. If after striking the wall the ball rolls a further 2,5 m along the ground away from the wall, calculate algebraically the ball's resultant displacement.

Solution to Exercise

Step 1.



Step 2. We know that the resultant displacement of the ball ($\vec{x}_R$) is equal to the sum of the ball's separate displacements ($\vec{x}_1$ and $\vec{x}_2$):

$$\vec{x}_R = \vec{x}_1 + \vec{x}_2 \tag{13.3}$$

Since the motion of the ball is in a straight line (i.e. the ball moves towards and away from the wall), we can use the method of algebraic addition just explained.

Step 3. Let's choose the **positive** direction to be towards the wall. This means that the **negative** direction is away from the wall.

Step 4. With right positive:

$$\begin{aligned} \vec{x}_1 &= +10,0 m \cdot s^{-1} \\ \vec{x}_2 &= -2,5 m \cdot s^{-1} \end{aligned} \tag{13.4}$$

Step 5. Next we simply add the two displacements to give the resultant:

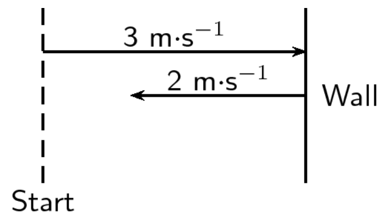
$$\begin{aligned} \vec{x}_R &= \left(+10 m \cdot s^{-1}\right) + \left(-2,5 m \cdot s^{-1}\right) \\ &= (+7,5)\, m \cdot s^{-1} \end{aligned} \tag{13.5}$$

Step 6. Finally, in this case towards the wall is the positive direction, so: $\vec{x}_R$ = 7,5 m towards the wall.

Exercise 13.6: Subtracting vectors algebraically I Suppose that a tennis ball is thrown horizontally towards a wall at an initial velocity of 3 m·s^{-1}to the right. After striking the wall, the ball returns to the thrower at 2 m·s^{-1}. Determine the change in velocity of the ball.

Solution to Exercise

Step 1. A quick sketch will help us understand the problem.



Step 2. Remember that velocity is a vector. The change in the velocity of the ball is equal to the difference between the ball's initial and final velocities:

$$\Delta \vec{v} = \vec{v}_f - \vec{v}_i \tag{13.6}$$

Since the ball moves along a straight line (i.e. left and right), we can use the algebraic technique of vector subtraction just discussed.

Step 3. Choose the **positive** direction to be towards the wall. This means that the **negative** direction is away from the wall.

Step 4.

$$\begin{aligned} \vec{v}_i &= +3m \cdot s^{-1} \\ \vec{v}_f &= -2m \cdot s^{-1} \end{aligned} \tag{13.7}$$

Step 5. Thus, the change in velocity of the ball is:

$$\begin{aligned} \Delta \vec{v} &= \left(-2m \cdot s^{-1}\right) - \left(+3m \cdot s^{-1}\right) \\ &= (-5)\, m \cdot s^{-1} \end{aligned} \tag{13.8}$$

Step 6. Remember that in this case towards the wall means a positive velocity, so away from the wall means a negative velocity: $\Delta \vec{v} = 5m \cdot s^{-1}$ away from the wall.

Resultant Vectors

1. Harold walks to school by walking 600 m Northeast and then 500 m N 40° W. Determine his resultant displacement by using accurate scale drawings.

2. A dove flies from her nest, looking for food for her chick. She flies at a velocity of 2 m·s^{-1} on a bearing of 135° and then at a velocity of 1,2 m·s^{-1} on a bearing of 230°. Calculate her resultant velocity by using accurate scale drawings.

3. A squash ball is dropped to the floor with an initial velocity of 2,5 m·s^{-1}. It rebounds (comes back up) with a velocity of 0,5 m·s^{-1}.
   - **a.** What is the change in velocity of the squash ball?
   - **b.** What is the resultant velocity of the squash ball?

Remember that the technique of addition and subtraction just discussed can only be applied to vectors acting along a straight line. When vectors are not in a straight line, i.e. at an angle to each other, the following method can be used:

Find the answers with the shortcodes:

(1.) l2b (2.) l2j (3.) l4F

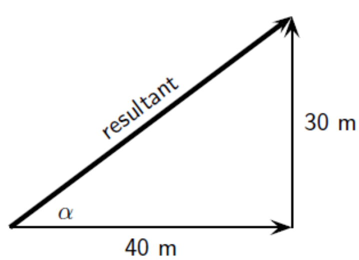
A More General Algebraic technique

Simple geometric and trigonometric techniques can be used to find resultant vectors.

Exercise 13.7: An Algebraic Solution I A man walks 40 m East, then 30 m North. Calculate the man's resultant displacement.

Solution to Exercise

Step 1. As before, the rough sketch looks as follows:



Step 2. Note that the triangle formed by his separate displacement vectors and his resultant displacement vector is a right-angle triangle. We can thus use the Theorem of Pythagoras to determine the length of the resultant. Let x_R represent the length of the resultant vector. Then:

$$\begin{aligned} x_R^2 &= \left(40m \cdot s^{-1}\right)^2 + \left(30m \cdot s^{-1}\right)^2 \\ x_R^2 &= 2\ 500m \cdot s^{-1^2} \\ x_R &= 50m \cdot s^{-1} \end{aligned} \tag{13.9}$$

Step 3. Now we have the length of the resultant displacement vector but not yet its direction. To determine its direction we calculate the angle α between the resultant displacement vector and East, by using simple trigonometry:

$$\begin{aligned} tan\alpha &= \frac{\text{oppositeside}}{\text{adjacentside}} \\ tan\alpha &= \frac{30}{40} \\ \alpha &= tan^{-1}(0,75) \\ \alpha &= 36,9^{\circ} \end{aligned} \qquad (13.10)$$

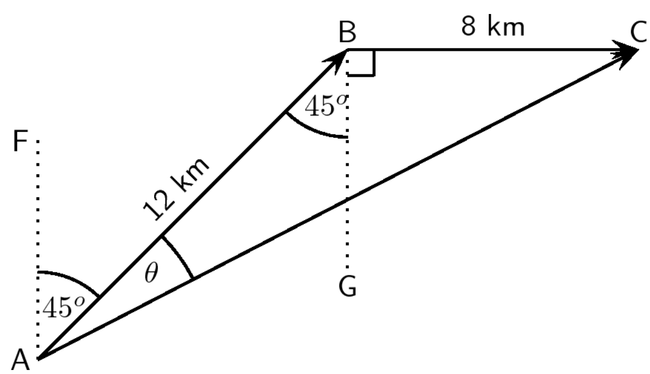
Step 4. The resultant displacement is then 50 m at 36,9° North of East. This is exactly the same answer we arrived at after drawing a scale diagram!

In the previous example we were able to use simple trigonometry to calculate the resultant displacement. This was possible since the directions of motion were perpendicular (north and east). Algebraic techniques, however, are not limited to cases where the vectors to be combined are along the same straight line or at right angles to one another. The following example illustrates this.

Exercise 13.8: An Algebraic Solution II A man walks from point A to point B which is 12 km away on a bearing of 45°. From point B the man walks a further 8 km east to point C. Calculate the resultant displacement.

Solution to Exercise

Step 1.



$B\hat{A}F = 45°$ since the man walks initially on a bearing of 45°.
Then, $A\hat{B}G = B\hat{A}F = 45°$ (parallel lines, alternate angles). Both of these angles are included in the rough sketch.

Step 2. The resultant is the vector AC. Since we know both the lengths of AB and BC and the included angle $A\hat{B}C$, we can use the cosine rule:

$$\begin{aligned} AC^2 &= AB^2 + BC^2 - 2 \cdot AB \cdot BC cos\left(A\hat{B}C\right) \\ &= (12)^2 + (8)^2 - 2 \cdot (12)(8) cos(135°) \\ &= 343,8 \\ AC &= 18,5\text{km} \end{aligned} \tag{13.11}$$

Step 3. Next we use the sine rule to determine the angle θ:

$$\begin{aligned} \frac{sin\theta}{8} &= \frac{sin135°}{18,5} \\ sin\theta &= \frac{8 \times sin135°}{18,5} \\ \theta &= sin^{-1}(0,3058) \\ \theta &= 17,8° \end{aligned} \tag{13.12}$$

To find $F\hat{A}C$, we add 45°. Thus, $F\hat{A}C = 62,8°$.

Step 4. The resultant displacement is therefore 18,5 km on a bearing of 062,8°.

More Resultant Vectors

1. A frog is trying to cross a river. It swims at 3 m·s^{-1}in a northerly direction towards the opposite bank. The water is flowing in a westerly direction at 5 m·s^{-1}. Find the frog's resultant velocity by using appropriate calculations. Include a rough sketch of the situation in your answer.
2. Sandra walks to the shop by walking 500 m Northwest and then 400 m N 30° E. Determine her resultant displacement by doing appropriate calculations.

Find the answers with the shortcodes:

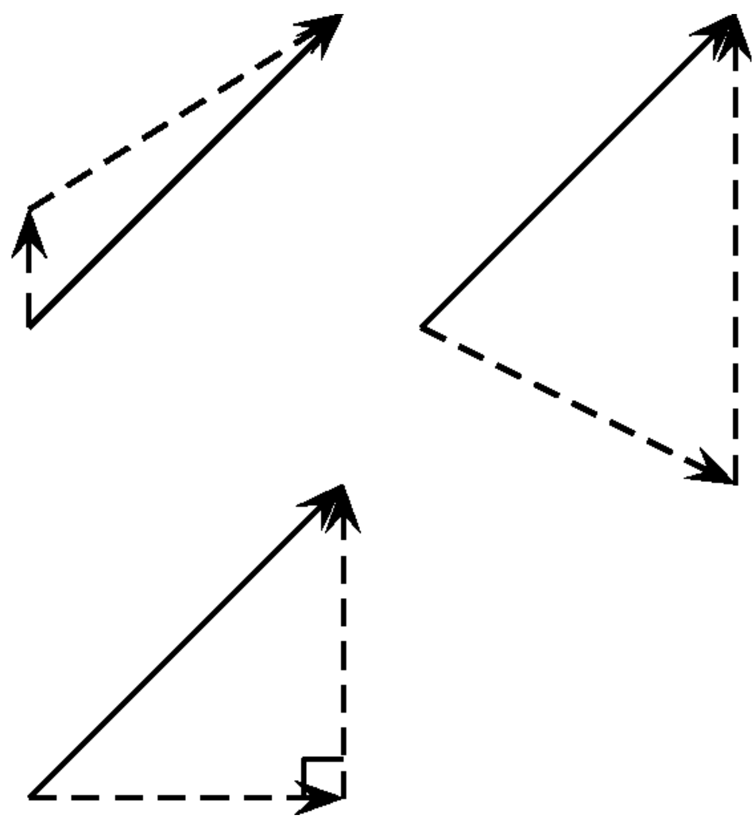
(1.) l4L (2.) l4M

1.101 Components of Vectors

(section shortcode: P10007)

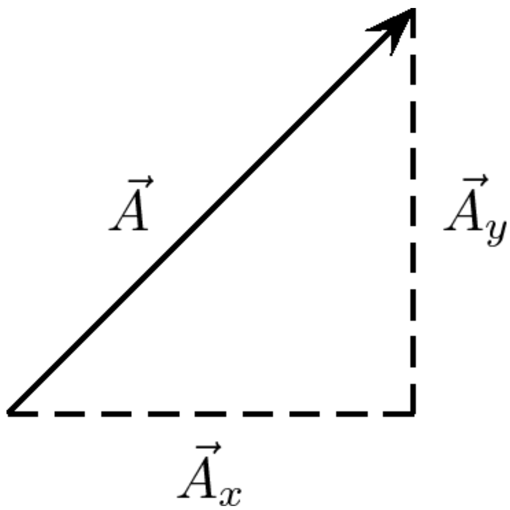
In the discussion of vector addition we saw that a number of vectors acting together can be combined to give a single vector (the resultant). In much the same way a single vector can be broken down into a number of vectors which when added give that original vector. These vectors which sum to the original are called **components** of the original vector. The process of breaking a vector into its components is called **resolving into components**.

While summing a given set of vectors gives just one answer (the resultant), a single vector can be resolved into infinitely many sets of components. In the diagrams below the same black vector is resolved into different pairs of components. These components are shown as dashed lines. When added together the dashed vectors give the original black vector (i.e. the original vector is the resultant of its components).



In practice it is most useful to resolve a vector into components which are at right angles to one another, usually horizontal and vertical.

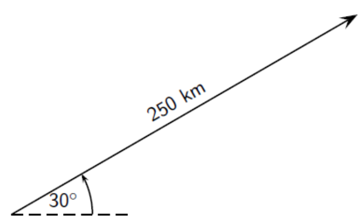
Any vector can be resolved into a horizontal and a vertical component. If $\vec{A}$ is a vector, then the horizontal component of $\vec{A}$ is $\vec{A}_x$ and the vertical component is $\vec{A}_y$.



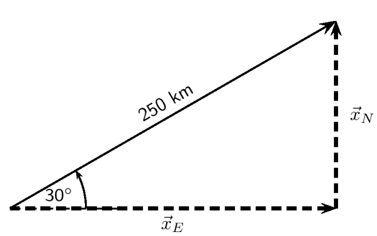
Exercise 13.9: Resolving a vector into components A motorist undergoes a displacement of 250 km in a direction 30° north of east. Resolve this displacement into components in the directions north ($\vec{x}_N$) and east ($\vec{x}_E$).

Solution to Exercise

Step 1.



Step 2. Next we resolve the displacement into its components north and east. Since these directions are perpendicular to one another, the components form a right-angled triangle with the original displacement as its hypotenuse.



Notice how the two components acting together give the original vector as their resultant.

Step 3. Now we can use trigonometry to calculate the magnitudes of the components of the original displacement:

$$\begin{aligned} x_N &= (250)\,(sin30°) \\ &= 125\ \text{km} \end{aligned} \tag{13.13}$$

and

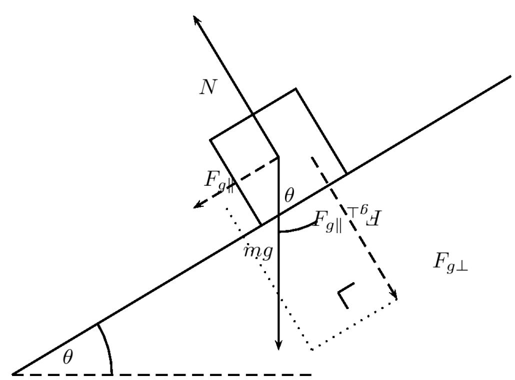
$$\begin{aligned} x_E &= (250)\,(cos30°) \\ &= 216,5\ \text{km} \end{aligned} \tag{13.14}$$

Remember x_N and x_E are the magnitudes of the components – they are in the directions north and east respectively.

1.101.1 Block on an incline

As a further example of components let us consider a block of mass m placed on a frictionless surface inclined at some angle θ to the horizontal. The block will obviously slide down the incline, but what causes this motion?

The forces acting on the block are its weight mg and the normal force N exerted by the surface on the object. These two forces are shown in the diagram below.



Now the object's weight can be resolved into components parallel and perpendicular to the inclined surface. These components are shown as dashed arrows in the diagram above and are at right angles to each other. The components have been drawn acting from the same point. Applying the parallelogram method, the two components of the block's weight sum to the weight vector.

To find the components in terms of the weight we can use trigonometry:

$$\begin{aligned} F_{g\|} &= mgsin\theta \\ F_{g\perp} &= mgcos\theta \end{aligned} \qquad (13.15)$$

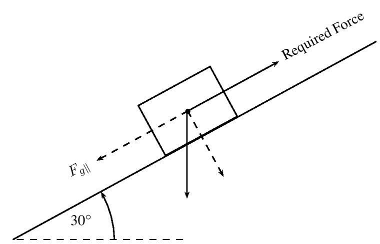
The component of the weight perpendicular to the slope $F_{g\perp}$ exactly balances the normal force N exerted by the surface. The parallel component, however, $F_{g\|}$ is unbalanced and causes the block to slide down the slope.

1.101.2 Worked example

Exercise 13.10: Block on an incline plane Determine the force needed to keep a 10 kg block from sliding down a frictionless slope. The slope makes an angle of 30° with the horizontal.

Solution to Exercise

Step 1.



Step 2. The force that will keep the block from sliding is equal to the parallel component of the weight, but its direction is up the slope.

$$\begin{aligned} F_{g\|} &= mgsin\theta \\ &= (10)\,(9,8)\,(sin30^\circ) \\ &= 49\mathrm{N} \end{aligned} \qquad (13.16)$$

Step 3. The force is 49 N up the slope.

1.101.3 Vector addition using components

Components can also be used to find the resultant of vectors. This technique can be applied to both graphical and algebraic methods of finding the resultant. The method is simple: make a rough sketch of the problem, find the horizontal and vertical components of each vector, find the sum of all horizontal components and the sum of all the vertical components and then use them to find the resultant.

Consider the two vectors, $\overrightarrow{A}$ and $\overrightarrow{B}$, in Figure 13.64, together with their resultant, $\overrightarrow{R}$.

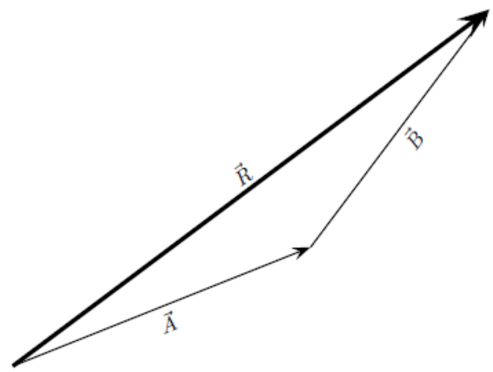


Figure 13.64: An example of two vectors being added to give a resultant

Each vector in Figure 13.64 can be broken down into one component in the x-direction (horizontal) and one in the y-direction (vertical). These components are two vectors which when added give you the original vector as the resultant. This is shown in Figure 13.65 where we can see that:

$$\begin{array}{rcl} \overrightarrow{A} & = & \overrightarrow{A}_x + \overrightarrow{A}_y \\ \overrightarrow{B} & = & \overrightarrow{B}_x + \overrightarrow{B}_y \\ \overrightarrow{R} & = & \overrightarrow{R}_x + \overrightarrow{R}_y \end{array} \tag{13.17}$$

$$\begin{array}{rcl} \mathrm{But}, \overrightarrow{\mathrm{R}}_{\mathrm{x}} & = & \overrightarrow{A}_x + \overrightarrow{B}_x \\ \mathrm{and} \overrightarrow{\mathrm{R}}_{\mathrm{y}} & = & \overrightarrow{A}_y + \overrightarrow{B}_y \end{array} \tag{13.18}$$

In summary, addition of the x components of the two original vectors gives the x component of the resultant. The same applies to the y components. So if we just added all the components together we would get the same answer! This is another important property of vectors.

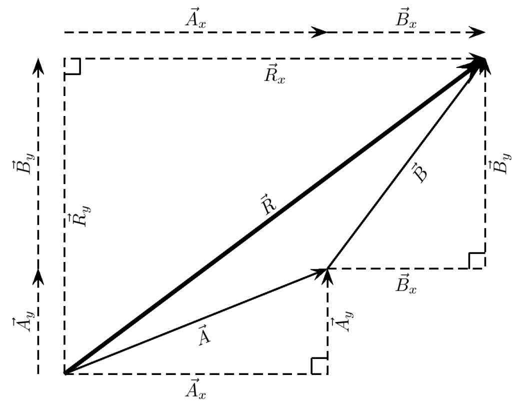


Figure 13.65: Adding vectors using components.

Exercise 13.11: Adding Vectors Using Components If in Figure 13.65, $\overrightarrow{A}= 5,385m \cdot s^{-1}$ at an angle of 21.8° to the horizontal and $\overrightarrow{B}= 5m \cdot s^{-1}$ at an angle of 53,13° to the horizontal, find $\overrightarrow{R}$.

Solution to Exercise

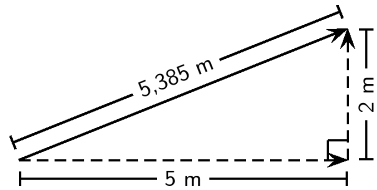
Step 1. The first thing we must realise is that the order that we add the vectors does not matter. Therefore, we can work through the vectors to be added in any order.

Step 2. We find the components of $\overrightarrow{A}$ by using known trigonometric ratios. First we find the magnitude of the vertical component, A_y:

$$\begin{array}{rcl} sin\theta & = & \frac{A_y}{A} \\ sin21,8^\circ & = & \frac{A_y}{5,385} \\ A_y & = & (5,385)\,(sin21,8^\circ) \\ & = & 2m \cdot s^{-1} \end{array} \tag{13.19}$$

Secondly we find the magnitude of the horizontal component, A_x:

$$\begin{array}{rcl} cos\theta & = & \frac{A_x}{A} \\ cos21.8^\circ & = & \frac{A_x}{5,385} \\ A_x & = & (5,385)\,(cos21,8^\circ) \\ & = & 5m \cdot s^{-1} \end{array} \tag{13.20}$$



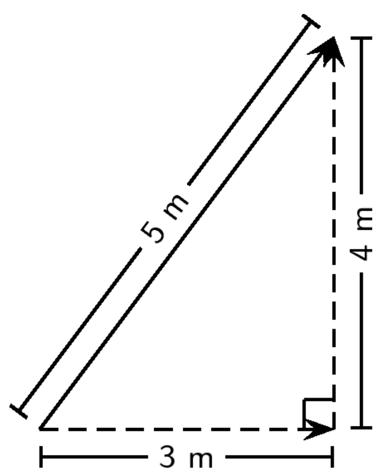
The components give the sides of the right angle triangle, for which the original vector, $\vec{A}$, is the hypotenuse.

Step 3. We find the components of $\vec{B}$ by using known trigonometric ratios. First we find the magnitude of the vertical component, B_y:

$$\begin{aligned} sin\theta &= \frac{B_y}{B} \\ sin53,13^\circ &= \frac{B_y}{5} \\ B_y &= (5)\,(sin53,13^\circ) \\ &= 4m \cdot s^{-1} \end{aligned} \tag{13.21}$$

Secondly we find the magnitude of the horizontal component, B_x:

$$\begin{aligned} cos\theta &= \frac{B_x}{B} \\ cos21,8^\circ &= \frac{B_x}{5,385} \\ B_x &= (5,385)\,(cos53,13^\circ) \\ &= 5m \cdot s^{-1} \end{aligned} \tag{13.22}$$



Step 4. Now we have all the components. If we add all the horizontal components then we will have the x-component of the resultant vector, $\vec{R}_x$. Similarly, we add all the vertical components then we will have

the y-component of the resultant vector, $\vec{R}_y$.

$$\begin{aligned} R_x &= A_x + B_x \\ &= 5m \cdot s^{-1} + 3m \cdot s^{-1} \\ &= 8m \cdot s^{-1} \end{aligned} \tag{13.23}$$

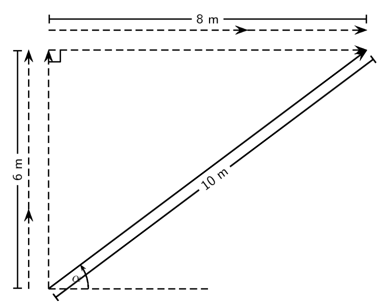
Therefore, $\vec{R}_x$ is 8 m to the right.

$$\begin{aligned} R_y &= A_y + B_y \\ &= 2m \cdot s^{-1} + 4m \cdot s^{-1} \\ &= 6m \cdot s^{-1} \end{aligned} \tag{13.24}$$

Therefore, $\vec{R}_y$ is 6 m up.

Step 5. Now that we have the components of the resultant, we can use the Theorem of Pythagoras to determine the magnitude of the resultant, R.

$$\begin{aligned} R^2 &= (R_x)^2 + (R_y)^2 \\ R^2 &= (6)^2 + (8)^2 \\ R^2 &= 100 \\ \therefore R &= 10m \cdot s^{-1} \end{aligned} \tag{13.25}$$



The magnitude of the resultant, R is 10 m. So all we have to do is calculate its direction. We can specify the direction as the angle the vectors makes with a known direction. To do this you only need to visualise the vector as starting at the origin of a coordinate system. We have drawn this explicitly below and the angle we will calculate is labeled α.

Using our known trigonometric ratios we can calculate the value of α;

$$\begin{aligned} tan\alpha &= \frac{6m \cdot s^{-1}}{8m \cdot s^{-1}} \\ \alpha &= tan^{-1}\frac{6m \cdot s^{-1}}{8m \cdot s^{-1}} \\ \alpha &= 36,8^\circ. \end{aligned} \tag{13.26}$$

Step 6. $\vec{R}$ is 10 m at an angle of $36,8°$ to the positive x-axis.

Adding and Subtracting Components of Vectors

1. Harold walks to school by walking 600 m Northeast and then 500 m N 40° W. Determine his resultant displacement by means of addition of components of vectors.

2. A dove flies from her nest, looking for food for her chick. She flies at a velocity of 2 m·s^{-1} on a bearing of 135^o in a wind with a velocity of 1,2 m·s^{-1} on a bearing of 230^o. Calculate her resultant velocity by adding the horizontal and vertical components of vectors.

Find the answers with the shortcodes:

(1.) l4e (2.) l2Z

Vector Multiplication

Vectors are special, they are more than just numbers. This means that multiplying vectors is not necessarily the same as just multiplying their magnitudes. There are two different types of multiplication defined for vectors. You can find the dot product of two vectors or the cross product.

The *dot* product is most similar to regular multiplication between scalars. To take the dot product of two vectors, you just multiply their magnitudes to get out a scalar answer. The mathematical definition of the dot product is:

$$\vec{a} \bullet \vec{b} = |\vec{a}| \cdot |\vec{b}| cos\theta \tag{13.27}$$

Take two vectors $\vec{a}$ and $\vec{b}$:

You can draw in the component of $\vec{b}$ that is parallel to $\vec{a}$:

In this way we can arrive at the definition of the dot product. You find how much of $\vec{b}$ is lined up with $\vec{a}$ by finding the component of $\vec{b}$ parallel to $\vec{a}$. Then multiply the magnitude of that component, $|\vec{b}|cos\theta$, with the magnitude of $\vec{a}$ to get a scalar.

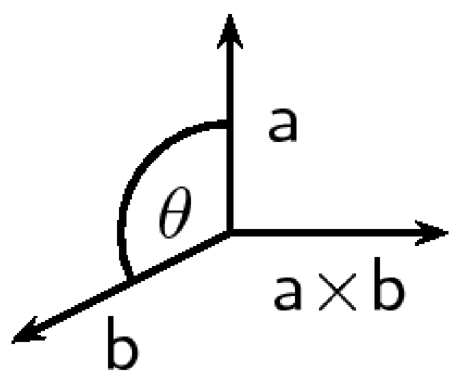
The second type of multiplication, the cross product, is more subtle and uses the directions of the vectors in a more complicated way. The cross product of two vectors, $\vec{a}$ and $\vec{b}$, is written $\vec{a} \times \vec{b}$ and the result of this operation on $\vec{a}$ and $\vec{b}$ is another vector. The magnitude of the cross product of these two vectors is:

$$|\vec{a} \times \vec{b}| = |\vec{a}||\vec{b}|sin\theta \tag{13.28}$$

We still need to find the direction of $\vec{a} \times \vec{b}$. We do this by applying the *right hand rule*.

Method: Right Hand Rule

1. Using your right hand:
2. Point your index finger in the direction of $\vec{a}$.
3. Point the middle finger in the direction of $\vec{b}$.
4. Your thumb will show the direction of $\vec{a} \times \vec{b}$.

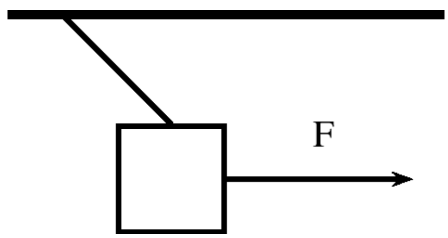


1.101.4 Summary

1. A scalar is a physical quantity with magnitude only.
2. A vector is a physical quantity with magnitude and direction.
3. Vectors may be represented as arrows where the length of the arrow indicates the magnitude and the arrowhead indicates the direction of the vector.
4. The direction of a vector can be indicated by referring to another vector or a fixed point (eg. 30° from the river bank); using a compass (eg. N 30° W); or bearing (eg. 053°).
5. Vectors can be added using the head-to-tail method, the parallelogram method or the component method.
6. The resultant of a number of vectors is the single vector whose effect is the same as the individual vectors acting together.

1.101.5 End of chapter exercises: Vectors

1. An object is suspended by means of a light string. The sketch shows a horizontal force F which pulls the object from the vertical position until it reaches an equilibrium position as shown. Which one of the following vector diagrams best represents all the forces acting on the object?



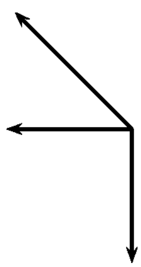
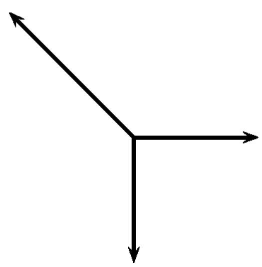
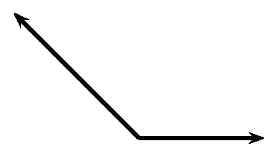
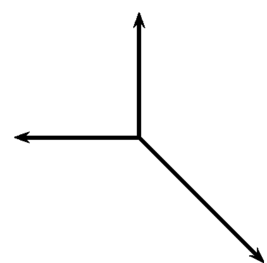
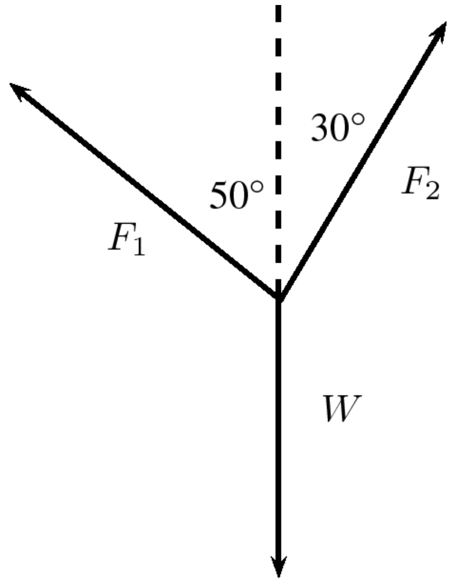
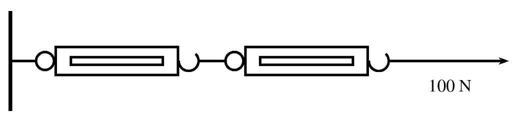
A	B	C	D
			

Table 13.9

2. A load of weight W is suspended from two strings. F_1 and F_2 are the forces exerted by the strings on the load in the directions show in the figure above. Which one of the following equations is valid for this situation?
   a. $W = F_1^2 + F_2^2$
   b. $F_1 sin50° = F_2 sin30°$
   c. $F_1 cos50° = F_2 cos30°$
   d. $W = F_1 + F_2$



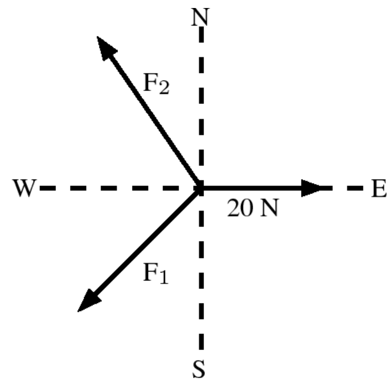
3. Two spring balances P and Q are connected by means of a piece of string to a wall as shown. A horizontal force of 100 N is exerted on spring balance Q. What will be the readings on spring balances P and Q?



	P	Q
A	100 N	0 N
B	25 N	75 N
C	50 N	50 N
D	100 N	100 N

Table 13.10

4. A point is acted on by two forces in equilibrium. The forces
   - **a.** have equal magnitudes and directions.
   - **b.** have equal magnitudes but opposite directions.
   - **c.** act perpendicular to each other.
   - **d.** act in the same direction.
5. A point in equilibrium is acted on by three forces. Force F_1 has components 15 N due south and 13 N due west. What are the components of force F_2?
   - **a.** 13 N due north and 20 due west
   - **b.** 13 N due north and 13 N due west
   - **c.** 15 N due north and 7 N due west
   - **d.** 15 N due north and 13 N due east



6. Which of the following contains two vectors and a scalar?
   a. distance, acceleration, speed
   b. displacement, velocity, acceleration
   c. distance, mass, speed
   d. displacement, speed, velocity
7. Two vectors act on the same point. What should the angle between them be so that a maximum resultant is obtained?
   a. 0°
   b. 90°
   c. 180°
   d. cannot tell
8. Two forces, 4 N and 11 N, act on a point. Which one of the following cannot be the magnitude of a resultant?
   a. 4 N
   b. 7 N
   c. 11 N
   d. 15 N

Find the answers with the shortcodes:

(1.) l2B (2.) l2K (3.) l2k (4.) l20 (5.) l28 (6.) l29
(7.) l2X (8.) l2I

1.101.6 End of chapter exercises: Vectors - Long questions

1. A helicopter flies due east with an air speed of 150 km.h^{-1}. It flies through an air current which moves at 200 km.h^{-1} north. Given this information, answer the following questions:
   a. In which direction does the helicopter fly?
   b. What is the ground speed of the helicopter?
   c. Calculate the ground distance covered in 40 minutes by the helicopter.
2. A plane must fly 70 km due north. A cross wind is blowing to the west at 30 km.h^{-1}. In which direction must the pilot steer if the plane flies at a speed of 200 km.h^{-1} in windless conditions?

3. A stream that is 280 m wide flows along its banks with a velocity of 1.80m.s^{-1}. A raft can travel at a speed of 2.50 m.s^{-1} across the stream. Answer the following questions:

a. What is the shortest time in which the raft can cross the stream?
b. How far does the raft drift downstream in that time?
c. In what direction must the raft be steered against the current so that it crosses the stream perpendicular to its banks?
d. How long does it take to cross the stream in part c?

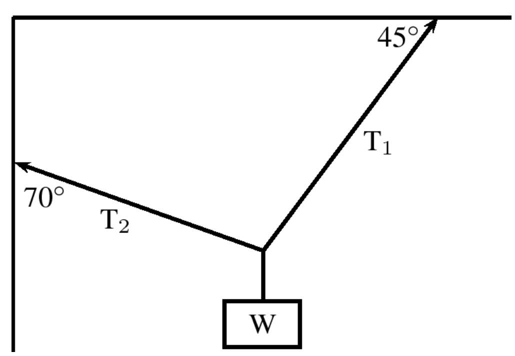
4. A helicopter is flying from place X to place Y. Y is 1000 km away in a direction $50°$ east of north and the pilot wishes to reach it in two hours. There is a wind of speed 150 km.h^{-1} blowing from the northwest. Find, by accurate construction and measurement (with a scale of $1 \text{ cm} = 50 \text{ km.h}^{-1}$), the

a. the direction in which the helicopter must fly, and
b. the magnitude of the velocity required for it to reach its destination on time.

5. An aeroplane is flying towards a destination 300 km due south from its present position. There is a wind blowing from the north east at 120 km.h^{-1}. The aeroplane needs to reach its destination in 30 minutes. Find, by accurate construction and measurement (with a scale of $1 \text{ cm} = 30 \text{ km.s}^{-1}$), or otherwise,

a. the direction in which the aeroplane must fly and
b. the speed which the aeroplane must maintain in order to reach the destination on time.
c. Confirm your answers in the previous 2 subquestions with calculations.

6. An object of weight W is supported by two cables attached to the ceiling and wall as shown. The tensions in the two cables are T_1 and T_2 respectively. Tension $T_1 = 1200$ N. Determine the tension T_2 and weight W of the object by accurate construction and measurement or by calculation.



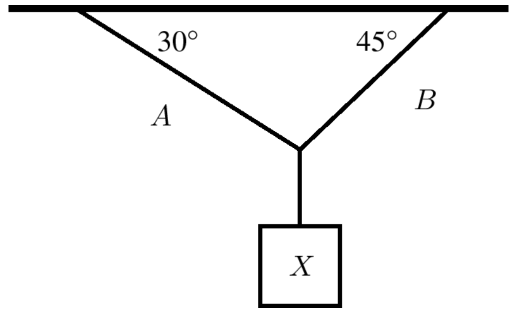
7. In a map-work exercise, hikers are required to walk from a tree marked A on the map to another tree marked B which lies 2,0 km due East of A. The hikers then walk in a straight line to a waterfall in position C which has components measured from B of 1,0 km E and 4,0 km N.

a. Distinguish between quantities that are described as being *vector* and *scalar*.
b. Draw a labelled displacement-vector diagram (not necessarily to scale) of the hikers' complete journey.

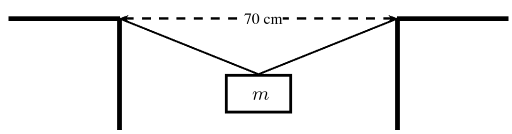
c. What is the total distance walked by the hikers from their starting point at A to the waterfall C?
d. What are the magnitude and bearing, to the nearest degree, of the displacement of the hikers from their starting point to the waterfall?

8. An object X is supported by two strings, A and B, attached to the ceiling as shown in the sketch. Each of these strings can withstand a maximum force of 700 N. The weight of X is increased gradually.

a. Draw a rough sketch of the triangle of forces, and use it to explain which string will break first.
b. Determine the maximum weight of X which can be supported.



9. A rope is tied at two points which are 70 cm apart from each other, on the same horizontal line. The total length of rope is 1 m, and the maximum tension it can withstand in any part is 1000 N. Find the largest mass (m), in kg, that can be carried at the midpoint of the rope, without breaking the rope. Include a vector diagram in your answer.



Find the answers with the shortcodes:

(1.) l25 (2.) l2N (3.) l2R (4.) l2n (5.) l2E (6.) l2m
(7.) l2y (8.) l2V (9.) l2p

Mechanical Energy

1.102 Introduction

(section shortcode: P10008)

All objects have energy. The word energy comes from the Greek word energeia ($[U+03AD]\nu[U+03AD]\rho\gamma\epsilon\iota\alpha$), meaning activity or operation. Energy is closely linked to mass and cannot be created or destroyed. In this chapter we will consider potential and kinetic energy.

1.103 Potential Energy

(section shortcode: P10009)

The potential energy of an object is generally defined as the energy an object has because of its position relative to other objects that it interacts with. There are different kinds of potential energy such as gravitional potential energy, chemical potential energy, electrical potential energy, to name a few. In this section we will be looking at gravitational potential energy.

Definition: Potential energy
Potential energy is the energy an object has due to its position or state.

Gravitational potential energy is the energy of an object due to its position above the surface of the Earth. The symbol PE is used to refer to gravitational potential energy. You will often find that the words potential energy are used where *gravitational* potential energy is meant. We can define potential energy (or gravitational potential energy, if you like) as:

$$E_P = mgh \quad (14.1)$$

where E_P = potential energy measured in joules (J)

m = mass of the object (measured in kg)

g = gravitational acceleration ($9,8\,\mathrm{m}\cdot\mathrm{s}^{-2}$)

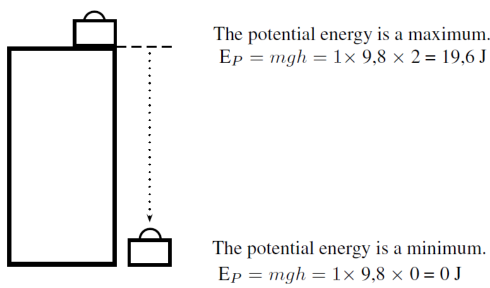
h = perpendicular height from the reference point (measured in m)

TIP: You may sometimes see potential energy written as PE. We will not use this notation in this book, but you may see it in other books.

A suitcase, with a mass of $1\,\mathrm{kg}$, is placed at the top of a $2\,\mathrm{m}$ high cupboard. By lifting the suitcase against the force of gravity, we give the suitcase potential energy. This potential energy can be calculated using (2.1).

If the suitcase falls off the cupboard, it will lose its potential energy. Halfway down the cupboard, the suitcase will have lost half its potential energy and will have only $9,8\,\mathrm{J}$ left. At the bottom of the cupboard the suitcase will have lost all it's potential energy and it's potential energy will be equal to zero.

Objects have maximum potential energy at a maximum height and will lose their potential energy as they fall.



Exercise 14.1: Gravitational potential energy A brick with a mass of $1\,\mathrm{kg}$ is lifted to the top of a $4\,m$ high roof. It slips off the roof and falls to the ground. Calculate the potential energy of the brick at the top of the roof and on the ground once it has fallen.

Solution to Exercise

Step 1.
- The mass of the brick is $m = 1\,\mathrm{kg}$
- The height lifted is $h = 4\,\mathrm{m}$

All quantities are in SI units.

Step 2.
- We are asked to find the gain in potential energy of the brick as it is lifted onto the roof.
- We also need to calculate the potential energy once the brick is on the ground again.

Step 3. Since the block is being lifted we are dealing with gravitational potential energy. To work out E_P, we need to know the mass of the object and the height lifted. As both of these are given, we just substitute them into the equation for E_P.

Step 4.

$$\begin{aligned} E_P &= mgh \\ &= (1)\,(9,8)\,(4) \\ &= 39,2\,\mathrm{J} \end{aligned} \qquad (14.2)$$

1.103.1 Gravitational Potential Energy

1. Describe the relationship between an object's gravitational potential energy and its:
   a. mass and
   b. height above a reference point.
2. A boy, of mass $30\,\mathrm{kg}$, climbs onto the roof of a garage. The roof is $2,5\,\mathrm{m}$ from the ground. He now jumps off the roof and lands on the ground.
   a. How much potential energy has the boy gained by climbing on the roof?
   b. The boy now jumps down. What is the potential energy of the boy when he is $1\,m$ from the ground?
   c. What is the potential energy of the boy when he lands on the ground?
3. A hiker walks up a mountain, $800\,\mathrm{m}$ above sea level, to spend the night at the top in the first overnight hut. The second day he walks to the second overnight hut, $500\,\mathrm{m}$ above sea level. The third day he returns to his starting point, $200\,\mathrm{m}$ above sea level.
   a. What is the potential energy of the hiker at the first hut (relative to sea level)?
   b. How much potential energy has the hiker lost during the second day?
   c. How much potential energy did the hiker have when he started his journey (relative to sea level)?
   d. How much potential energy did the hiker have at the end of his journey?

Find the answers with the shortcodes:

(1.) lxE (2.) lxm (3.) lxy

1.104 Kinetic Energy

(section shortcode: P10010)

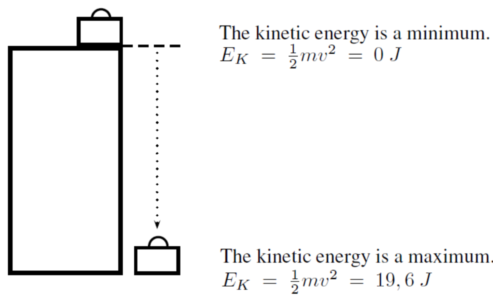
Definition: Kinetic Energy
Kinetic energy is the energy an object has due to its motion.

Kinetic energy is the energy an object has because of its motion. This means that any moving object has kinetic energy. The faster it moves, the more kinetic energy it has. Kinetic energy (E_K) is therefore dependent on the velocity of the object. The mass of the object also plays a role. A truck of $2000\,\mathrm{kg}$, moving at $100\,\mathrm{km}\cdot\mathrm{hr}^{-1}$, will have more kinetic energy than a car of $500\,\mathrm{kg}$, also moving at $100\,\mathrm{km}\cdot\mathrm{hr}^{-1}$. Kinetic energy is defined as:

NOTE: You may sometimes see kinetic energy written as KE. This is simply another way to write kinetic energy. We will not use this form in this book, but you may see it written like this in other books.

$$E_K = \frac{1}{2}mv^2 \tag{14.3}$$

Consider the $1\,\mathrm{kg}$ suitcase on the cupboard that was discussed earlier. When the suitcase falls, it will gain velocity (fall faster), until it reaches the ground with a maximum velocity. The suitcase will not have any kinetic energy when it is on top of the cupboard because it is not moving. Once it starts to fall it will gain kinetic energy, because it gains velocity. Its kinetic energy will increase until it is a maximum when the suitcase reaches the ground.



Exercise 14.2: Calculation of Kinetic Energy A $1\,\mathrm{kg}$ brick falls off a $4\,\mathrm{m}$ high roof. It reaches the ground with a velocity of $8{,}85\,\mathrm{m\cdot s^{-1}}$. What is the kinetic energy of the brick when it starts to fall and when it reaches the ground?

Solution to Exercise

Step 1.
- The mass of the rock $m = 1\,\mathrm{kg}$
- The velocity of the rock at the bottom $v_{\mathrm{bottom}} = 8{,}85\,\mathrm{m{\cdot}s^{-1}}$

These are both in the correct units so we do not have to worry about unit conversions.

Step 2. We are asked to find the kinetic energy of the brick at the top and the bottom. From the definition we know that to work out E_K, we need to know the mass and the velocity of the object and we are given both of these values.

Step 3. Since the brick is not moving at the top, its kinetic energy is zero.

Step 4.

$$\begin{aligned} E_K &= \tfrac{1}{2}mv^2 \\ &= \tfrac{1}{2}(1\ \mathrm{kg})\left(8{,}85\ \mathrm{m\cdot s^{-1}}\right)^2 \\ &= 39{,}2\,\mathrm{J} \end{aligned} \tag{14.4}$$

1.104.1 Checking units

According to the equation for kinetic energy, the unit should be $\mathrm{kg} \cdot \mathrm{m}^2 \cdot \mathrm{s}^{-2}$. We can prove that this unit is equal to the joule, the unit for energy.

$$\begin{array}{rcl} (\mathrm{kg})\left(\mathrm{m}\cdot\mathrm{s}^{-1}\right)^2 & = & \left(\mathrm{kg}\cdot\mathrm{m}\cdot\mathrm{s}^{-2}\right)\cdot\mathrm{m} \\ & = & \mathrm{N}\cdot\mathrm{m} \quad \left(\text{because Force (N)} = \text{mass (kg)} \times \text{acceleration } \left(\mathrm{m}\cdot\mathrm{s}^{-2}\right)\right) \\ & = & \mathrm{J} \quad (\text{Work (J)} = \text{Force (N)} \times \text{distance (m)}) \end{array} \tag{14.5}$$

We can do the same to prove that the unit for potential energy is equal to the joule:

$$\begin{array}{rcl} (\mathrm{kg})\left(\mathrm{m}\cdot\mathrm{s}^{-2}\right)(\mathrm{m}) & = & \mathrm{N}\cdot\mathrm{m} \\ & = & \mathrm{J} \end{array} \tag{14.6}$$

Exercise 14.3: Mixing Units & Energy Calculations A bullet, having a mass of $150\,\mathrm{g}$, is shot with a muzzle velocity of $960\,\mathrm{m}\cdot\mathrm{s}^{-1}$. Calculate its kinetic energy.

Solution to Exercise

Step 1.
- We are given the mass of the bullet $m = 150\,\mathrm{g}$. This is not the unit we want mass to be in. We need to convert to kg.

$$\begin{array}{rcl} \text{Mass in grams} \div 1000 & = & \text{Mass in kg} \\ 150\ \mathrm{g} \div 1000 & = & 0,150\ \mathrm{kg} \end{array} \tag{14.7}$$

- We are given the initial velocity with which the bullet leaves the barrel, called the muzzle velocity, and it is $v = 960\,\mathrm{m}\cdot\mathrm{s}^{-1}$.

Step 2.
- We are asked to find the kinetic energy.

Step 3. We just substitute the mass and velocity (which are known) into the equation for kinetic energy:

$$\begin{array}{rcl} E_K & = & \frac{1}{2}mv^2 \\ & = & \frac{1}{2}(0,150)(960)^2 \\ & = & 69\,120\,\mathrm{J} \end{array} \tag{14.8}$$

Kinetic Energy

1. Describe the relationship between an object's kinetic energy and its:
   a. mass and
   b. velocity
2. A stone with a mass of $100\,\mathrm{g}$ is thrown up into the air. It has an initial velocity of $3\,\mathrm{m \cdot s^{-1}}$. Calculate its kinetic energy
   a. as it leaves the thrower's hand.
   b. when it reaches its turning point.
3. A car with a mass of $700\,\mathrm{kg}$ is travelling at a constant velocity of $100\,\mathrm{km \cdot hr^{-1}}$. Calculate the kinetic energy of the car.

Find the answers with the shortcodes:

(1.) lad (2.) law (3.) lav

1.105 Mechanical Energy

(section shortcode: P10011)

TIP: Mechanical energy is the sum of the gravitational potential energy and the kinetic energy.

Mechanical energy, E_M, is simply the sum of gravitational potential energy (E_P) and the kinetic energy (E_K). Mechanical energy is defined as:

$$E_M = E_P + E_K \tag{14.9}$$

$$\begin{aligned} E_M &= E_P + E_K \\ E_M &= mgh + \tfrac{1}{2}mv^2 \end{aligned} \tag{14.10}$$

TIP: You may see mechanical energy written as U. We will not use this notation in this book, but you should be aware that this notation is sometimes used.

1.105.1 Conservation of Mechanical Energy

The Law of Conservation of Energy states:

Energy cannot be created or destroyed, but is merely changed from one form into another.

Definition: Conservation of Energy
The Law of Conservation of Energy: Energy cannot be created or destroyed, but is merely changed from one form into another.

So far we have looked at two types of energy: gravitational potential energy and kinetic energy. The sum of the gravitational potential energy and kinetic energy is called the mechanical energy. In a closed system, one where there are no external forces acting, the mechanical energy will remain constant. In other words, it will not change (become more or less). This is called the Law of Conservation of Mechanical Energy and it states:

The total amount of mechanical energy in a closed system remains constant.

Definition: Conservation of Mechanical Energy
Law of Conservation of Mechanical Energy: The total amount of mechanical energy in a closed system remains constant.

This means that potential energy can become kinetic energy, or vice versa, but energy cannot 'disappear'. The mechanical energy of an object moving in the Earth's gravitational field (or accelerating as a result of gravity) is constant or conserved, unless external forces, like air resistance, acts on the object.

We can now use the conservation of mechanical energy to calculate the velocity of a body in freefall and show that the velocity is independent of mass.

Show by using the law of conservation of energy that the velocity of a body in free fall is independent of its mass.

TIP: In problems involving the use of conservation of energy, the path taken by the object can be ignored. The only important quantities are the object's velocity (which gives its kinetic energy) and height above the reference point (which gives its gravitational potential energy).

TIP: In the absence of friction, mechanical energy is conserved and

$$E_{\text{M before}} = E_{\text{M after}} \tag{14.11}$$

In the presence of friction, mechanical energy is **not** conserved. The mechanical energy lost is equal to the work done against friction.

$$\Delta E_M = E_{\text{M before}} - E_{\text{M after}} = \text{work done against friction} \tag{14.12}$$

In general, mechanical energy is conserved in the absence of external forces. Examples of external forces are: applied forces, frictional forces and air resistance.

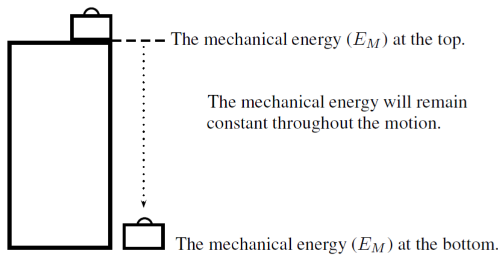
In the presence of internal forces like the force due to gravity or the force in a spring, mechanical energy is conserved.

The following simulation covers the law of conservation of energy.
(Simulation: ITd)

1.105.2 Using the Law of Conservation of Energy

Mechanical energy is conserved (in the absence of friction). Therefore we can say that the sum of the E_P and the E_K anywhere during the motion must be equal to the sum of the E_P and the E_K anywhere else in the motion.

We can now apply this to the example of the suitcase on the cupboard. Consider the mechanical energy of the suitcase at the top and at the bottom. We can say:



$$
\begin{aligned}
E_{\text{M top}} &= E_{\text{M bottom}} \\
E_{\text{P top}} + E_{\text{K top}} &= E_{\text{P bottom}} + E_{\text{K bottom}} \\
mgh + \frac{1}{2}mv^2 &= mgh + \frac{1}{2}mv^2 \\
(1)(9,8)(2) + 0 &= 0 + \frac{1}{2}(1)\left(v^2\right) \\
19,6 \text{ J} &= \frac{1}{2}v^2 \\
39,2 &= v^2 \\
v &= 6,26\,\mathrm{m \cdot s^{-1}}
\end{aligned}
\tag{14.13}
$$

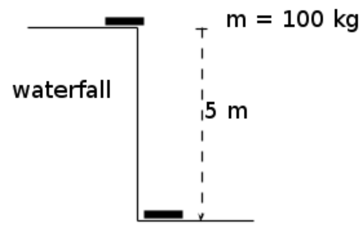
The suitcase will strike the ground with a velocity of $6,26\,\mathrm{m \cdot s^{-1}}$.

From this we see that when an object is lifted, like the suitcase in our example, it gains potential energy. As it falls back to the ground, it will lose this potential energy, but gain kinetic energy. We know that energy cannot be created or destroyed, but only changed from one form into another. In our example, the potential energy that the suitcase loses is changed to kinetic energy.

The suitcase will have maximum potential energy at the top of the cupboard and maximum kinetic energy at the bottom of the cupboard. Halfway down it will have half kinetic energy and half potential energy. As it moves down, the potential energy will be converted (changed) into kinetic energy until all the potential energy is gone and only kinetic energy is left. The $19,6$ J of potential energy at the top will become $19,6$ J of kinetic energy at the bottom.

Exercise 14.4: Using the Law of Conservation of Mechanical Energy During a flood a tree truck of mass $100\,\mathrm{kg}$ falls down a waterfall. The waterfall is $5\,\mathrm{m}$ high. If air resistance is ignored, calculate

1. the potential energy of the tree trunk at the top of the waterfall.
2. the kinetic energy of the tree trunk at the bottom of the waterfall.
3. the magnitude of the velocity of the tree trunk at the bottom of the waterfall.



Solution to Exercise

Step 1.
- The mass of the tree trunk $m = 100\,\mathrm{kg}$
- The height of the waterfall $h = 5\,\mathrm{m}$. These are all in SI units so we do not have to convert.

Step 2.
- Potential energy at the top
- Kinetic energy at the bottom
- Velocity at the bottom

Step 3.

$$\begin{aligned} E_P &= mgh \\ E_P &= (100)\,(9,8)\,(5) \\ E_P &= 4900 \text{ J} \end{aligned} \tag{14.14}$$

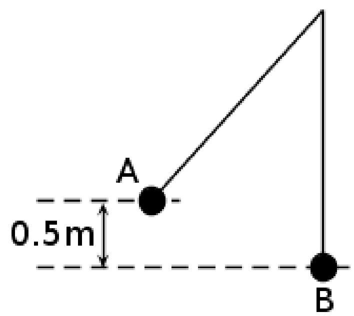
Step 4. The kinetic energy of the tree trunk at the bottom of the waterfall is equal to the potential energy it had at the top of the waterfall. Therefore $KE = 4900\,\mathrm{J}$.

Step 5. To calculate the velocity of the tree trunk we need to use the equation for kinetic energy.

$$\begin{aligned} E_K &= \tfrac{1}{2}mv^2 \\ 4900 &= \tfrac{1}{2}\,(100)\,\left(v^2\right) \\ 98 &= v^2 \\ v &= 9,899... \\ v &= 9,90\,\mathrm{m}\cdot\mathrm{s}^{-1}\mathrm{downwards} \end{aligned} \tag{14.15}$$

Exercise 14.5: Pendulum A $2\,\mathrm{kg}$ metal ball is suspended from a rope. If it is released from point A and swings down to the point B (the bottom of its arc):

1. Show that the velocity of the ball is independent of it mass.
2. Calculate the velocity of the ball at point B.



Solution to Exercise

Step 1.
- The mass of the metal ball is $m = 2\,\mathrm{kg}$
- The change in height going from point A to point B is $h = 0,5\,\mathrm{m}$
- The ball is released from point A so the velocity at point, $v_A = 0\,\mathrm{m \cdot s^{-1}}$.

All quantities are in SI units.

Step 2.
- Prove that the velocity is independent of mass.
- Find the velocity of the metal ball at point B.

Step 3. As there is no friction, mechanical energy is conserved. Therefore:

$$\begin{aligned} E_{M_A} &= E_{M_A} \\ E_{P_A} + E_{K_A} &= E_{P_A} + E_{K_A} \\ mgh_A + \tfrac{1}{2}m(v_A)^2 &= mgh_B + \tfrac{1}{2}m(v_B)^2 \\ mgh_A + 0 &= 0 + \tfrac{1}{2}m(v_B)^2 \\ mgh_A &= \tfrac{1}{2}m(v_B)^2 \end{aligned} \quad (14.16)$$

As the mass of the ball m appears on both sides of the equation, it can be eliminated so that the equation becomes:

$$gh_A = \frac{1}{2}(v_B)^2 \quad (14.17)$$

$$2gh_A = (v_B)^2 \quad (14.18)$$

This proves that the velocity of the ball is independent of its mass. It does not matter what its mass is, it will always have the same velocity when it falls through this height.

Step 4. We can use the equation above, or do the calculation from 'first principles':

$$\begin{aligned} (v_B)^2 &= 2gh_A \\ (v_B)^2 &= (2)\,(9.8)\,(0,5) \\ (v_B)^2 &= 9,8 \\ v_B &= \sqrt{9,8}\ \mathrm{m \cdot s^{-1}} \end{aligned} \quad (14.19)$$

Potential Energy

1. A tennis ball, of mass $120\,\mathrm{g}$, is dropped from a height of $5\,\mathrm{m}$. Ignore air friction.
   a. What is the potential energy of the ball when it has fallen $3\,\mathrm{m}$?
   b. What is the velocity of the ball when it hits the ground?
2. A bullet, mass $50\,\mathrm{g}$, is shot vertically up in the air with a muzzle velocity of $200\,\mathrm{m\cdot s^{-1}}$. Use the Principle of Conservation of Mechanical Energy to determine the height that the bullet will reach. Ignore air friction.
3. A skier, mass $50\,\mathrm{kg}$, is at the top of a $6,4\,\mathrm{m}$ ski slope.
   a. Determine the maximum velocity that she can reach when she skies to the bottom of the slope.
   b. Do you think that she will reach this velocity? Why/Why not?
4. A pendulum bob of mass $1,5\,\mathrm{kg}$, swings from a height A to the bottom of its arc at B. The velocity of the bob at B is $4\,\mathrm{m\cdot s^{-1}}$. Calculate the height A from which the bob was released. Ignore the effects of air friction.
5. Prove that the velocity of an object, in free fall, in a closed system, is independent of its mass.

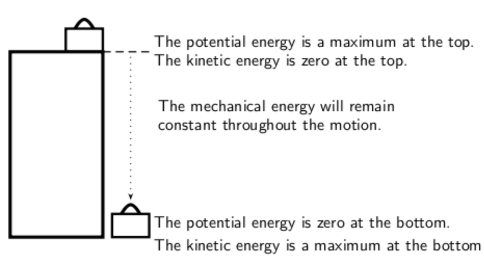
Find the answers with the shortcodes:

(1.) lxV (2.) lxp (3.) lxd (4.) lxv (5.) lxw

1.106 Energy graphs - (Not in CAPS, included for completeness)

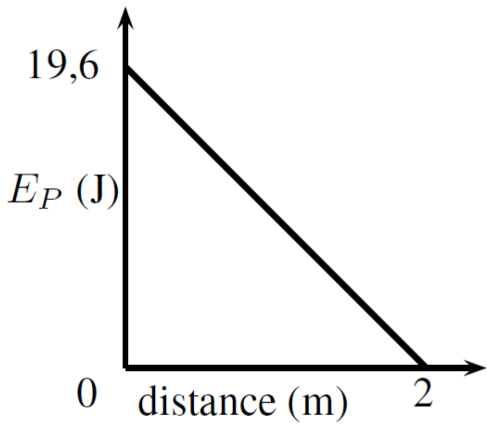
(section shortcode: P10012)

Let us consider our example of the suitcase on the cupboard, once more.

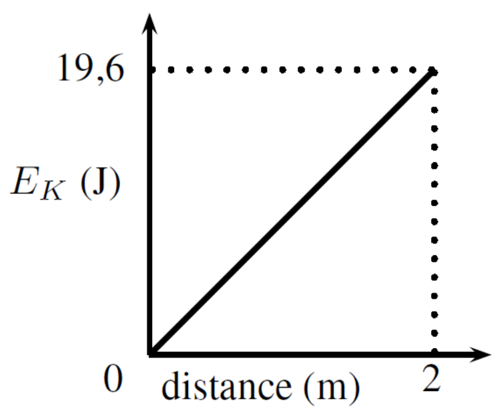


Let's look at each of these quantities and draw a graph for each. We will look at how each quantity changes as the suitcase falls from the top to the bottom of the cupboard.

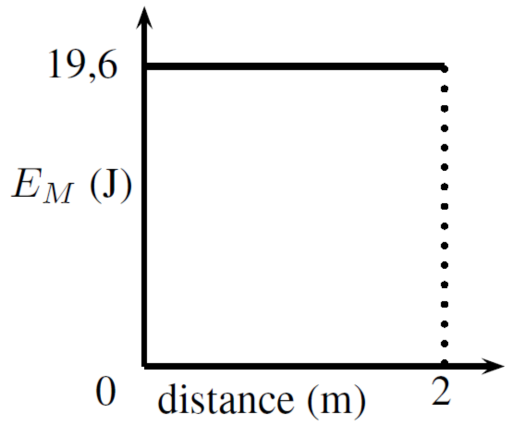
- **Potential energy**: The potential energy starts off at a maximum and decreases until it reaches zero at the bottom of the cupboard. It had fallen a distance of 2 metres.



- **Kinetic energy**: The kinetic energy is zero at the start of the fall. When the suitcase reaches the ground, the kinetic energy is a maximum. We also use distance on the x-axis.



- **Mechanical energy**: The mechanical energy is constant throughout the motion and is always a maximum. At any point in time, when we add the potential energy and the kinetic energy, we will get the same number.



1.107 Summary

(section shortcode: P10013)

- The potential energy of an object is the energy the object has due to his position above a reference point.
- The kinetic energy of an object is the energy the object has due to its motion.
- Mechanical energy of an object is the sum of the potential energy and kinetic energy of the object.
- The unit for energy is the joule (J).
- The Law of Conservation of Energy states that energy cannot be created or destroyed, but can only be changed from one form into another.
- The Law of Conservation of Mechanical Energy states that the total mechanical energy of an isolated system remains constant.
- The table below summarises the most important equations:

Potential Energy	$E_P = mgh$
Kinetic Energy	$E_K = \frac{1}{2}mv^2$
Mechanical Energy	$E_M = E_K + E_P$

Table 14.1

1.108 End of Chapter Exercises: Mechanical Energy

(section shortcode: P10014)

1. Give one word/term for the following descriptions.
   a. The force with which the Earth attracts a body.
   b. The unit for energy.

- **c.** The movement of a body in the Earth's gravitational field when no other forces act on it.
- **d.** The sum of the potential and kinetic energy of a body.
- **e.** The amount of matter an object is made up of.

2. Consider the situation where an apple falls from a tree. Indicate whether the following statements regarding this situation are TRUE or FALSE. Write only 'true' or 'false'. If the statement is false, write down the correct statement.

- **a.** The potential energy of the apple is a maximum when the apple lands on the ground.
- **b.** The kinetic energy remains constant throughout the motion.
- **c.** To calculate the potential energy of the apple we need the mass of the apple and the height of the tree.
- **d.** The mechanical energy is a maximum only at the beginning of the motion.
- **e.** The apple falls at an acceleration of $9,8\,\mathrm{m \cdot s^{-2}}$.

3. A man fires a rock out of a slingshot directly upward. The rock has an initial velocity of $15\,\mathrm{m \cdot s^{-1}}$.

- **a.** How long will it take for the rock to reach its highest point?
- **b.** What is the maximum height that the rock will reach?
- **c.** Draw graphs to show how the potential energy, kinetic energy and mechanical energy of the rock changes as it moves to its highest point.

4. A metal ball of mass $200\,\mathrm{g}$ is tied to a light string to make a pendulum. The ball is pulled to the side to a height (A), $10\,\mathrm{cm}$ above the lowest point of the swing (B). Air friction and the mass of the string can be ignored. The ball is let go to swing freely.

- **a.** Calculate the potential energy of the ball at point A.
- **b.** Calculate the kinetic energy of the ball at point B.
- **c.** What is the maximum velocity that the ball will reach during its motion?

5. A truck of mass $1,2\,\mathrm{tons}$ is parked at the top of a hill, $150\,\mathrm{m}$ high. The truck driver lets the truck run freely down the hill to the bottom.

- **a.** What is the maximum velocity that the truck can achieve at the bottom of the hill?
- **b.** Will the truck achieve this velocity? Why/why not?

6. A stone is dropped from a window, $3\,\mathrm{m}$ above the ground. The mass of the stone is $25\,\mathrm{g}$. Use the Principle of Conservation of Energy to calculate the velocity of the stone as it reaches the ground.

Find the answers with the shortcodes:

(1.) lxf (2.) lxG (3.) lxA (4.) lxo (5.) lxs (6.) lxH

Motion in One Dimension

1.109 Introduction

(section shortcode: P10015)

This chapter is about how things move in a straight line or more scientifically how things move *in one dimension*. This is useful for learning how to describe the movement of cars along a straight road or of trains along straight railway tracks. If you want to understand how any object moves, for example a car on the freeway, a soccer ball being kicked towards the goal or your dog chasing the neighbour's cat, then you have to understand three basic ideas about what it means when something *is moving*. These three ideas describe different parts of exactly how an object moves. They are:

1. position or displacement which tells us exactly where the object is,
2. speed or velocity which tells us exactly how fast the object's position is changing or more familiarly, how fast the object is moving, and
3. acceleration which tells us exactly how fast the object's velocity is changing.

You will also learn how to use position, displacement, speed, velocity and acceleration to describe the motion of simple objects. You will learn how to read and draw graphs that summarise the motion of a moving object. You will also learn about the equations that can be used to describe motion and how to apply these equations to objects moving in one dimension.

1.110 Reference Point, Frame of Reference and Position

(section shortcode: P10016)

The most important idea when studying motion, is you have to know where you are. The word *position* describes your location (where you are). However, saying that you are *here* is meaningless, and you have to specify your position *relative* to a known reference point. For example, if you are 2 m from the doorway, inside your classroom then your reference point is the doorway. This defines your position inside the classroom. Notice that you need a reference point (the doorway) and a direction (inside) to define your location.

1.110.1 Frames of Reference

Definition: Frame of Reference
A frame of reference is a reference point combined with a set of directions.

A *frame of reference* is similar to the idea of a reference point. A frame of reference is defined as a reference point combined with a set of directions. For example, a boy is standing still inside a train as it pulls out of a station. You are standing on the platform watching the train move from left to right. To you it looks as if the boy is moving from left to right, because relative to where you are standing (the platform), he is moving. According to the boy, and his *frame of reference* (the train), he is not moving.

A frame of reference must have an origin (where you are standing on the platform) and at least a positive direction. The train was moving from left to right, making to your right positive and to your left negative. If someone else was looking at the same boy, his frame of reference will be different. For example, if he was standing on the other side of the platform, the boy will be moving from right to left.

Another great example of frames of reference is a car overtaking another car on a road. Think about sitting in a taxi passing a car. If you sit in the taxi it is your origin and the direction it is moving in will be the positive direction. You will see the car slowly move further behind you (in the negative direction). The very important thing is that the driver of the car has their own reference frame in which things look different. In the driver's reference frame the taxi is moving ahead. In one case someone is falling behind, but in the other case someone is moving ahead. This is just a matter of perspective (from which reference frame you choose to view the situation).

For this chapter, we will only use frames of reference in the x-direction. Frames of reference will be covered in more detail in Grade 12.

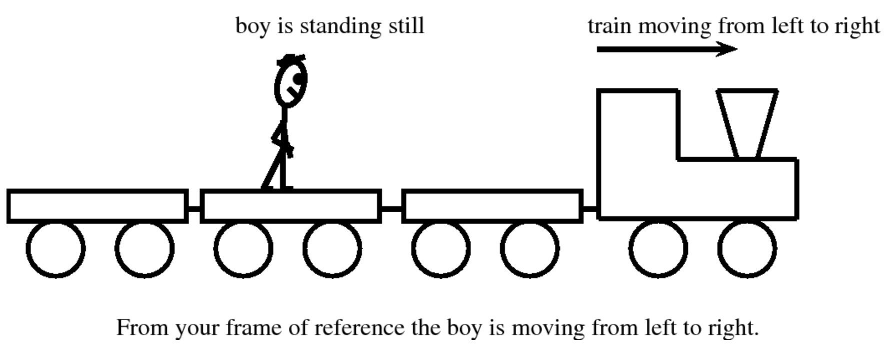
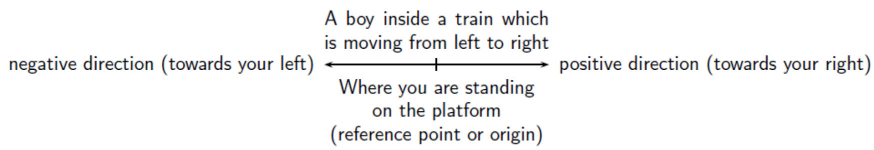


Figure 15.1: Frames of Reference



(Presentation: lb0)

1.110.2 Position

Definition: Position
Position is a measurement of a location, with reference to an origin.

A position is a measurement of a location, with reference to an origin. Positions can therefore be negative or positive. The symbol x is used to indicate position. x has units of length for example cm, m or km. Figure 15.4 shows the position of a school. Depending on what reference point we choose, we can say that the school is $300\,\mathrm{m}$ from Joan's house (with Joan's house as the reference point or origin) or $500\,\mathrm{m}$ from Joel's house (with Joel's house as the reference point or origin).

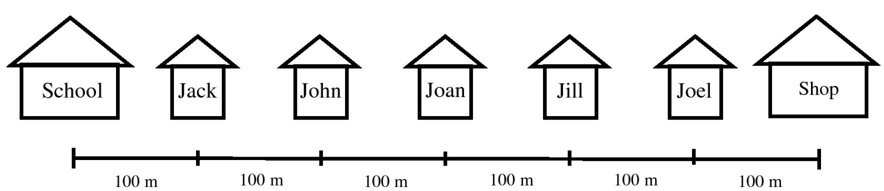


Figure 15.4: Illustration of position

The shop is also $300\,m$ from Joan's house, but in the opposite direction as the school. When we choose a reference point, we have a positive direction and a negative direction. If we choose the direction towards the school as positive, then the direction towards the shop is negative. A negative direction is always opposite to the direction chosen as positive.

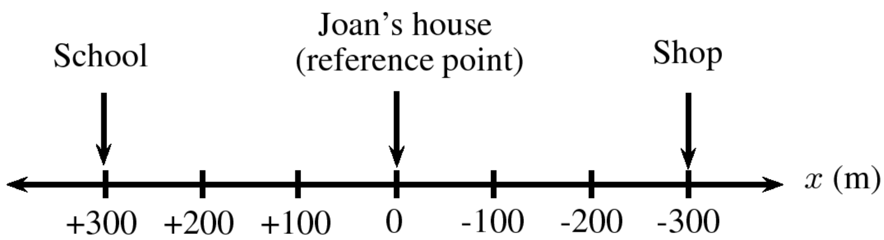


Figure 15.5: The origin is at Joan's house and the position of the school is +300 m. Positions towards the left are defined as positive and positions towards the right are defined as negative.

Discussion : Reference Points

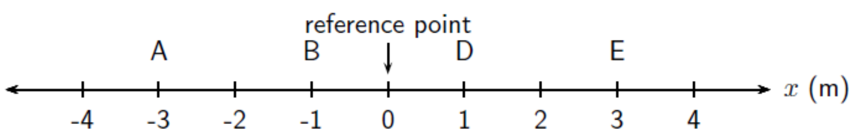
Divide into groups of 5 for this activity. On a straight line, choose a reference point. Since position can have both positive and negative values, discuss the advantages and disadvantages of choosing

1. either end of the line,
2. the middle of the line.

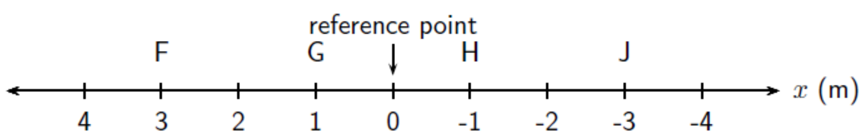
This reference point can also be called "the origin".

Position

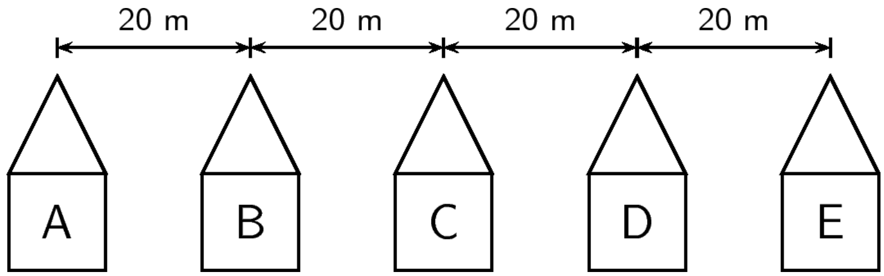
1. Write down the positions for objects at A, B, D and E. Do not forget the units.



2. Write down the positions for objects at F, G, H and J. Do not forget the units.



3. There are 5 houses on Newton Street, A, B, C, D and E. For all cases, assume that positions to the right are positive.



a. Draw a frame of reference with house A as the origin and write down the positions of houses B, C, D and E.
b. You live in house C. What is your position relative to house E?
c. What are the positions of houses A, B and D, if house B is taken as the reference point?

Find the answers with the shortcodes:

(1.) laG (2.) la7 (3.) laA

1.111 Displacement and Distance

(section shortcode: P10018)

Definition: Displacement
Displacement is the change in an object's position.

The displacement of an object is defined as its change in position (final position minus initial position). Displacement has a magnitude and direction and is therefore a vector. For example, if the initial position of a car is x_i and it moves to a final position of x_f, then the displacement is:

$$x_f - x_i \tag{15.1}$$

However, subtracting an initial quantity from a final quantity happens often in Physics, so we use the shortcut Δ to mean *final* - *initial*. Therefore, displacement can be written:

$$\Delta x = x_f - x_i \tag{15.2}$$

TIP: The symbol Δ is read out as *delta*. Δ is a letter of the Greek alphabet and is used in Mathematics and Science to indicate a change in a certain quantity, or a final value minus an initial value. For example, Δx means change in x while Δt means change in t.

TIP: The words *initial* and *final* will be used very often in Physics. *Initial* will always refer to something that happened earlier in time and *final* will always refer to something that happened later in time. It will often happen that the final value is smaller than the initial value, such that the difference is negative. This is ok!

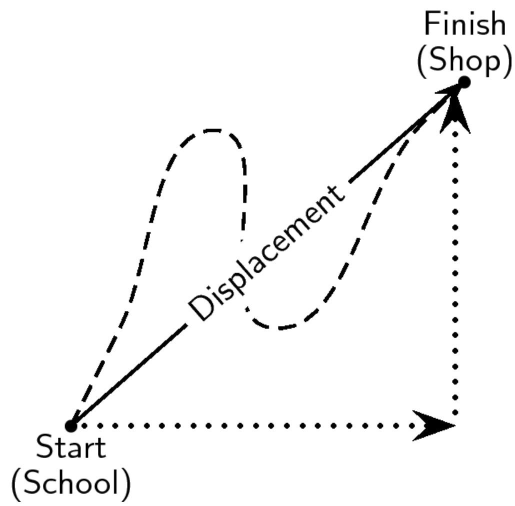


Figure 15.9: Illustration of displacement

Displacement does not depend on the path travelled, but only on the initial and final positions (Figure 15.9). We use the word *distance* to describe how far an object travels along a particular path. Distance is the actual distance that was covered. Distance (symbol d) does not have a direction, so it is a scalar. Displacement is the shortest distance from the starting point to the endpoint – from the school to the shop in the figure. Displacement has direction and is therefore a vector.

Figure 15.4 shows the five houses we discussed earlier. Jack walks to school, but instead of walking straight to school, he decided to walk to his friend Joel's house first to fetch him so that they can walk to school together. Jack covers a distance of $400\,\mathrm{m}$ to Joel's house and another $500\,\mathrm{m}$ to school. He covers a distance of $900\,\mathrm{m}$. His displacement, however, is only $100\,\mathrm{m}$ towards the school. This is because displacement only looks at the starting position (his house) and the end position (the school). It does not depend on the path he travelled.

To calculate his distance and displacement, we need to choose a reference point and a direction. Let's choose Jack's house as the reference point, and towards Joel's house as the positive direction (which means that towards the school is negative). We would do the calculations as follows:

$$\begin{aligned} \text{Distance}\,(\text{D}) &= \text{path travelled} \\ &= 400\ \text{m} + 500\ \text{m} \\ &= 900\ \text{m} \end{aligned} \qquad (15.3)$$

$$\begin{aligned}
\text{Displacement}\,(\Delta \text{x}) &= x_f - x_i \\
&= -100 \text{ m} + 0 \text{ m} \\
&= -100 \text{ m}
\end{aligned} \tag{15.4}$$

NOTE: You may also see d used for distance. We will use D in this book, but you may see d used in other books.

Joel walks to school with Jack and after school walks back home. What is Joel's displacement and what distance did he cover? For this calculation we use Joel's house as the reference point. Let's take towards the school as the positive direction.

$$\begin{aligned}
\text{Distance}\,(\text{D}) &= \text{path travelled} \\
&= 500 \text{ m} + 500 \text{ m} \\
&= 1000 \text{ m}
\end{aligned} \tag{15.5}$$

$$\begin{aligned}
\text{Displacement}\,(\Delta \text{x}) &= x_f - x_i \\
&= 0 \text{ m} + 0 \text{ m} \\
&= 0 \text{ m}
\end{aligned} \tag{15.6}$$

It is possible to have a displacement of $0\,\mathrm{m}$ and a distance that is not $0\,\mathrm{m}$. This happens when an object completes a round trip back to its original position, like an athlete running around a track.

1.111.1 Interpreting Direction

Very often in calculations you will get a negative answer. For example, Jack's displacement in the example above, is calculated as $-100\,\mathrm{m}$. The minus sign in front of the answer means that his displacement is $100\,\mathrm{m}$ in the opposite direction (opposite to the direction chosen as positive in the beginning of the question). When we start a calculation we choose a frame of reference and a positive direction. In the first example above, the reference point is Jack's house and the positive direction is towards Joel's house. Therefore Jack's displacement is $100\,\mathrm{m}$ towards the school. Notice that distance has no direction, but displacement has.

1.111.2 Differences between Distance and Displacement

Definition: Vectors and Scalars
A vector is a physical quantity with magnitude (size) and direction. A scalar is a physical quantity with magnitude (size) only.

The differences between distance and displacement can be summarised as:

Distance	**Displacement**
1. depends on the path	1. independent of path taken
2. always positive	2. can be positive or negative
3. is a scalar	3. is a vector
4. does not have a direction	4. has a direction

Table 15.1

Point of Reference

1. Use Figure 15.4 to answer the following questions.
   a. Jill walks to Joan's house and then to school, what is her distance and displacement?
   b. John walks to Joan's house and then to school, what is his distance and displacement?
   c. Jack walks to the shop and then to school, what is his distance and displacement?
   d. What reference point did you use for each of the above questions?
2. You stand at the front door of your house (displacement, $\Delta x = 0\,\mathrm{m}$). The street is $10\,\mathrm{m}$ away from the front door. You walk to the street and back again.
   a. What is the distance you have walked?
   b. What is your final displacement?
   c. Is displacement a vector or a scalar? Give a reason for your answer.

Find the answers with the shortcodes:

(1.) lao (2.) las

1.112 Speed, Average Velocity and Instantaneous Velocity

(section shortcode: P10019)

Definition: Velocity
Velocity is the rate of change of displacement.

Definition: Instantaneous velocity
Instantaneous velocity is the velocity of a body at a specific instant in time.

Definition: Average velocity
Average velocity is the total displacement of a body over a time interval.

Velocity is the rate of change of position. It tells us how much an object's position changes in time. This is the same as the displacement divided by the time taken. Since displacement is a vector and time taken is a scalar, velocity is also a vector. We use the symbol v for velocity. If we have a displacement of Δx and a time taken of Δt, v is then defined as:

$$\begin{aligned} \text{velocity } \left(\text{in m}\cdot\text{s}^{-1}\right) &= \frac{\text{change in displacement (in m)}}{\text{change in time (in s)}} \\ v &= \frac{\Delta x}{\Delta t} \end{aligned} \tag{15.7}$$

Velocity can be positive or negative. Positive values of velocity mean that the object is moving away from the reference point or origin and negative values mean that the object is moving towards the reference point or origin.

TIP: An instant in time is different from the time taken or the time interval. It is therefore useful to use the symbol t for an instant in time (for example during the 4^{th} second) and the symbol Δt for the time taken (for example during the first 5 seconds of the motion).

Average velocity (symbol v) is the displacement for the whole motion divided by the time taken for the whole motion. Instantaneous velocity is the velocity at a specific instant in time.

This is terminology that occurs quite often and it is important to always remember that instantaneous and average quantities are not always the same. In fact, they can be very different. The magnitude of the instantaneous velocity is the same as the slope of the line which is a tangent to the displacement curve at the time of measurement. The magnitude of the average velocity is the same as the slope of the line between the start and end points of the interval.

If you want to think about why the tangent at a particular time gives us the velocity remember that the velocity is the slope of the displacement curve. Now, a simple why to start thinking about it is to image you could zoom (magnify) the displacement curve at the point. The more you zoom in the more it will look like a straight line at the point and that straight line will be very similar to the tangent line at the point. This isn't a mathematical proof but you will learn one later on in mathematics about limits and slopes.

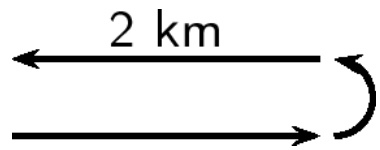
(Average) Speed (symbol s) is the distance travelled (D) divided by the time taken (Δt) for the journey. Distance and time are scalars and therefore speed will also be a scalar. Speed is calculated as follows:

$$\text{speed } \left(\text{in m}\cdot\text{s}^{-1}\right) = \frac{\text{distance (in m)}}{\text{time (in s)}} \tag{15.8}$$

$$s = \frac{D}{\Delta t} \tag{15.9}$$

Instantaneous speed is the magnitude of instantaneous velocity. It has the same value, but no direction.

Exercise 15.1: Average speed and average velocity James walks 2 km away from home in 30 minutes. He then turns around and walks back home along the same path, also in 30 minutes. Calculate James' average speed and average velocity.



Solution to Exercise

Step 1. The question explicitly gives

- the distance and time out (2 km in 30 minutes)
- the distance and time back (2 km in 30 minutes)

Step 2. The information is not in SI units and must therefore be converted. To convert km to m, we know that:

$$\begin{aligned} 1\ \text{km} &= 1\ 000\ \text{m} \\ \therefore 2\ \text{km} &= 2\ 000\ \text{m}\,(\text{multiply both sides by}\,2) \end{aligned} \quad (15.10)$$

Similarly, to convert 30 minutes to seconds,

$$\begin{aligned} 1\ \text{min} &= 60\,\text{s} \\ \therefore 30\ \text{min} &= 1\ 800\ \text{s}\,(\text{multiply both sides by}\,30) \end{aligned} \quad (15.11)$$

Step 3. James started at home and returned home, so his displacement is 0 m.

$$\Delta x = 0\ \text{m} \quad (15.12)$$

James walked a total distance of 4 000 m (2 000 m out and 2 000 m back).

$$D = 4\ 000\,\text{m} \quad (15.13)$$

Step 4. James took 1 800 s to walk out and 1 800 s to walk back.

$$\Delta t = 3\ 600\,\text{s} \quad (15.14)$$

Step 5.

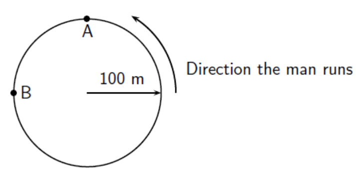
$$\begin{aligned} s &= \frac{D}{\Delta t} \\ &= \frac{4\ 000\ \text{m}}{3\ 600\ \text{s}} \\ &= 1,11\ \text{m} \cdot \text{s}^{-1} \end{aligned} \quad (15.15)$$

Step 6.

$$
\begin{aligned}
v &= \frac{\Delta x}{\Delta t} \\
&= \frac{0 \ \mathrm{m}}{3 \ 600 \ \mathrm{s}} \\
&= 0 \ \mathrm{m} \cdot \mathrm{s}^{-1}
\end{aligned}
\tag{15.16}
$$

Exercise 15.2: Instantaneous Speed and Velocity A man runs around a circular track of radius $100\,\mathrm{m}$. It takes him $120\,\mathrm{s}$ to complete a revolution of the track. If he runs at constant speed, calculate:

1. his speed,
2. his instantaneous velocity at point A,
3. his instantaneous velocity at point B,
4. his average velocity between points A and B,
5. his average speed during a revolution.
6. his average velocity during a revolution.



Solution to Exercise

Step 1. To determine the man's speed we need to know the distance he travels and how long it takes. We know it takes 120 s to complete one revolution of the track.(A revolution is to go around the track once.)

Step 2. What distance is one revolution of the track? We know the track is a circle and we know its radius, so we can determine the distance around the circle. We start with the equation for the circumference of a circle

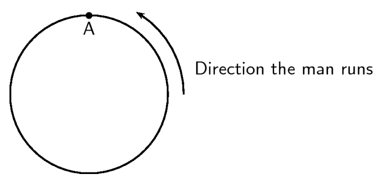
$$
\begin{aligned}
C &= 2\pi r \\
&= 2\pi\,(100\,\mathrm{m}) \\
&= 628,32\,\mathrm{m}
\end{aligned}
\tag{15.17}
$$

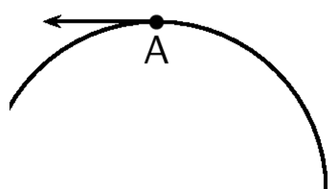
Therefore, the distance the man covers in one revolution is $628,32\,\mathrm{m}$.

Step 3. We know that speed is distance covered per unit time. So if we divide the distance covered by the time it took we will know how much distance was covered for every unit of time. No direction is used here because speed is a scalar.

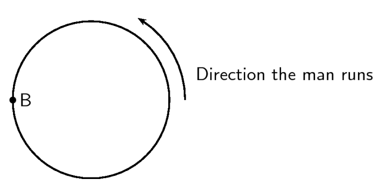
$$\begin{aligned} s &= \frac{D}{\Delta t} \\ &= \frac{628{,}32\,\mathrm{m}}{120\,\mathrm{s}} \\ &= 5{,}24\,\mathrm{m\cdot s^{-1}} \end{aligned} \tag{15.18}$$

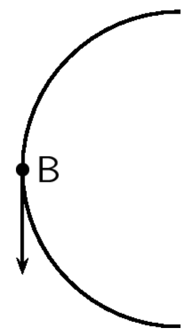
Step 4. Consider the point A in the diagram. We know which way the man is running around the track and we know his speed. His velocity at point A will be his speed (the magnitude of the velocity) plus his direction of motion (the direction of his velocity). The instant that he arrives at A he is moving as indicated in the diagram. His velocity will be $5,24\,\mathrm{m\cdot s^{-1}}$ West.





Step 5. Consider the point B in the diagram. We know which way the man is running around the track and we know his speed. His velocity at point B will be his speed (the magnitude of the velocity) plus his direction of motion (the direction of his velocity). The instant that he arrives at B he is moving as indicated in the diagram. His velocity will be $5,24\,\mathrm{m\cdot s^{-1}}$ South.



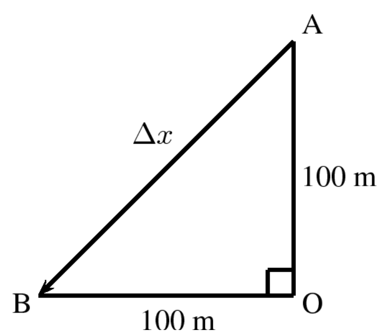


Step 6. To determine the average velocity between A and B, we need the change in displacement between A and B and the change in time between A and B. The displacement from A and B can be calculated by using the Theorem of Pythagoras:

$$\begin{aligned}(\Delta x)^2 &= (100\text{m})^2 + (100\text{m})^2 \\ &= 20000\text{m} \\ \Delta x &= 141,42135...\ \text{m}\end{aligned} \tag{15.19}$$

The time for a full revolution is 120 s, therefore the time for a $\frac{1}{4}$ of a revolution is 30 s.

$$\begin{aligned}v_{AB} &= \frac{\Delta x}{\Delta t} \\ &= \frac{141{,}42...\text{m}}{30\text{s}} \\ &= 4.71\ \ \text{m}\cdot\text{s}^{-1}\end{aligned} \tag{15.20}$$



Velocity is a vector and needs a direction.
Triangle AOB is isosceles and therefore angle BAO = 45°.
The direction is between west and south and is therefore southwest.
The final answer is: $v = 4.71\,\text{m}\cdot\text{s}^{-1}$, southwest.

Step 7. Because he runs at a constant rate, we know that his speed anywhere around the track will be the same. His average speed is $5,24\,\text{m}\cdot\text{s}^{-1}$.

TIP: Remember - displacement can be zero even when distance travelled is not!

To calculate average velocity we need his total displacement and his total time. His displacement is zero because he ends up where he started. His time is $120\,\mathrm{s}$. Using these we can calculate his average velocity:

$$\begin{aligned} v &= \frac{\Delta x}{\Delta t} \\ &= \frac{0\ \mathrm{m}}{120\ \mathrm{s}} \\ &= 0\ \mathrm{m} \cdot \mathrm{s}^{-1} \end{aligned} \tag{15.21}$$

1.112.1 Differences between Speed and Velocity

The differences between speed and velocity can be summarised as:

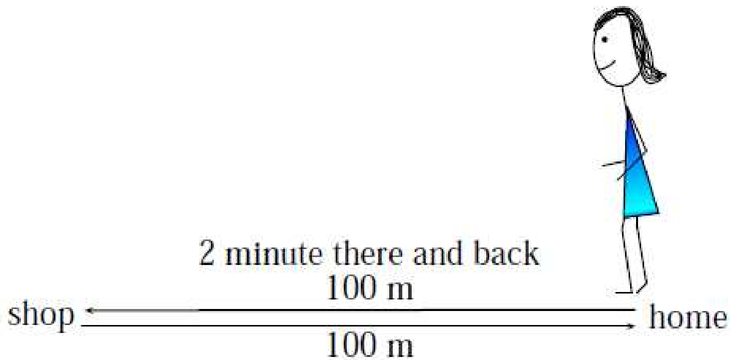
Speed	Velocity
1. depends on the path taken	1. independent of path taken
2. always positive	2. can be positive or negative
3. is a scalar	3. is a vector
4. no dependence on direction and so is only positive	4. direction can be guessed from the sign (i.e. positive or negative)

Table 15.2

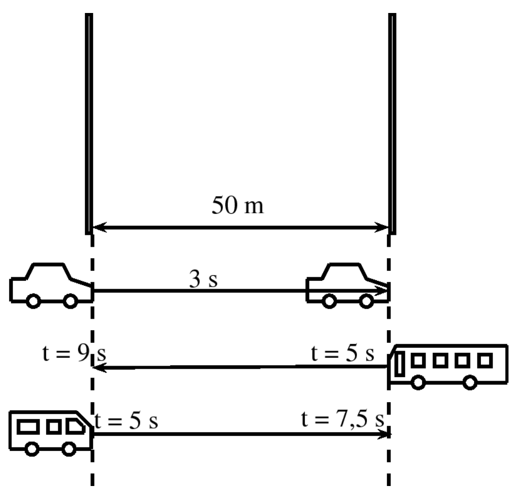
Additionally, an object that makes a round trip, i.e. travels away from its starting point and then returns to the same point has zero velocity but travels a non-zero speed.

Displacement and related quantities

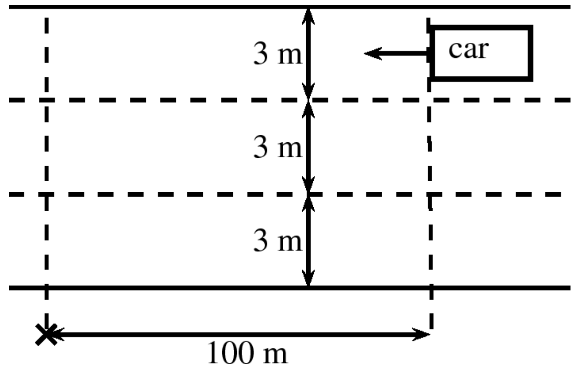
1. Theresa has to walk to the shop to buy some milk. After walking $100\,\mathrm{m}$, she realises that she does not have enough money, and goes back home. If it took her two minutes to leave and come back, calculate the following:
   a. How long was she out of the house (the time interval Δt in seconds)?
   b. How far did she walk (distance (D))?
   c. What was her displacement (Δx)?
   d. What was her average velocity (in m·s^{-1})?
   e. What was her average speed (in m·s^{-1})?



2. Desmond is watching a straight stretch of road from his classroom window. He can see two poles which he earlier measured to be $50\,\mathrm{m}$ apart. Using his stopwatch, Desmond notices that it takes $3\,\mathrm{s}$ for most cars to travel from the one pole to the other.
   a. Using the equation for velocity ($v = \frac{\Delta x}{\Delta t}$), show all the working needed to calculate the velocity of a car travelling from the left to the right.
   b. If Desmond measures the velocity of a red Golf to be $-16,67\,\mathrm{m \cdot s^{-1}}$, in which direction was the Gold travelling? Desmond leaves his stopwatch running, and notices that at $t = 5,0\,\mathrm{s}$, a taxi passes the left pole at the same time as a bus passes the right pole. At time $t = 7,5\,\mathrm{s}$ the taxi passes the right pole. At time $t = 9,0\,\mathrm{s}$, the bus passes the left pole.
   c. How long did it take the taxi and the bus to travel the distance between the poles? (Calculate the time interval (Δt) for both the taxi and the bus).
   d. What was the velocity of the taxi and the bus?
   e. What was the speed of the taxi and the bus?
   f. What was the speed of taxi and the bus in $\mathrm{km \cdot h^{-1}}$?



3. A rabbit runs across a freeway. There is a car, $100\,\mathrm{m}$ away travelling towards the rabbit.



a. If the car is travelling at $120\,\text{km}\cdot\text{h}^{-1}$, what is the car's speed in $\text{m}\cdot\text{s}^{-1}$.
b. How long will it take the a car to travel $100\,\text{m}$?
c. If the rabbit is running at $10\,\text{km}\cdot\text{h}^{-1}$, what is its speed in $\text{m}\cdot\text{s}^{-1}$?
d. If the freeway has 3 lanes, and each lane is $3\,\text{m}$ wide, how long will it take for the rabbit to cross all three lanes?
e. If the car is travelling in the furthermost lane from the rabbit, will the rabbit be able to cross all 3 lanes of the freeway safely?

Find the answers with the shortcodes:

(1.) laH (2.) la6 (3.) laF

Investigation : An Exercise in Safety

Divide into groups of 4 and perform the following investigation. Each group will be performing the same investigation, but the aim for each group will be different.

1. Choose an aim for your investigation from the following list and formulate a hypothesis:
   - Do cars travel at the correct speed limit?
   - Is is safe to cross the road outside of a pedestrian crossing?
   - Does the colour of your car determine the speed you are travelling at?
   - Any other relevant question that you would like to investigate.
2. On a road that you often cross, measure out $50\,\text{m}$ along a straight section, far away from traffic lights or intersections.
3. Use a stopwatch to record the time each of 20 cars take to travel the $50\,\text{m}$ section you measured.
4. Design a table to represent your results. Use the results to answer the question posed in the aim of the investigation. You might need to do some more measurements for your investigation. Plan in your group what else needs to be done.
5. Complete any additional measurements and write up your investigation under the following headings:
   - Aim and Hypothesis
   - Apparatus
   - Method
   - Results
   - Discussion
   - Conclusion

6. Answer the following questions:
   a. How many cars took less than $3\,\mathrm{s}$ to travel $50\,\mathrm{m}$?
   b. What was the shortest time a car took to travel $50\,\mathrm{m}$?
   c. What was the average time taken by the 20 cars?
   d. What was the average speed of the 20 cars?
   e. Convert the average speed to $\mathrm{km}\cdot\mathrm{h}^{-1}$.

1.113 Acceleration

(section shortcode: P10020)

Definition: Acceleration
Acceleration is the rate of change of velocity.

Acceleration (symbol a) is the rate of change of velocity. It is a measure of how fast the velocity of an object changes in time. If we have a change in velocity (Δv) over a time interval (Δt), then the acceleration (a) is defined as:

$$\text{acceleration}\ \left(\text{in m}\cdot\text{s}^{-2}\right) = \frac{\text{change in velocity}\ \left(\text{in m}\cdot\text{s}^{-1}\right)}{\text{change in time}\ (\text{in s})} \tag{15.22}$$

$$a = \frac{\Delta v}{\Delta t} \tag{15.23}$$

Since velocity is a vector, acceleration is also a vector. Acceleration does not provide any information about a motion, but only about how the motion changes. It is not possible to tell how fast an object is moving or in which direction from the acceleration.

Like velocity, acceleration can be negative or positive. We see that when the sign of the acceleration and the velocity are the same, the object is speeding up. If both velocity and acceleration are positive, the object is speeding up in a positive direction. If both velocity and acceleration are negative, the object is speeding up in a negative direction. If velocity is positive and acceleration is negative, then the object is slowing down. Similarly, if the velocity is negative and the acceleration is positive the object is slowing down. This is illustrated in the following worked example.

Exercise 15.3: Acceleration A car accelerates uniformly from an initial velocity of 2 m·s^{-1} to a final velocity of 10 m·s^{1} in 8 seconds. It then slows down uniformly to a final velocity of 4 m·s^{-1} in 6 seconds. Calculate the acceleration of the car during the first 8 seconds and during the last 6 seconds.

Solution to Exercise

Step 1. Consider the motion of the car in two parts: the first 8 seconds and the last 6 seconds.
For the first 8 seconds:

$$\begin{aligned} v_i &= 2\ \mathrm{m \cdot s^{-1}} \\ v_f &= 10\ \mathrm{m \cdot s^{-1}} \\ t_i &= 0\ \mathrm{s} \\ t_f &= 8\ \mathrm{s} \end{aligned} \tag{15.24}$$

For the last 6 seconds:

$$\begin{aligned} v_i &= 10\ \mathrm{m \cdot s^{-1}} \\ v_f &= 4\ \mathrm{m \cdot s^{-1}} \\ t_i &= 8\ \mathrm{s} \\ t_f &= 14\ \mathrm{s} \end{aligned} \tag{15.25}$$

Step 2. For the first 8 seconds:

$$\begin{aligned} a &= \frac{\Delta v}{\Delta t} \\ &= \frac{10\mathrm{m} \cdot s^{-1} - 2\mathrm{m \cdot s^{-1}}}{8\mathrm{s} - 0\mathrm{s}} \\ &= 1\ \mathrm{m \cdot s^{-2}} \end{aligned} \tag{15.26}$$

For the next 6 seconds:

$$\begin{aligned} a &= \frac{\Delta v}{\Delta t} \\ &= \frac{4\mathrm{m \cdot s^{-1}} - 10\mathrm{m \cdot s^{-1}}}{14\mathrm{s} - 8\mathrm{s}} \\ &= -1\ \mathrm{m \cdot s^{-2}} \end{aligned} \tag{15.27}$$

During the first 8 seconds the car had a positive acceleration. This means that its velocity increased. The velocity is positive so the car is speeding up. During the next 6 seconds the car had a negative acceleration. This means that its velocity decreased. The velocity is negative so the car is slowing down.

TIP: Acceleration does not tell us about the direction of the motion. Acceleration only tells us how the velocity changes.

TIP: Avoid the use of the word *deceleration* to refer to a negative acceleration. This word usually means *slowing down* and it is possible for an object to slow down with both a positive and negative acceleration, because the sign of the velocity of the object must also be taken into account to determine whether the body is slowing down or not.

1.113.1 Acceleration

1. An athlete is accelerating uniformly from an initial velocity of 0 m·s^{-1}to a final velocity of 4 m·s^{-1}in 2 seconds. Calculate his acceleration. Let the direction that the athlete is running in be the positive direction.

2. A bus accelerates uniformly from an initial velocity of 15 m·s^{-1}to a final velocity of 7 m·s^{-1}in 4 seconds. Calculate the acceleration of the bus. Let the direction of motion of the bus be the positive direction.

3. An aeroplane accelerates uniformly from an initial velocity of 200 m·s^{-1}to a velocity of 100 m·s^{-1}in 10 seconds. It then accelerates uniformly to a final velocity of 240 m·s^{-1}in 20 seconds. Let the direction of motion of the aeroplane be the positive direction.
   a. Calculate the acceleration of the aeroplane during the first 10 seconds of the motion.
   b. Calculate the acceleration of the aeroplane during the next 14 seconds of its motion.

The following video provides a summary of distance, velocity and acceleration. Note that in this video a different convention for writing units is used. You should not use this convention when writing units in physics.

Khan academy video on motion - 1 (Video: P10021)

Find the answers with the shortcodes:

(1.) l1k (2.) l10 (3.) l18

1.114 Description of Motion

(section shortcode: P10022)

The purpose of this chapter is to describe motion, and now that we understand the definitions of displacement, distance, velocity, speed and acceleration, we are ready to start using these ideas to describe how an object is moving. There are many ways of describing motion:

1. words
2. diagrams
3. graphs

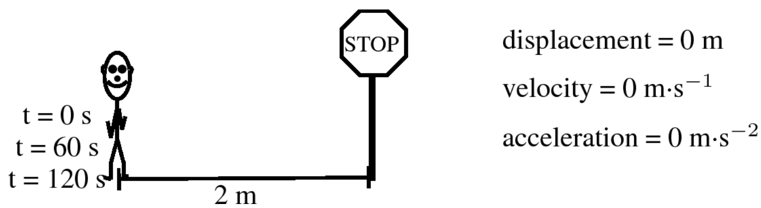
These methods will be described in this section.

We will consider three types of motion: when the object is not moving (stationary object), when the object is moving at a constant velocity (uniform motion) and when the object is moving at a constant acceleration (motion at constant acceleration).

1.114.1 Stationary Object

The simplest motion that we can come across is that of a stationary object. A stationary object does not move and so its position does not change, for as long as it is standing still. An example of this situation is when someone is waiting for something without moving. The person remains in the same position.

Lesedi is waiting for a taxi. He is standing two metres from a stop street at t = 0 s. After one minute, at t = 60 s, he is still 2 metres from the stop street and after two minutes, at t = 120 s, also 2 metres from the stop street. His position has not changed. His displacement is zero (because his position is the same), his velocity is zero (because his displacement is zero) and his acceleration is also zero (because his velocity is not changing).



We can now draw graphs of position vs. time (x vs. t), velocity vs. time (v vs. t) and acceleration vs. time (a vs. t) for a stationary object. The graphs are shown in Figure 15.22. Lesedi's position is 2 metres from the stop street. If the stop street is taken as the reference point, his position remains at 2 metres for 120 seconds. The graph is a horizontal line at 2 m. The velocity and acceleration graphs are also shown. They are both horizontal lines on the x-axis. Since his position is not changing, his velocity is $0\,\mathrm{m\cdot s^{-1}}$ and since velocity is not changing, acceleration is $0\,\mathrm{m\cdot s^{-2}}$.

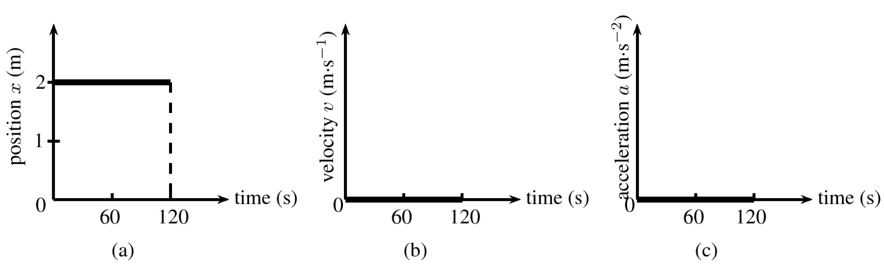


Figure 15.22: Graphs for a stationary object (a) position vs. time (b) velocity vs. time (c) acceleration vs. time.

Definition: Gradient
The gradient of a line can be calculated by dividing the change in the y-value by the change in the x-value.
m = $\frac{\Delta y}{\Delta x}$

Since we know that velocity is the rate of change of position, we can confirm the value for the velocity vs. time graph, by calculating the gradient of the x vs. t graph.

TIP: The gradient of a position vs. time graph gives the velocity.

If we calculate the gradient of the x vs. t graph for a stationary object we get:

$$\begin{aligned} v &= \frac{\Delta x}{\Delta t} \\ &= \frac{x_f - x_i}{t_f - t_i} \\ &= \frac{2\ \mathrm{m} - 2\ \mathrm{m}}{120\ \mathrm{s} - 60\ \mathrm{s}}\ (\text{initial position} = \text{final position}) \\ &= 0\ \ \mathrm{m}\cdot\mathrm{s}^{-1}\ (\text{for the time that Lesedi is stationary}) \end{aligned} \tag{15.28}$$

Similarly, we can confirm the value of the acceleration by calculating the gradient of the velocity vs. time graph.

TIP: The gradient of a velocity vs. time graph gives the acceleration.

If we calculate the gradient of the v vs. t graph for a stationary object we get:

$$\begin{aligned} a &= \frac{\Delta v}{\Delta t} \\ &= \frac{v_f - v_i}{t_f - t_i} \\ &= \frac{0\ \ \mathrm{m}\cdot\mathrm{s}^{-1} - 0\ \ \mathrm{m}\cdot\mathrm{s}^{-1}}{120\ \mathrm{s} - 60\ \mathrm{s}} \\ &= 0\ \ \mathrm{m}\cdot\mathrm{s}^{-2} \end{aligned} \tag{15.29}$$

Additionally, because the velocity vs. time graph is related to the position vs. time graph, we can use the area under the velocity vs. time graph to calculate the displacement of an object.

TIP: The area under the velocity vs. time graph gives the displacement.

The displacement of the object is given by the area under the graph, which is $0\,\mathrm{m}$. This is obvious, because the object is not moving.

1.114.2 Motion at Constant Velocity

Motion at a constant velocity or *uniform motion* means that the position of the object is changing at the same rate.

Assume that Lesedi takes $100\,\mathrm{s}$ to walk the $100\,\mathrm{m}$ to the taxi-stop every morning. If we assume that Lesedi's house is the origin, then Lesedi's velocity is:

$$\begin{aligned} v &= \frac{\Delta x}{\Delta t} \\ &= \frac{x_f - x_i}{t_f - t_i} \\ &= \frac{100\ \mathrm{m} - 0\ \mathrm{m}}{100\ \mathrm{s} - 0\ \mathrm{s}} \\ &= 1\ \mathrm{m} \cdot \mathrm{s}^{-1} \end{aligned} \tag{15.30}$$

Lesedi's velocity is 1 m·s^{-1}. This means that he walked $1\,\mathrm{m}$ in the first second, another metre in the second second, and another in the third second, and so on. For example, after $50\,\mathrm{s}$ he will be $50\,\mathrm{m}$ from home. His position increases by $1\,\mathrm{m}$ every $1\,\mathrm{s}$. A diagram of Lesedi's position is shown in Figure 15.23.

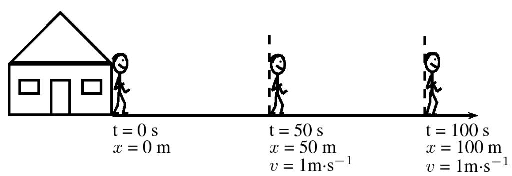


Figure 15.23: Diagram showing Lesedi's motion at a constant velocity of 1 m·s^{-1}

We can now draw graphs of position vs.time (x vs. t), velocity vs.time (v vs. t) and acceleration vs.time (a vs. t) for Lesedi moving at a constant velocity. The graphs are shown in Figure 15.24.

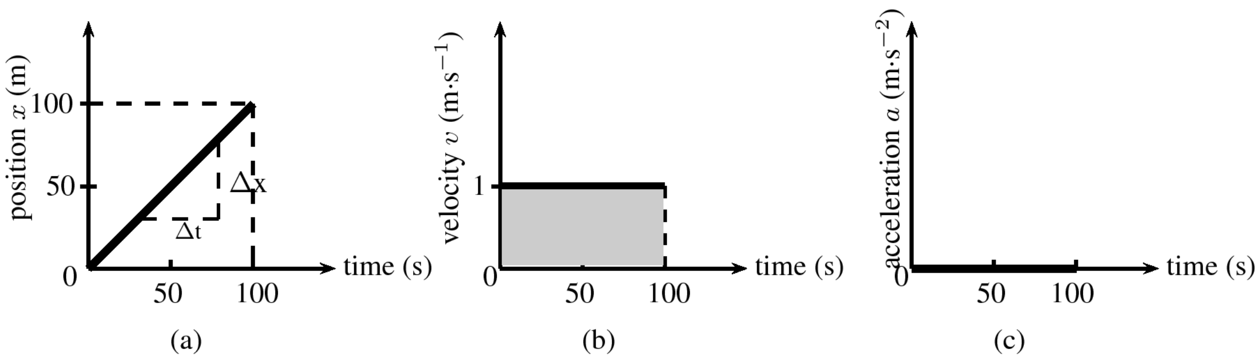


Figure 15.24: Graphs for motion at constant velocity (a) position vs. time (b) velocity vs. time (c) acceleration vs. time. The area of the shaded portion in the v vs. t graph corresponds to the object's displacement.

In the evening Lesedi walks $100\,\mathrm{m}$ from the bus stop to his house in $100\,\mathrm{s}$. Assume that Lesedi's house is the origin. The following graphs can be drawn to describe the motion.

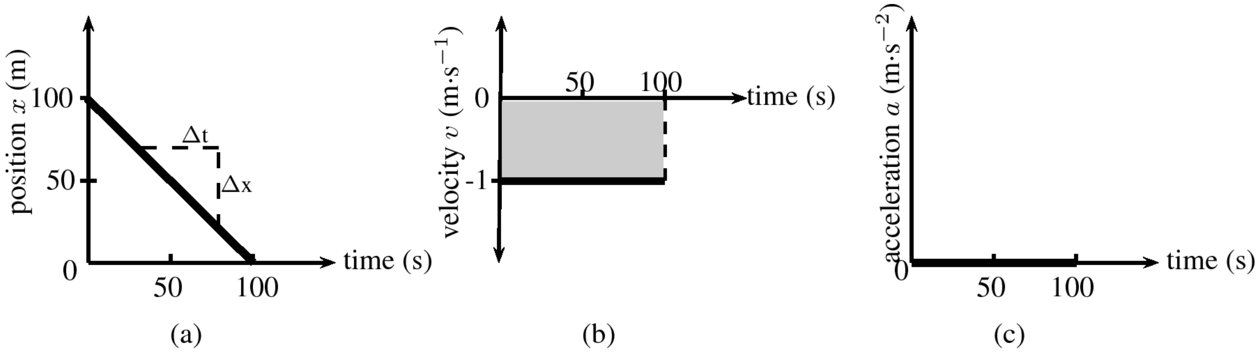


Figure 15.25: Graphs for motion with a constant negative velocity (a) position vs. time (b) velocity vs. time (c) acceleration vs. time. The area of the shaded portion in the v vs.t graph corresponds to the object's displacement.

We see that the v vs. t graph is a horisontal line. If the velocity vs. time graph is a horisontal line, it means that the velocity is *constant* (not changing). Motion at a constant velocity is known as *uniform motion*.

We can use the x vs. t to calculate the velocity by finding the gradient of the line.

$$\begin{aligned} v &= \frac{\Delta x}{\Delta t} \\ &= \frac{x_f - x_i}{t_f - t_i} \\ &= \frac{0\ \text{m} - 100\ \text{m}}{100\ \text{s} - 0\ \text{s}} \\ &= -1\ \text{m} \cdot \text{s}^{-1} \end{aligned} \tag{15.31}$$

Lesedi has a velocity of $-1\,\text{m}\cdot\text{s}^{-1}$, or $1\,\text{m}\cdot\text{s}^{-1}$ towards his house. You will notice that the v vs. t graph is a horisontal line corresponding to a velocity of $-1\,\text{m}\cdot\text{s}^{-1}$. The horizontal line means that the velocity stays the same (remains constant) during the motion. This is uniform velocity.

We can use the v vs. t to calculate the acceleration by finding the gradient of the line.

$$\begin{aligned} a &= \frac{\Delta v}{\Delta t} \\ &= \frac{v_f - v_i}{t_f - t_i} \\ &= \frac{1\ \text{m}\cdot\text{s}^{-1} - 1\ \text{m}\cdot\text{s}^{-1}}{100\ \text{s} - 0\ \text{s}} \\ &= 0\ \text{m} \cdot \text{s}^{-2} \end{aligned} \tag{15.32}$$

Lesedi has an acceleration of $0\,\text{m}\cdot\text{s}^{-2}$. You will notice that the graph of a vs.t is a horisontal line corresponding to an acceleration value of $0\,\text{m}\cdot\text{s}^{-2}$. There is no acceleration during the motion because his velocity does not change.

We can use the v vs. t to calculate the displacement by finding the area under the graph.

$$\begin{aligned} v &= \text{Area under graph} \\ &= \ell \times b \\ &= 100 \times (-1) \\ &= -100 \text{ m} \end{aligned} \tag{15.33}$$

This means that Lesedi has a displacement of $100\,\mathrm{m}$ towards his house.

Velocity and acceleration

1. Use the graphs in Figure 15.24 to calculate each of the following:
   a. Calculate Lesedi's velocity between $50\,\mathrm{s}$ and $100\,\mathrm{s}$ using the x vs. t graph. Hint: Find the gradient of the line.
   b. Calculate Lesedi's acceleration during the whole motion using the v vs. t graph.
   c. Calculate Lesedi's displacement during the whole motion using the v vs. t graph.
2. Thandi takes $200\,\mathrm{s}$ to walk $100\,\mathrm{m}$ to the bus stop every morning. In the evening Thandi takes $200\,\mathrm{s}$ to walk $100\,\mathrm{m}$ from the bus stop to her home.
   a. Draw a graph of Thandi's position as a function of time for the morning (assuming that Thandi's home is the reference point). Use the gradient of the x vs. t graph to draw the graph of velocity vs. time. Use the gradient of the v vs. t graph to draw the graph of acceleration vs. time.
   b. Draw a graph of Thandi's position as a function of time for the evening (assuming that Thandi's home is the origin). Use the gradient of the x vs. t graph to draw the graph of velocity vs. time. Use the gradient of the v vs. t graph to draw the graph of acceleration vs. time.
   c. Discuss the differences between the two sets of graphs in questions 2 and 3.

www Find the answers with the shortcodes:

(1.) l19 (2.) l19

Experiment : Motion at constant velocity

Aim: To measure the position and time during motion at constant velocity and determine the average velocity as the gradient of a "Position vs. Time" graph.

Apparatus: A battery operated toy car, stopwatch, meter stick or measuring tape.

Method

1. Work with a friend. Copy the table below into your workbook.
2. Complete the table by timing the car as it travels each distance.
3. Time the car twice for each distance and take the average value as your accepted time.
4. Use the distance and average time values to plot a graph of "Distance vs. Time" **onto graph paper**. Stick the graph paper into your workbook. (Remember that "A vs. B" always means "y vs. x").
5. Insert all axis labels and units onto your graph.
6. Draw the best straight line through your data points.

7. Find the gradient of the straight line. This is the average velocity.

Results:

Distance (m)	Time (s)		
	1	2	Ave.
0			
0,5			
1,0			
1,5			
2,0			
2,5			
3,0			

Table 15.3

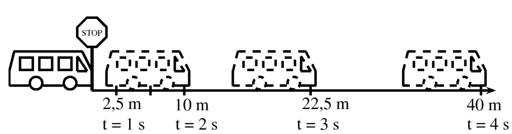
Conclusions: Answer the following questions in your workbook:

1. Did the car travel with a constant velocity?
2. How can you tell by looking at the "Distance vs. Time" graph if the velocity is constant?
3. How would the "Distance vs. Time" look for a car with a faster velocity?
4. How would the "Distance vs. Time" look for a car with a slower velocity?

1.114.3 Motion at Constant Acceleration

The final situation we will be studying is motion at constant acceleration. We know that acceleration is the rate of change of velocity. So, if we have a constant acceleration, this means that the velocity changes at a constant rate.

Let's look at our first example of Lesedi waiting at the taxi stop again. A taxi arrived and Lesedi got in. The taxi stopped at the stop street and then accelerated as follows: After $1\,\mathrm{s}$ the taxi covered a distance of $2,5\,\mathrm{m}$, after $2\,\mathrm{s}$ it covered $10\,\mathrm{m}$, after $3\,\mathrm{s}$ it covered $22,5\,\mathrm{m}$ and after $4\,\mathrm{s}$ it covered $40\,\mathrm{m}$. The taxi is covering a larger distance every second. This means that it is accelerating.



To calculate the velocity of the taxi you need to calculate the gradient of the line at each second:

$$\begin{aligned} v_{1s} &= \frac{\Delta x}{\Delta t} \\ &= \frac{x_f - x_i}{t_f - t_i} \\ &= \frac{5\ \text{m} - 0\ \text{m}}{1{,}5\ \text{s} - 0{,}5\ \text{s}} \\ &= 5\ \text{m} \cdot \text{s}^{-1} \end{aligned} \tag{15.34}$$

$$\begin{aligned} v_{2s} &= \frac{\Delta x}{\Delta t} \\ &= \frac{x_f - x_i}{t_f - t_i} \\ &= \frac{15\ \text{m} - 5\ \text{m}}{2{,}5\ \text{s} - 1{,}5\ \text{s}} \\ &= 10\ \text{m} \cdot \text{s}^{-1} \end{aligned} \tag{15.35}$$

$$\begin{aligned} v_{3s} &= \frac{\Delta x}{\Delta t} \\ &= \frac{x_f - x_i}{t_f - t_i} \\ &= \frac{30\ \text{m} - 15\ \text{m}}{3{,}5\ \text{s} - 2{,}5\ \text{s}} \\ &= 15\ \text{m} \cdot \text{s}^{-1} \end{aligned} \tag{15.36}$$

From these velocities, we can draw the velocity-time graph which forms a straight line.

The acceleration is the gradient of the v vs. t graph and can be calculated as follows:

$$\begin{aligned} a &= \frac{\Delta v}{\Delta t} \\ &= \frac{v_f - v_i}{t_f - t_i} \\ &= \frac{15\ \text{m}\cdot\text{s}^{-1} - 5\ \text{m}\cdot\text{s}^{-1}}{3\ \text{s} - 1\ \text{s}} \\ &= 5\ \text{m} \cdot \text{s}^{-2} \end{aligned} \tag{15.37}$$

The acceleration does not change during the motion (the gradient stays constant). This is motion at constant or uniform acceleration.

The graphs for this situation are shown in Figure 15.27.

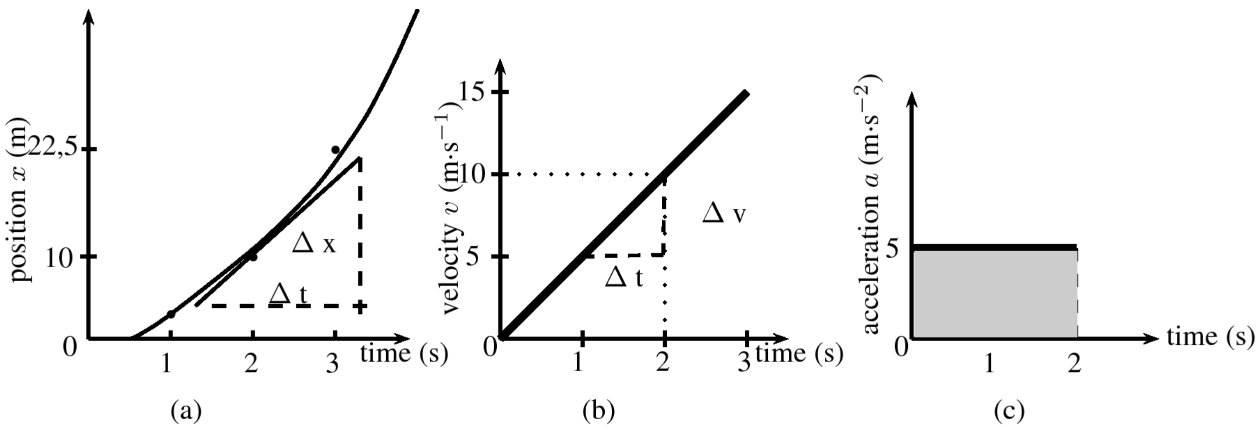


Figure 15.27: Graphs for motion with a constant acceleration (a) position vs. time (b) velocity vs. time (c) acceleration vs. time.

Velocity from Acceleration vs. Time Graphs

Just as we used velocity vs. time graphs to find displacement, we can use acceleration vs. time graphs to find the velocity of an object at a given moment in time. We simply calculate the area under the acceleration vs. time graph, at a given time. In the graph below, showing an object at a constant positive acceleration, the increase in velocity of the object after 2 seconds corresponds to the shaded portion.

$$\begin{aligned} v = \text{area of rectangle} &= a \times \Delta t \\ &= 5\ \text{m} \cdot \text{s}^{-2} \times 2\ \text{s} \\ &= 10\,\text{m} \cdot \text{s}^{-1} \end{aligned} \tag{15.38}$$

The velocity of the object at $t = 2\text{s}$ is therefore $10\,\text{m} \cdot \text{s}^{-1}$. This corresponds with the values obtained in Figure 15.27.

1.114.4 Summary of Graphs

The relation between graphs of position, velocity and acceleration as functions of time is summarised in Figure 15.28.

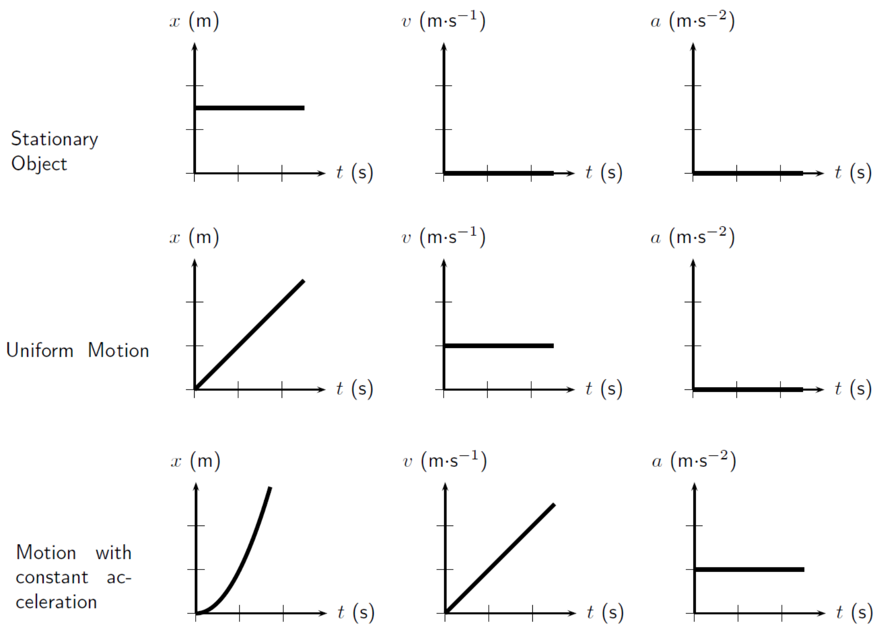


Figure 15.28: Position-time, velocity-time and acceleration-time graphs.

TIP: Often you will be required to describe the motion of an object that is presented as a graph of either position, velocity or acceleration as functions of time. The description of the motion represented by a graph should include the following (where possible):

1. whether the object is moving in the positive or negative direction
2. whether the object is at rest, moving at constant velocity or moving at constant positive acceleration (speeding up) or constant negative acceleration (slowing down)

You will also often be required to draw graphs based on a description of the motion in words or from a diagram. Remember that these are just different methods of presenting the same information. If you keep in mind the general shapes of the graphs for the different types of motion, there should not be any difficulty with explaining what is happening.

1.114.5 Experiment: Position versus time using a ticker timer

Aim: To measure the position and time during motion and to use that data to plot a “Position vs. Time" graph.

Apparatus: Trolley, ticker tape apparatus, tape, graph paper, ruler, ramp

Method:

1. Work with a friend. Copy the table below into your workbook.

2. Attach a length of tape to the trolley.
3. Run the other end of the tape through the ticker timer.
4. Start the ticker timer going and roll the trolley down the ramp.
5. Repeat steps 1 - 3.
6. On each piece of tape, measure the distance between successive dots. Note these distances in the table below.
7. Use the frequency of the ticker timer to work out the time intervals between successive dots. Note these times in the table below,
8. Work out the average values for distance and time.
9. Use the average distance and average time values to plot a graph of “Distance vs. Time" **onto graph paper**. Stick the graph paper into your workbook. (Remember that “A vs. B" always means “y vs. x").
10. Insert all axis labels and units onto your graph.
11. Draw the best straight line through your data points.

Results:

Distance (m)			Time (s)		
1	2	Ave.	1	2	Ave.

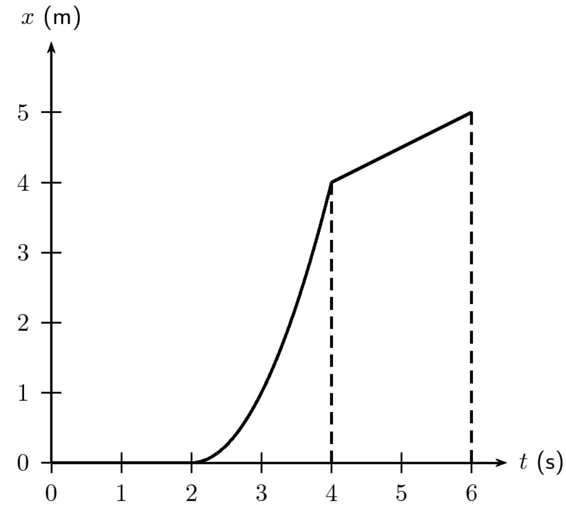
Table 15.4

Discussion: Describe the motion of the trolley down the ramp.

1.114.6 Worked Examples

The worked examples in this section demonstrate the types of questions that can be asked about graphs.

Exercise 15.4: Description of motion based on a position-time graph The position vs. time graph for the motion of a car is given below. Draw the corresponding velocity vs. time and acceleration vs. time graphs, and then describe the motion of the car.



Solution to Exercise

Step 1. The question gives a position vs. time graph and the following three things are required:

a. Draw a v vs. t graph.
b. Draw an a vs. t graph.
c. Describe the motion of the car.

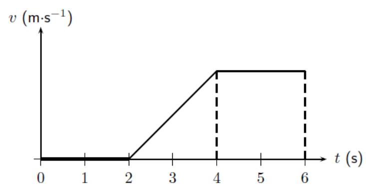
To answer these questions, break the motion up into three sections: 0 – 2 seconds, 2 – 4 seconds and 4 – 6 seconds.

Step 2. For the first 2 seconds we can see that the displacement remains constant - so the object is not moving, thus it has zero velocity during this time. We can reach this conclusion by another path too: remember that the gradient of a displacement vs. time graph is the velocity. For the first 2 seconds we can see that the displacement vs. time graph is a horizontal line, ie. it has a gradient of zero. Thus the velocity during this time is zero and the object is stationary.

Step 3. For the next 2 seconds, displacement is increasing with time so the object is moving. Looking at the gradient of the displacement graph we can see that it is not constant. In fact, the slope is getting steeper (the gradient is increasing) as time goes on. Thus, remembering that the gradient of a displacement vs. time graph is the velocity, the velocity must be increasing with time during this phase.

Step 4. For the final 2 seconds we see that displacement is still increasing with time, but this time the gradient is constant, so we know that the object is now travelling at a constant velocity, thus the velocity vs. time graph will be a horizontal line during this stage. We can now draw the graphs:

So our velocity vs. time graph looks like this one below. Because we haven't been given any values on the vertical axis of the displacement vs. time graph, we cannot figure out what the exact gradients are and therefore what the values of the velocities are. In this type of question it is just important to show whether velocities are positive or negative, increasing, decreasing or constant.

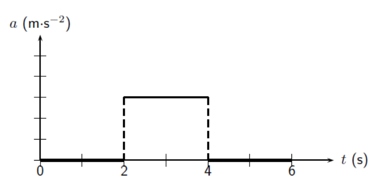


Once we have the velocity vs. time graph its much easier to get the acceleration vs. time graph as we know that the gradient of a velocity vs. time graph is the just the acceleration.

Step 5. For the first 2 seconds the velocity vs. time graph is horisontal and has a value of zero, thus it has a gradient of zero and there is no acceleration during this time. (This makes sense because we know from the displacement time graph that the object is stationary during this time, so it can't be accelerating).

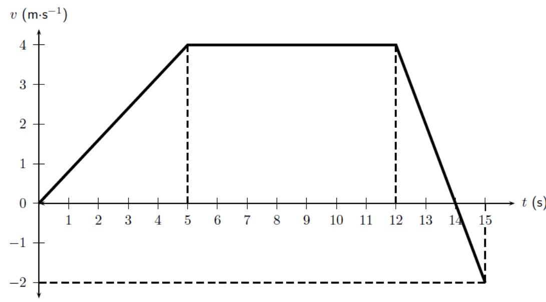
Step 6. For the next 2 seconds the velocity vs. time graph has a positive gradient. This gradient is not changing (i.e. its constant) throughout these 2 seconds so there must be a constant positive acceleration.

Step 7. For the final 2 seconds the object is traveling with a constant velocity. During this time the gradient of the velocity vs. time graph is once again zero, and thus the object is not accelerating. The acceleration vs. time graph looks like this:



Step 8. A brief description of the motion of the object could read something like this: At $t = 0$ s and object is stationary at some position and remains stationary until $t = 2$ s when it begins accelerating. It accelerates in a positive direction for 2 seconds until $t = 4$ s and then travels at a constant velocity for a further 2 seconds.

Exercise 15.5: Calculations from a velocity vs. time graph The velocity vs. time graph of a truck is plotted below. Calculate the distance and displacement of the truck after 15 seconds.



Solution to Exercise

Step 1. We are asked to calculate the distance and displacement of the car. All we need to remember here is that we can use the area between the velocity vs. time graph and the time axis to determine the distance and displacement.

Step 2. Break the motion up: 0 – 5 seconds, 5 – 12 seconds, 12 – 14 seconds and 14 – 15 seconds.

For 0 – 5 seconds: The displacement is equal to the area of the triangle on the left:

$$\begin{aligned} \text{Area}_{\triangle} &= \tfrac{1}{2}\, b \times h \\ &= \tfrac{1}{2} \times 5\ \text{s} \ \times 4\,\text{m}\cdot\text{s}^{-1} \\ &= 10\ \text{m} \end{aligned} \tag{15.39}$$

For 5 – 12 seconds: The displacement is equal to the area of the rectangle:

$$\begin{aligned} \text{Area}_{\square} &= \ell \times b \\ &= 7\ \text{s} \ \times 4\,\text{m}\cdot\text{s}^{-1} \\ &= 28\ \text{m}^2 \end{aligned} \tag{15.40}$$

For 12 – 14 seconds the displacement is equal to the area of the triangle above the time axis on the right:

$$\begin{aligned} \text{Area}_{\triangle} &= \tfrac{1}{2}\, b \times h \\ &= \tfrac{1}{2} \times 2\ \text{s} \ \times 4\,\text{m}\cdot\text{s}^{-1} \\ &= 4\ \text{m} \end{aligned} \tag{15.41}$$

For 14 – 15 seconds the displacement is equal to the area of the triangle below the time axis:

$$\begin{aligned} \text{Area}_{\triangle} &= \tfrac{1}{2}\, b \times h \\ &= \tfrac{1}{2} \times 1\ \text{s} \ \times 2\,\text{m}\cdot\text{s}^{-1} \\ &= 1\ \text{m} \end{aligned} \tag{15.42}$$

Step 3. Now the total distance of the car is the sum of all of these areas:

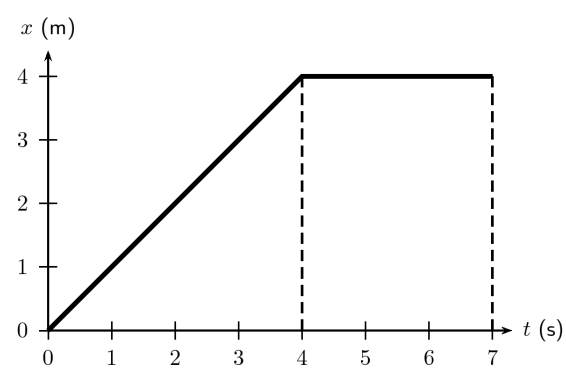
$$\begin{aligned} \Delta x &= 10\ \mathrm{m} + 28\ \mathrm{m} + 4\ \mathrm{m} + 1\ \mathrm{m} \\ &= 43\ \mathrm{m} \end{aligned} \tag{15.43}$$

Step 4. Now the total displacement of the car is just the sum of all of these areas. HOWEVER, because in the last second (from $t = 14$ s to $t = 15$ s) the velocity of the car is negative, it means that the car was going in the opposite direction, i.e. back where it came from! So, to find the total displacement, we have to add the first 3 areas (those with positive displacements) and subtract the last one (because it is a displacement in the opposite direction).

$$\begin{aligned} \Delta x &= 10\ \mathrm{m} + 28\ \mathrm{m} + 4\ \mathrm{m} - 1\ \mathrm{m} \\ &= 41\,\mathrm{min\ the\ positive\ direction} \end{aligned} \tag{15.44}$$

Exercise 15.6: Velocity from a position vs. time graph The position vs. time graph below describes the motion of an athlete.

1. What is the velocity of the athlete during the first 4 seconds?
2. What is the velocity of the athlete from $t = 4$ s to $t = 7$ s?



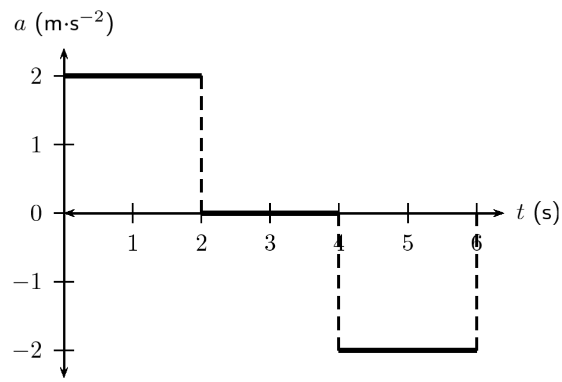
Solution to Exercise

Step 1. The velocity is given by the gradient of a position vs. time graph. During the first 4 seconds, this is

$$\begin{aligned} v &= \frac{\Delta x}{\Delta t} \\ &= \frac{4\ \mathrm{m} - 0\ \mathrm{m}}{4\ \mathrm{s} - 0\ \mathrm{s}} \\ &= 1\ \mathrm{m} \cdot \mathrm{s}^{-1} \end{aligned} \tag{15.45}$$

Step 2. For the last 3 seconds we can see that the displacement stays constant. The graph shows a horisontal line and therefore the gradient is zero. Thus $v = 0\ \mathrm{m} \cdot \mathrm{s}^{-1}$.

Exercise 15.7: Drawing a v vs. t graph from an a vs. t graph The acceleration vs. time graph for a car starting from rest, is given below. Calculate the velocity of the car and hence draw the velocity vs. time graph.



Solution to Exercise

Step 1. The motion of the car can be divided into three time sections: 0 – 2 seconds; 2 – 4 seconds and 4 – 6 seconds. To be able to draw the velocity vs. time graph, the velocity for each time section needs to be calculated. The velocity is equal to the area of the square under the graph:
For 0 – 2 seconds:

$$\begin{aligned} \mathrm{Area}_{\square} &= \ell \times b \\ &= 2\ \mathrm{s}\ \times 2\,\mathrm{m} \cdot \mathrm{s}^{-2} \\ &= 4\ \mathrm{m} \cdot \mathrm{s}^{-1} \end{aligned} \tag{15.46}$$

The velocity of the car is 4 m·s^{-1} at t = 2s. For 2 – 4 seconds:

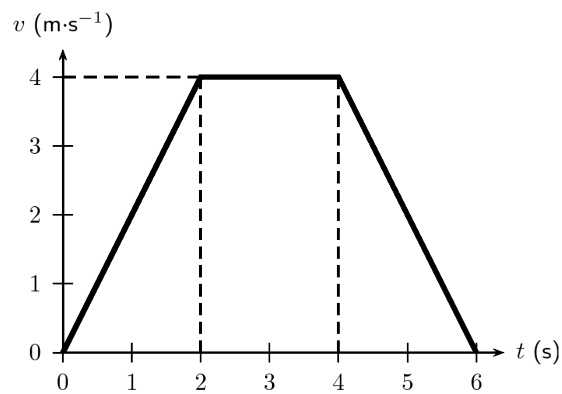
$$\begin{aligned} \text{Area}_{\square} &= \ell \times b \\ &= 2\ \text{s} \times 0\,\text{m}\cdot\text{s}^{-2} \\ &= 0\ \text{m}\cdot\text{s}^{-1} \end{aligned} \tag{15.47}$$

The velocity of the car is 0 m·s^{-1} from $t = 2$ s to $t = 4$ s. For 4 – 6 seconds:

$$\begin{aligned} \text{Area}_{\square} &= \ell \times b \\ &= 2\ \text{s} \times -2\,\text{m}\cdot\text{s}^{-2} \\ &= -4\ \text{m}\cdot\text{s}^{-1} \end{aligned} \tag{15.48}$$

The acceleration had a negative value, which means that the velocity is decreasing. It starts at a velocity of 4 m·s^{-1} and decreases to 0 m·s^{-1}.

Step 2. The velocity vs. time graph looks like this:



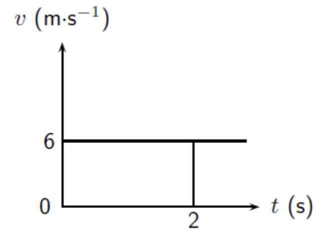
1.114.7 Graphs

1. A car is parked $10\,\text{m}$ from home for 10 minutes. Draw a displacement-time, velocity-time and acceleration-time graphs for the motion. Label all the axes.

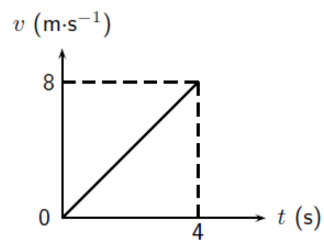
2. A bus travels at a constant velocity of $12\,\text{m}\cdot\text{s}^{-1}$for 6 seconds. Draw the displacement-time, velocity-time and acceleration-time graph for the motion. Label all the axes.

3. An athlete runs with a constant acceleration of $1\,\text{m}\cdot\text{s}^{-2}$ for $4\,\text{s}$. Draw the acceleration-time, velocity-time and displacement time graphs for the motion. Accurate values are only needed for the acceleration-time and velocity-time graphs.

4. The following velocity-time graph describes the motion of a car. Draw the displacement-time graph and the acceleration-time graph and explain the motion of the car according to the three graphs.



5. The following velocity-time graph describes the motion of a truck. Draw the displacement-time graph and the acceleration-time graph and explain the motion of the truck according to the three graphs.



This simulation allows you the opportunity to plot graphs of motion and to see how the graphs of motion change when you move the man. (Simulation: lb8)

Find the answers with the shortcodes:

(1.) l1I (2.) l15 (3.) l1N (4.) l1R (5.) l1n

1.115 Equations of Motion

(section shortcode: P10023)

In this chapter we will look at the third way to describe motion. We have looked at describing motion in terms of graphs and words. In this section we examine equations that can be used to describe motion.

This section is about solving problems relating to uniformly accelerated motion. In other words, motion at constant acceleration.

The following are the variables that will be used in this section:

$$\begin{aligned} v_i &= \text{initial velocity } \left(\mathrm{m \cdot s^{-1}}\right) \text{ at } \mathrm{t = 0s} \\ v_f &= \text{final velocity } \left(\mathrm{m \cdot s^{-1}}\right) \text{ at time } \mathrm{t} \\ \Delta x &= \text{displacement } (\mathrm{m}) \\ t &= \text{time } (\mathrm{s}) \\ \Delta t &= \text{time interval } (\mathrm{s}) \\ a &= \text{acceleration } \left(\mathrm{m \cdot s^{-1}}\right) \end{aligned} \tag{15.49}$$

$$v_f = v_i + at \tag{15.50}$$

$$\Delta x = \frac{(v_i + v_f)}{2} t \tag{15.51}$$

$$\Delta x = v_i t + \frac{1}{2} a t^2 \tag{15.52}$$

$$v_f^2 = v_i^2 + 2a\Delta x \tag{15.53}$$

The questions can vary a lot, but the following method for answering them will always work. Use this when attempting a question that involves motion with constant acceleration. You need any three known quantities (v_i, v_f, Δx, t or a) to be able to calculate the fourth one.

1. Read the question carefully to identify the quantities that are given. Write them down.
2. Identify the equation to use. *Write it down!!!*
3. Ensure that all the values are in the correct unit and fill them in your equation.
4. Calculate the answer and fill in its unit.

NOTE: Galileo Galilei of Pisa, Italy, was the first to determined the correct mathematical law for acceleration: the total distance covered, starting from rest, is proportional to the square of the time. He also concluded that objects retain their velocity unless a force – often friction – acts upon them, refuting the accepted Aristotelian hypothesis that objects "naturally" slow down and stop unless a force acts upon them. This principle was incorporated into Newton’s laws of motion (1st law).

1.115.1 Finding the Equations of Motion

The following does not form part of the syllabus and can be considered additional information.

Derivation of (15.50)

According to the definition of acceleration:

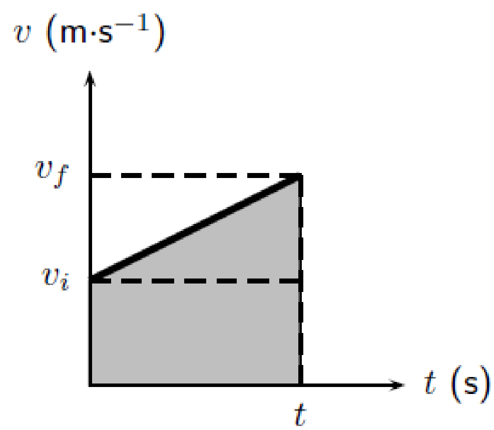
$$a = \frac{\Delta v}{t} \tag{15.54}$$

where Δv is the change in velocity, i.e. $\Delta v = v_f$ - v_i. Thus we have

$$\begin{aligned} a &= \frac{v_f - v_i}{t} \\ v_f &= v_i + at \end{aligned} \tag{15.55}$$

Derivation of (15.51)

We have seen that displacement can be calculated from the area under a velocity vs. time graph. For *uniformly accelerated motion* the most complicated velocity vs. time graph we can have is a straight line. Look at the graph below - it represents an object with a starting velocity of v_i, accelerating to a final velocity v_f over a total time t.



To calculate the final displacement we must calculate the area under the graph - this is just the area of the rectangle added to the area of the triangle. This portion of the graph has been shaded for clarity.

$$\begin{aligned} \text{Area}_{\triangle} &= \frac{1}{2} b \times h \\ &= \frac{1}{2} t \times (v_f - v_i) \\ &= \frac{1}{2} v_f t - \frac{1}{2} v_i t \end{aligned} \tag{15.56}$$

$$\begin{aligned} \text{Area}_{\square} &= \ell \times b \\ &= t \times v_i \\ &= v_i t \end{aligned} \tag{15.57}$$

$$\begin{aligned} \text{Displacement} &= \text{Area}_{\square} + \text{Area}_{\triangle} \\ \Delta x &= v_i t + \frac{1}{2} v_f t - \frac{1}{2} v_i t \\ \Delta x &= \frac{(v_i + v_f)}{2} t \end{aligned} \tag{15.58}$$

Derivation of (15.52)

This equation is simply derived by eliminating the final velocity v_f in (15.51). Remembering from (15.50) that

$$v_f = v_i + at \tag{15.59}$$

then (15.51) becomes

$$\begin{aligned} \Delta x &= \frac{v_i + v_i + at}{2} t \\ &= \frac{2v_i t + at^2}{2} \\ \Delta x &= v_i t + \frac{1}{2} a t^2 \end{aligned} \tag{15.60}$$

Derivation of (15.53)

This equation is just derived by eliminating the time variable in the above equation. From (15.50) we know

$$t = \frac{v_f - v_i}{a} \tag{15.61}$$

Substituting this into (15.52) gives

$$\begin{aligned} \Delta x &= v_i \left(\frac{v_f - v_i}{a}\right) + \frac{1}{2} a \left(\frac{v_f - v_i}{a}\right)^2 \\ &= \frac{v_i v_f}{a} - \frac{v_i^2}{a} + \frac{1}{2} a \left(\frac{v_f^2 - 2v_i v_f + v_i^2}{a^2}\right) \\ &= \frac{v_i v_f}{a} - \frac{v_i^2}{a} + \frac{v_f^2}{2a} - \frac{v_i v_f}{a} + \frac{v_i^2}{2a} \\ 2a\Delta x &= -2v_i^2 + v_f^2 + v_i^2 \\ v_f^2 &= v_i^2 + 2a\Delta x \end{aligned} \tag{15.62}$$

This gives us the final velocity in terms of the initial velocity, acceleration and displacement and is independent of the time variable.

Exercise 15.8: Equations of motion A racing car is travelling north. It accelerates uniformly covering a distance of 725 m in 10 s. If it has an initial velocity of 10 m·s^{-1}, find its acceleration.

Solution to Exercise

Step 1. We are given:

$$\begin{aligned} v_i &= 10\ \mathrm{m \cdot s^{-1}} \\ \Delta x &= 725\ \mathrm{m} \\ t &= 10\ \mathrm{s} \\ a &= ? \end{aligned} \tag{15.63}$$

Step 2. If you struggle to find the correct equation, find the quantity that is not given and then look for an equation that has this quantity in it.
We can use equation (15.52)

$$\Delta x = v_i t + \frac{1}{2} a t^2 \tag{15.64}$$

Step 3.

$$\begin{aligned} \Delta x &= v_i t + \tfrac{1}{2} a t^2 \\ 725\ \mathrm{m} &= \left(10\,\mathrm{m}\cdot\mathrm{s}^{-1} \times 10\ \mathrm{s}\right) + \tfrac{1}{2}\mathrm{a} \times (10\ \mathrm{s})^2 \\ 725\ \mathrm{m} - 100\ \mathrm{m} &= \left(50\ \mathrm{s}^2\right)\ \mathrm{a} \\ a &= 12{,}5\ \mathrm{m}\cdot\mathrm{s}^{-2} \end{aligned} \tag{15.65}$$

Step 4. The racing car is accelerating at 12,5 m·s^{-2} north.

Exercise 15.9: Equations of motion A motorcycle, travelling east, starts from rest, moves in a straight line with a constant acceleration and covers a distance of 64 m in 4 s. Calculate

1. its acceleration
2. its final velocity
3. at what time the motorcycle had covered half the total distance
4. what distance the motorcycle had covered in half the total time.

Solution to Exercise

Step 1. We are given:

$$\begin{aligned} v_i &= 0\mathrm{m}\cdot\mathrm{s}^{-1}\ (\text{because the object starts from rest.}) \\ \Delta x &= 64\mathrm{m} \\ t &= 4\mathrm{s} \\ a &= ? \\ v_f &= ? \\ t &= ?\text{at half the distance}\Delta\mathrm{x}\ =\ 32\ \mathrm{m.} \\ \Delta x &= ?\text{at half the time t}\ =\ 2\ \mathrm{s.} \end{aligned} \tag{15.66}$$

All quantities are in SI units.

Step 2. We can use (15.52)

$$\Delta x = v_i t + \frac{1}{2} a t^2 \tag{15.67}$$

Step 3.

$$\begin{aligned} \Delta x &= v_i t + \tfrac{1}{2} a t^2 \\ 64\ \mathrm{m} &= \left(0 m \cdot s^{-1} \times 4\ \mathrm{s}\right) + \tfrac{1}{2} a \times (4\ \mathrm{s})^2 \\ 64\ \mathrm{m} &= \left(8\ \mathrm{s}^2\right) \mathrm{a} \\ a &= 8\, m \cdot s^{-2}\ \mathrm{east} \end{aligned} \tag{15.68}$$

Step 4. We can use (15.53) - remember we now also know the acceleration of the object.

$$v_f = v_i + at \tag{15.69}$$

Step 5.

$$\begin{aligned} v_f &= v_i + at \\ v_f &= 0 \mathrm{m} \cdot \mathrm{s}^{-1} + \left(8 \mathrm{m} \cdot \mathrm{s}^{-2}\right)(4\ \mathrm{s}) \\ &= 32\, \mathrm{m} \cdot \mathrm{s}^{-1}\ \mathrm{east} \end{aligned} \tag{15.70}$$

Step 6. We can use (15.52):

$$\begin{aligned} \Delta x &= v_i + \tfrac{1}{2} a t^2 \\ 32\ \mathrm{m} &= \left(0 \mathrm{m} \cdot \mathrm{s}^{-1}\right) t + \tfrac{1}{2}\left(8 \mathrm{m} \cdot \mathrm{s}^{-2}\right)(t)^2 \\ 32\ \mathrm{m} &= 0 + \left(4 \mathrm{m} \cdot \mathrm{s}^{-2}\right) t^2 \\ 8\ \mathrm{s}^2 &= t^2 \\ t &= 2,83\ \mathrm{s} \end{aligned} \tag{15.71}$$

Step 7. Half the time is 2 s, thus we have v_i, a and t - all in the correct units. We can use (15.52) to get the distance:

$$\begin{aligned} \Delta x &= v_i t + \tfrac{1}{2} a t^2 \\ &= (0)(2) + \tfrac{1}{2}(8)(2)^2 \\ &= 16\, \mathrm{m\, east} \end{aligned} \tag{15.72}$$

Step 8.
a. The acceleration is $8\, \mathrm{m} \cdot \mathrm{s}^{-2}$ east
b. The velocity is $32\, \mathrm{m} \cdot \mathrm{s}^{-1}$ east
c. The time at half the distance is $2,83\, \mathrm{s}$
d. The distance at half the time is $16\, \mathrm{m}$ east

Equations of motion

1. A car starts off at 10 m·s^{-1} and accelerates at 1 m·s^{-2} for 10 s. What is its final velocity?

2. A train starts from rest, and accelerates at 1 m·s^{-2} for 10 s. How far does it move?

3. A bus is going 30 m·s^{-1} and stops in 5 s. What is its stopping distance for this speed?

4. A racing car going at 20 m·s^{-1} stops in a distance of 20 m. What is its acceleration?

5. A ball has a uniform acceleration of 4 m·s^{-1}. Assume the ball starts from rest. Determine the velocity and displacement at the end of 10 s.

6. A motorcycle has a uniform acceleration of 4 m·s^{-1}. Assume the motorcycle has an initial velocity of 20 m·s^{-1}. Determine the velocity and displacement at the end of 12 s.

7. An aeroplane accelerates uniformly such that it goes from rest to 144 km·hr^{-1}in 8 s. Calculate the acceleration required and the total distance that it has traveled in this time.

Find the answers with the shortcodes:

(1.) l1Q (2.) l1U (3.) l1P (4.) l1E (5.) l1m (6.) l1y
(7.) l1V

1.116 Applications in the Real-World

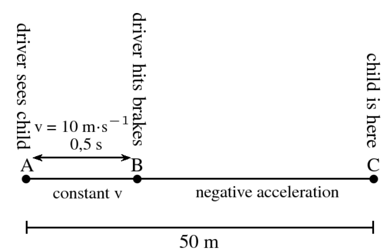
(section shortcode: P10024)

What we have learnt in this chapter can be directly applied to road safety. We can analyse the relationship between speed and stopping distance. The following worked example illustrates this application.

Exercise 15.10: Stopping distance A truck is travelling at a constant velocity of 10 m·s^{-1}when the driver sees a child 50 m in front of him in the road. He hits the brakes to stop the truck. The truck accelerates at a rate of -1.25 m·s^{-2}. His reaction time to hit the brakes is 0,5 seconds. Will the truck hit the child?

Solution to Exercise

Step 1. It is useful to draw a timeline like this one:



We need to know the following:

- What distance the driver covers before hitting the brakes.
- How long it takes the truck to stop after hitting the brakes.
- What total distance the truck covers to stop.

Step 2. Before the driver hits the brakes, the truck is travelling at constant velocity. There is no acceleration and therefore the equations of motion are not used. To find the distance traveled, we use:

$$\begin{aligned} v &= \frac{D}{t} \\ 10 &= \frac{d}{0{,}5} \\ d &= 5 \text{ m} \end{aligned} \tag{15.73}$$

The truck covers 5 m before the driver hits the brakes.

Step 3. We have the following for the motion between B and C:

$$\begin{aligned} v_i &= 10\,\mathrm{m}\cdot\mathrm{s}^{-1} \\ v_f &= 0\,\mathrm{m}\cdot\mathrm{s}^{-1} \\ a &= -1,25\,\mathrm{m}\cdot\mathrm{s}^{-2} \\ t &= ? \end{aligned} \tag{15.74}$$

We can use (3.50)

$$\begin{aligned} v_f &= v_i + at \\ 0 &= 10 + (-1,25)\,t \\ -10 &= -1,25t \\ t &= 8 \text{ s} \end{aligned} \tag{15.75}$$

Step 4. For the distance we can use (15.51) or (15.52). We will use (15.51):

$$\begin{aligned} \Delta x &= \frac{(v_i+v_f)}{2}t \\ \Delta x &= \frac{10+0}{s}\,(8) \\ \Delta x &= 40 \text{ m} \end{aligned} \tag{15.76}$$

Step 5. The total distance that the truck covers is D_{AB} + $D_{BC} = 5 + 40 = 45$ meters. The child is 50 meters ahead. The truck will not hit the child.

1.117 Summary

(section shortcode: P10025)

- A reference point is a point from where you take your measurements.

- A frame of reference is a reference point with a set of directions.
- Your position is where you are located with respect to your reference point.
- The displacement of an object is how far it is from the reference point. It is the shortest distance between the object and the reference point. It has magnitude and direction because it is a vector.
- The distance of an object is the length of the path travelled from the starting point to the end point. It has magnitude only because it is a scalar.
- A vector is a physical quantity with magnitude and direction.
- A scalar is a physical quantity with magnitude only.
- Speed (s) is the distance covered (D) divided by the time taken (Δt):

$$s = \frac{D}{\Delta t} \tag{15.77}$$

- Average velocity (v) is the displacement (Δx) divided by the time taken (Δt):

$$v = \frac{\Delta x}{\Delta t} \tag{15.78}$$

- Instantaneous speed is the speed at a specific instant in time.
- Instantaneous velocity is the velocity at a specific instant in time.
- Acceleration (a) is the change in velocity (Δx) over a time interval (Δt):

$$a = \frac{\Delta v}{\Delta t} \tag{15.79}$$

- The gradient of a position - time graph (x vs. t) give the velocity.
- The gradient of a velocity - time graph (v vs. t) give the acceleration.
- The area under a velocity - time graph (v vs. t) give the displacement.
- The area under an acceleration - time graph (a vs. t) gives the velocity.
- The graphs of motion are summarised in Figure 15.28.
- The equations of motion are used where constant acceleration takes place:

$$\begin{aligned} v_f &= v_i + at \\ \Delta x &= \frac{(v_i+v_f)}{2}t \\ \Delta x &= v_i t + \frac{1}{2}at^2 \\ v_f^2 &= v_i^2 + 2a\Delta x \end{aligned} \tag{15.80}$$

1.118 End of Chapter Exercises: Motion in One Dimension

(section shortcode: P10026)

1. Give one word/term for the following descriptions.
   a. The shortest path from start to finish.
   b. A physical quantity with magnitude and direction.
   c. The quantity defined as a change in velocity over a time period.
   d. The point from where you take measurements.
   e. The distance covered in a time interval.
   f. The velocity at a specific instant in time.

2. Choose an item from column B that match the description in column A. Write down only the letter next to the question number. You may use an item from column B more than once.

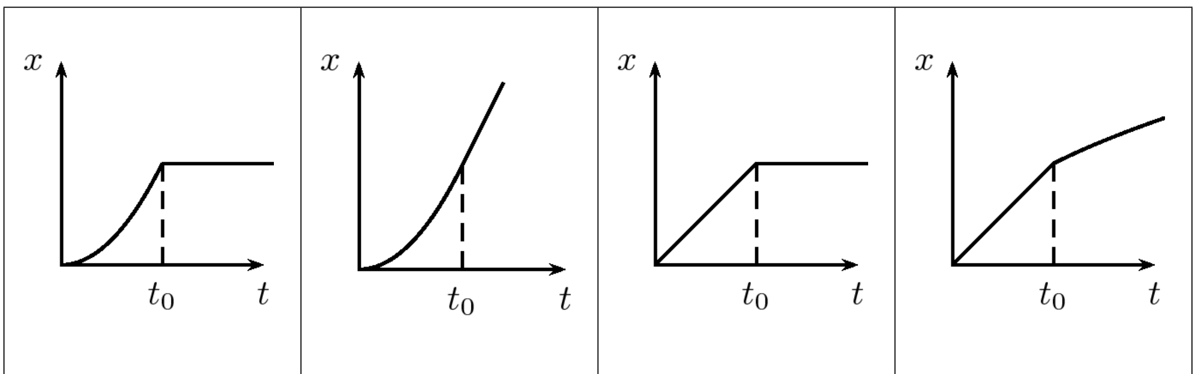
Column A	Column B
a. The area under a velocity - time graph	gradient
b. The gradient of a velocity - time graph	area
c. The area under an acceleration - time graph	velocity
d. The gradient of a displacement - time graph	displacement
	acceleration
	slope

Table 15.5

3. Indicate whether the following statements are TRUE or FALSE. Write only 'true' or 'false'. If the statement is false, write down the correct statement.

a. A scalar is the displacement of an object over a time interval.
b. The position of an object is where it is located.
c. The sign of the velocity of an object tells us in which direction it is travelling.
d. The acceleration of an object is the change of its displacement over a period in time.

4. (SC 2003/11) A body accelerates uniformly from rest for t_0 seconds after which it continues with a constant velocity. Which graph is the correct representation of the body's motion?



continued on next page

(a)	(b)	(c)	(d)

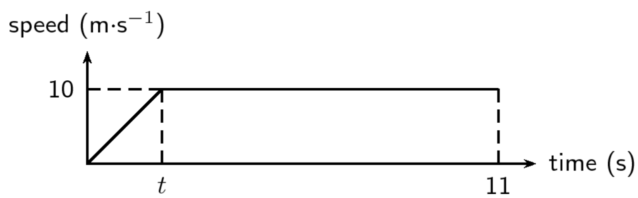
Table 15.6

5. (SC 2003/11) The velocity-time graphs of two cars are represented by P and Q as shown

The difference in the distance travelled by the two cars (in m) after 4 s is ...

a. 12
b. 6
c. 2
d. 0

6. (IEB 2005/11 HG) The graph that follows shows how the speed of an athlete varies with time as he sprints for 100 m.



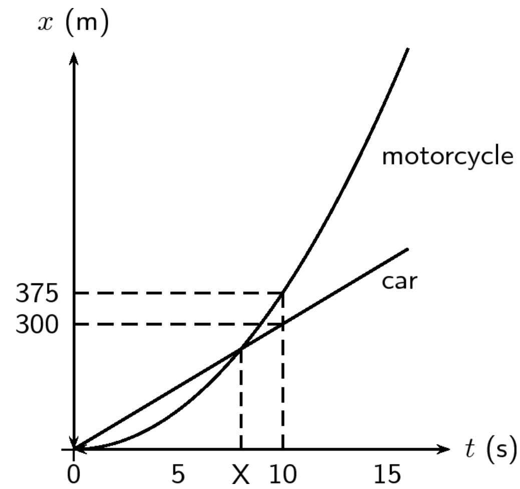
Which of the following equations can be used to correctly determine the time t for which he accelerates?

a. $100 = (10)(11) - \frac{1}{2}(10)t$
b. $100 = (10)(11) + \frac{1}{2}(10)t$
c. $100 = 10t + \frac{1}{2}(10)t^2$
d. $100 = \frac{1}{2}(0)t + \frac{1}{2}(10)t^2$

7. (SC 2002/03 HG1) In which one of the following cases will the distance covered and the magnitude of the displacement be the same?

a. A girl climbs a spiral staircase.
b. An athlete completes one lap in a race.
c. A raindrop falls in still air.
d. A passenger in a train travels from Cape Town to Johannesburg.

8. (SC 2003/11) A car, travelling at constant velocity, passes a stationary motor cycle at a traffic light. As the car overtakes the motorcycle, the motorcycle accelerates uniformly from rest for 10 s. The following displacement-time graph represents the motions of both vehicles from the traffic light onwards.



a. Use the graph to find the magnitude of the constant velocity of the car.

b. Use the information from the graph to show by means of calculation that the magnitude of the acceleration of the motorcycle, for the first 10 s of its motion is 7,5 m·s^{-2}.

c. Calculate how long (in seconds) it will take the motorcycle to catch up with the car (point X on the time axis).

d. How far behind the motorcycle will the car be after 15 seconds?

9. (IEB 2005/11 HG) Which of the following statements is **true** of a body that accelerates uniformly?

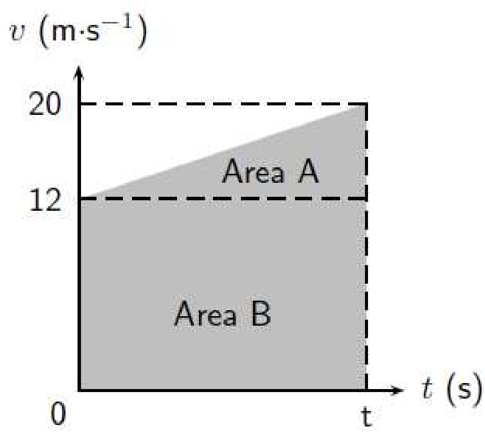
a. Its rate of change of position with time remains constant.

b. Its position changes by the same amount in equal time intervals.

c. Its velocity increases by increasing amounts in equal time intervals.

d. Its rate of change of velocity with time remains constant.

10. (IEB 2003/11 HG1) The velocity-time graph for a car moving along a straight horizontal road is shown below.



Which of the following expressions gives the magnitude of the average velocity of the car?

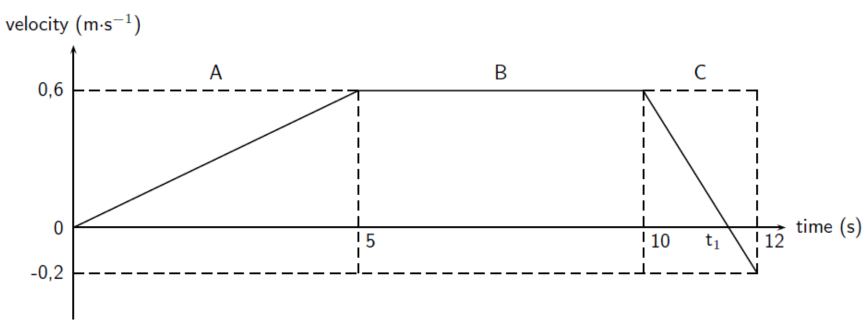
a. $\frac{\text{AreaA}}{t}$

b. $\frac{\text{AreaA + AreaB}}{t}$

c. $\frac{\text{AreaB}}{t}$

d. $\frac{\text{AreaA} - \text{AreaB}}{t}$

11. (SC 2002/11 SG) A car is driven at 25 m·s^{-1} in a municipal area. When the driver sees a traffic officer at a speed trap, he realises he is travelling too fast. He immediately applies the brakes of the car while still 100 m away from the speed trap.
    a. Calculate the magnitude of the minimum acceleration which the car must have to avoid exceeding the speed limit, if the municipal speed limit is 16.6 m·s^{-1}.
    b. Calculate the time from the instant the driver applied the brakes until he reaches the speed trap. Assume that the car's velocity, when reaching the trap, is 16.6 m·s^{-1}.
12. A traffic officer is watching his speed trap equipment at the bottom of a valley. He can see cars as they enter the valley 1 km to his left until they leave the valley 1 km to his right. Nelson is recording the times of cars entering and leaving the valley for a school project. Nelson notices a white Toyota enter the valley at 11:01:30 and leave the valley at 11:02:42. Afterwards, Nelson hears that the traffic officer recorded the Toyota doing 140 km·hr^{-1}.
    a. What was the time interval (Δt) for the Toyota to travel through the valley?
    b. What was the average speed of the Toyota?
    c. Convert this speed to km·hr^{-1}.
    d. Discuss whether the Toyota could have been travelling at 140km·hr^{-1} at the bottom of the valley.
    e. Discuss the differences between the instantaneous speed (as measured by the speed trap) and average speed (as measured by Nelson).
13. (IEB 2003/11HG) A velocity-time graph for a ball rolling along a track is shown below. The graph has been divided up into 3 sections, A, B and C for easy reference. (Disregard any effects of friction.)



    a. Use the graph to determine the following:
        i. the speed 5 s after the start
        ii. the distance travelled in Section A
        iii. the acceleration in Section C
    b. At time t_1 the velocity-time graph intersects the time axis. Use an appropriate equation of motion to calculate the value of time t_1 (in s).
    c. Sketch a displacement-time graph for the motion of the ball for these 12 s. (You do not need to calculate the actual values of the displacement for each time interval, but do pay attention to the general shape of this graph during each time interval.)
14. In towns and cities, the speed limit is 60 km·hr^{-1}. The length of the average car is 3.5 m, and the width of the average car is 2 m. In order to cross the road, you need to be able to walk further than the width of a car, before that car reaches you. To cross safely, you should be able to walk at least 2 m further than the width of the car (4 m in total), before the car reaches you.
    a. If your walking speed is 4 km·hr^{-1}, what is your walking speed in m·s^{-1}?
    b. How long does it take you to walk a distance equal to the width of the average car?

c. What is the speed in m·s^{-1} of a car travelling at the speed limit in a town?
d. How many metres does a car travelling at the speed limit travel, in the same time that it takes you to walk a distance equal to the width of car?
e. Why is the answer to the previous question important?
f. If you see a car driving toward you, and it is 28 m away (the same as the length of 8 cars), is it safe to walk across the road?
g. How far away must a car be, before you think it might be safe to cross? How many car-lengths is this distance?

15. A bus on a straight road starts from rest at a bus stop and accelerates at 2 m·s^{-2} until it reaches a speed of 20 m·s^{-1}. Then the bus travels for 20 s at a constant speed until the driver sees the next bus stop in the distance. The driver applies the brakes, stopping the bus in a uniform manner in 5 s.
   a. How long does the bus take to travel from the first bus stop to the second bus stop?
   b. What is the average velocity of the bus during the trip?

www Find the answers with the shortcodes:

(1.) lrr	(2.) lrY	(3.) lrg	(4.) lr4	(5.) lr2	(6.) lrT
(7.) lrb	(8.) lrj	(9.) lrD	(10.) lrW	(11.) lrZ	(12.) lrB
(13.) lrK	(14.) lrk	(15.) lr0			

Transverse Pulses

1.119 Introduction

(section shortcode: P10027)

This chapter forms the basis of the discussion into mechanical waves. Waves are all around us, even though most of us are not aware of it. The most common waves are waves in the sea, but waves can be created in any container of water, ranging from an ocean to a tea-cup. Waves do not only occur in water, they occur in any kind of medium. Earthquakes generate waves that travel through the rock of the Earth. When your friend speaks to you he produces *sound waves* that travel through the air to your ears. Light is made up of electromagnetic waves. A wave is simply moving energy.

1.120 What is a *medium*?

(section shortcode: P10028)

In this chapter, as well as in the following chapters, we will speak about waves moving in a medium. A medium is just the substance or material through which waves move. In other words the medium carries the wave from one place to another. The medium does not create the wave and the medium is not the wave. Therefore the medium does not travel with the wave as the wave propagates through it. Air is a medium for sound waves, water is a medium for water waves and rock is a medium for earthquakes (which are also a type of wave). Air, water and rock are therefore examples of media (media is the plural of medium).

Definition: Medium
A medium is the substance or material in which a wave will move.

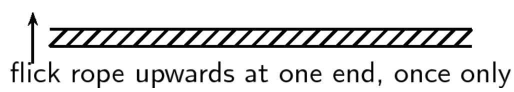
In each medium, the atoms that make up the medium are moved *temporarily* from their rest position. In order for a wave to travel, the different parts of the medium must be able to interact with each other.

1.121 What is a *pulse*?

(section shortcode: P10029)

1.121.1 Investigation : Observation of Pulses

Take a heavy rope. Have two people hold the rope stretched out horizontally. Flick the rope at one end only once.



What happens to the disturbance that you created in the rope? Does it stay at the place where it was created or does it move down the length of the rope?

In the activity, we created a *pulse*. A pulse is a *single* disturbance that moves through a medium. In a transverse pulse the displacement of the medium is perpendicular to the direction of motion of the pulse. Figure 16.2 shows an example of a transverse pulse. In the activity, the rope or spring was held horizontally and the pulse moved the rope up and down. This was an example of a transverse pulse.

Definition: Pulse
A pulse is a single disturbance that moves through a medium.

Definition: Transverse Pulse
A pulse where all of the particles disturbed by the pulse move perpendicular (at a right angle) to the direction in which the pulse is moving.

1.121.2 Pulse Length and Amplitude

The amplitude of a pulse is a measurement of how far the medium is displaced momentarily from a position of rest. The pulse length is a measurement of how long the pulse is. Both these quantities are shown in Figure 16.2.

Definition: Amplitude
The amplitude of a pulse is a measurement of how far the medium is displaced from rest.

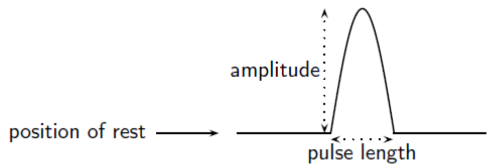
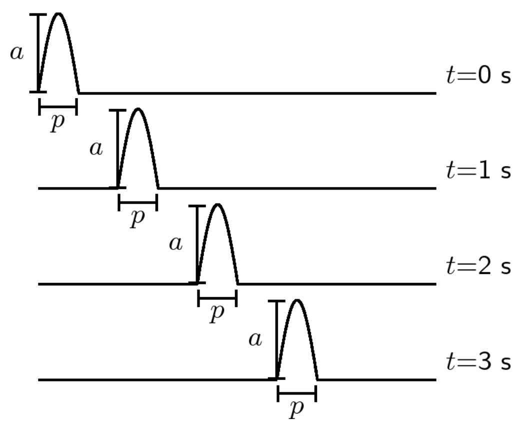


Figure 16.2: Example of a transverse pulse

The position of rest is the position the medium would be in if it were undisturbed. This is also called the equilibrium position. Sometimes people will use rest and sometimes equilibrium but they will also use to the two in the same discussion to mean the same thing.

Investigation : Pulse Length and Amplitude

The graphs below show the positions of a pulse at different times.



Use your ruler to measure the lengths of a and p. Fill your answers in the table.

Time	a	p
$t = 0$ s		
$t = 1$ s		
$t = 2$ s		
$t = 3$ s		

Table 16.1

What do you notice about the values of a and p?

In the activity, we found that the values for how high the pulse (a) is and how wide the pulse (p) is the same at different times. *Pulse length* and *amplitude* are two important quantities of a pulse.

1.121.3 Pulse Speed

Definition: Pulse Speed
Pulse speed is the distance a pulse travels per unit time.

In Motion in one dimension (Chapter 15) we saw that speed was defined as the distance traveled per unit time. We can use the same definition of speed to calculate how fast a pulse travels. If the pulse travels a distance D in a time t, then the pulse speed v is:

$$v = \frac{D}{t} \tag{16.1}$$

Exercise 16.1: Pulse Speed A pulse covers a distance of $2\ \mathrm{m}$ in $4\ \mathrm{s}$ on a heavy rope. Calculate the pulse speed.

Solution to Exercise

Step 1. We are given:

- the distance travelled by the pulse: $D = 2\,\mathrm{m}$
- the time taken to travel $2\,\mathrm{m}$: $t = 4\,\mathrm{s}$

We are required to calculate the speed of the pulse.

Step 2. We can use:

$$v = \frac{D}{t} \tag{16.2}$$

to calculate the speed of the pulse.

Step 3.

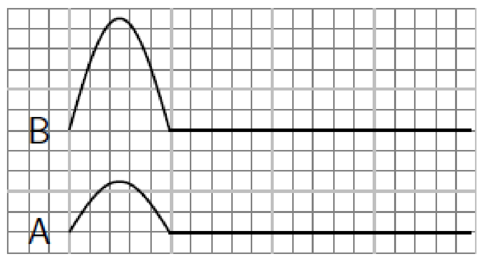
$$\begin{aligned} v &= \tfrac{D}{t} \\ &= \tfrac{2\,\mathrm{m}}{4\,\mathrm{s}} \\ &= 0,5\,\mathrm{m}\cdot\mathrm{s}^{-1} \end{aligned} \tag{16.3}$$

Step 4. The pulse speed is $0,5\,\mathrm{m}\cdot\mathrm{s}^{-1}$.

TIP: The pulse speed depends on the properties of the medium and not on the amplitude or pulse length of the pulse.

Pulse Speed

1. A pulse covers a distance of $5\,\mathrm{m}$ in $15\,\mathrm{s}$. Calculate the speed of the pulse.
2. A pulse has a speed of $5\,\mathrm{cm}\cdot\mathrm{s}^{-1}$. How far does it travel in $2,5\,\mathrm{s}$?
3. A pulse has a speed of $0,5\,\mathrm{m}\cdot\mathrm{s}^{-1}$. How long does it take to cover a distance of $25\,\mathrm{cm}$?
4. How long will it take a pulse moving at $0,25\,\mathrm{m}\cdot\mathrm{s}^{-1}$ to travel a distance of $20\,\mathrm{m}$?
5. The diagram shows two pulses in the same medium. Which has the higher speed? Explain your answer.



Find the answers with the shortcodes:

(1.) l1f (2.) l1G (3.) l17 (4.) l1A (5.) l1o

1.122 Superposition of Pulses

(section shortcode: P10030)

Two or more pulses can pass through the same medium at that same time in the same place. When they do they interact with each other to form a different disturbance at that point. The resulting pulse is obtained by using the *principle of superposition*. The principle of superposition states that the effect of the different pulses is the sum of their individual effects. After pulses pass through each other, each pulse continues along its original direction of travel, and their original amplitudes remain unchanged.

Constructive interference takes place when two pulses meet each other to create a larger pulse. The amplitude of the resulting pulse is the sum of the amplitudes of the two initial pulses. This is shown in Figure 16.5.

Definition: Constructive interference
Constructive interference is when two pulses meet, resulting in a bigger pulse.

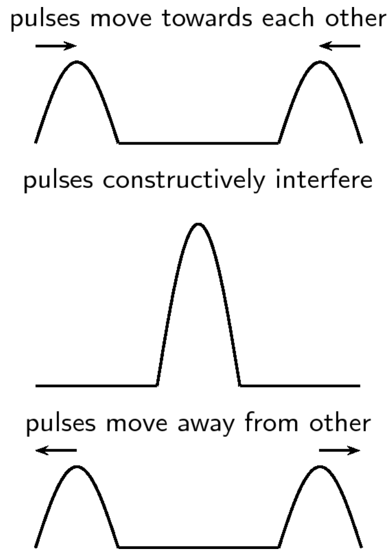


Figure 16.5: Superposition of two pulses: constructive interference.

Destructive interference takes place when two pulses meet and cancel each other. The amplitude of the resulting pulse is the sum of the amplitudes of the two initial pulses, but the one amplitude will be a negative number. This is shown in Figure 16.6. In general, amplitudes of individual pulses add together to give the amplitude of the resultant pulse.

Definition: Destructive interference
Destructive interference is when two pulses meet, resulting in a smaller pulse.

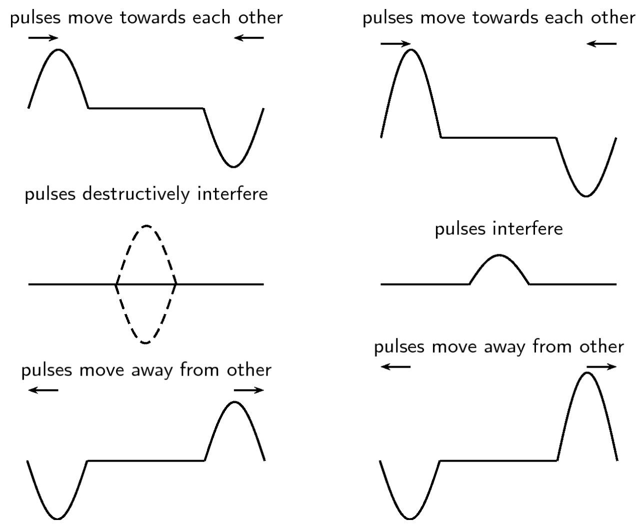
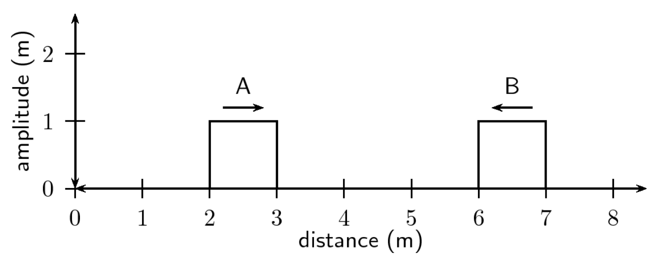


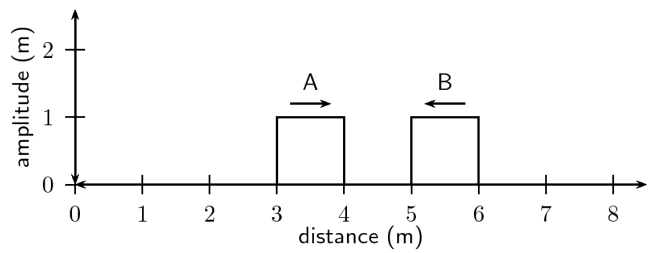
Figure 16.6: Superposition of two pulses. The left-hand series of images demonstrates destructive interference, since the pulses cancel each other. The right-hand series of images demonstrate a partial cancelation of two pulses, as their amplitudes are not the same in magnitude.

Exercise 16.2: Superposition of Pulses The two pulses shown below approach each other at $1\,\mathrm{m}\cdot\mathrm{s}^{-1}$. Draw what the waveform would look like after $1\,\mathrm{s}$, $2\,\mathrm{s}$ and $5\,\mathrm{s}$.

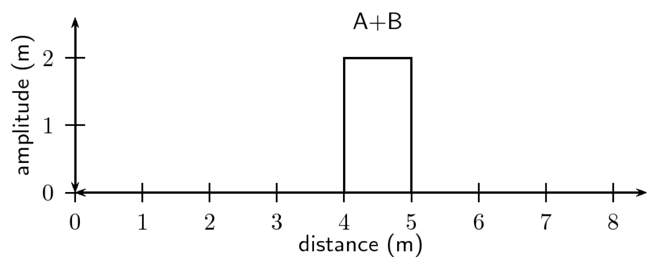


Solution to Exercise

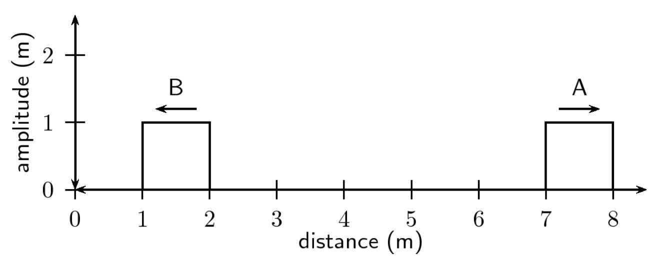
Step 1. After $1\,\mathrm{s}$, pulse A has moved $1\,\mathrm{m}$ to the right and pulse B has moved $1\,\mathrm{m}$ to the left.



Step 2. After $1\ \mathrm{s}$ more, pulse A has moved $1\ \mathrm{m}$ to the right and pulse B has moved $1\ \mathrm{m}$ to the left.



Step 3. After $5\ \mathrm{s}$, pulse A has moved $5\ \mathrm{m}$ to the right and pulse B has moved $5\ \mathrm{m}$ to the left.

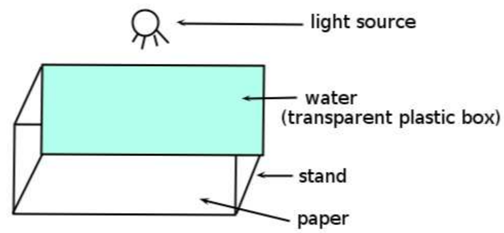


TIP: The idea of superposition is one that occurs often in physics. You will see *much, much more* of superposition!

1.122.1 Experiment: Constructive and destructive interference

Aim: To demonstrate constructive and destructive interference

Apparatus: Ripple tank apparatus



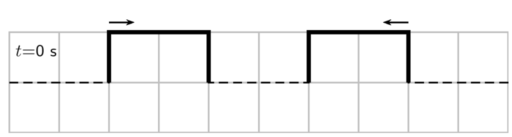
Method:

1. Set up the ripple tank
2. Produce a single pulse and observe what happens
3. Produce two pulses simultaneously and observe what happens
4. Produce two pulses at slightly different times and observe what happens

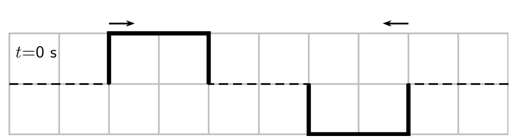
Results and conclusion: You should observe that when you produce two pulses simultaneously you see them interfere constructively and when you produce two pulses at slightly different times you see them interfere destructively.

1.122.2 Problems Involving Superposition of Pulses

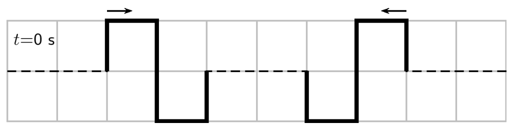
1. For the following pulse, draw the resulting wave forms after $1\,\mathrm{s}$, $2\,\mathrm{s}$, $3\,\mathrm{s}$, $4\,\mathrm{s}$ and $5\,\mathrm{s}$. Each pulse is travelling at $1\,\mathrm{m}\cdot\mathrm{s}^{-1}$. Each block represents $1\,\mathrm{m}$. The pulses are shown as thick black lines and the undisplaced medium as dashed lines.



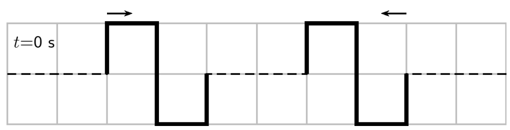
2. For the following pulse, draw the resulting wave forms after $1\,\mathrm{s}$, $2\,\mathrm{s}$, $3\,\mathrm{s}$, $4\,\mathrm{s}$ and $5\,\mathrm{s}$. Each pulse is travelling at $1\,\mathrm{m}\cdot\mathrm{s}^{-1}$. Each block represents $1\,\mathrm{m}$. The pulses are shown as thick black lines and the undisplaced medium as dashed lines.



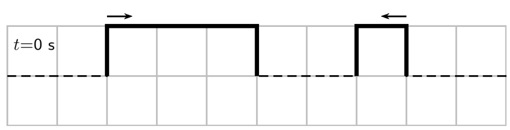
3. For the following pulse, draw the resulting wave forms after $1\,\mathrm{s}$, $2\,\mathrm{s}$, $3\,\mathrm{s}$, $4\,\mathrm{s}$ and $5\,\mathrm{s}$. Each pulse is travelling at $1\,\mathrm{m}\cdot\mathrm{s}^{-1}$. Each block represents $1\,\mathrm{m}$. The pulses are shown as thick black lines and the undisplaced medium as dashed lines.



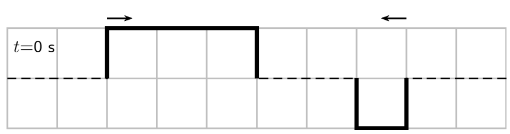
4. For the following pulse, draw the resulting wave forms after $1\,\text{s}$, $2\,\text{s}$, $3\,\text{s}$, $4\,\text{s}$ and $5\,\text{s}$. Each pulse is travelling at $1\,\text{m}\cdot\text{s}^{-1}$. Each block represents $1\,\text{m}$. The pulses are shown as thick black lines and the undisplaced medium as dashed lines.



5. For the following pulse, draw the resulting wave forms after $1\,\text{s}$, $2\,\text{s}$, $3\,\text{s}$, $4\,\text{s}$ and $5\,\text{s}$. Each pulse is travelling at $1\,\text{m}\cdot\text{s}^{-1}$. Each block represents $1\,\text{m}$. The pulses are shown as thick black lines and the undisplaced medium as dashed lines.



6. For the following pulse, draw the resulting wave forms after $1\,\text{s}$, $2\,\text{s}$, $3\,\text{s}$, $4\,\text{s}$ and $5\,\text{s}$. Each pulse is travelling at $1\,\text{m}\cdot\text{s}^{-1}$. Each block represents $1\,\text{m}$. The pulses are shown as thick black lines and the undisplaced medium as dashed lines.



7. What is superposition of waves?

8. What is constructive interference?

9. What is destructive interference?

The following presentation provides a summary of the work covered in this chapter. Although the presentation is titled waves, the presentation covers pulses only.

(Presentation: P10031)

Find the answers with the shortcodes:

(1.) l1M (2.) l1e (3.) l1t (4.) l1z (5.) l1u (6.) l1J
(7.) l1S (8.) l1h (9.) lrg

1.123 Graphs of Position and Velocity (Not in CAPS - Included for Completeness)

(section shortcode: P10032)

When a pulse moves through a medium, there are two different motions: the motion of the particles of the medium and the motion of the pulse. These two motions are at right angles to each other when the pulse is transverse. Each motion will be discussed.

Consider the situation shown in Figure 16.20. The dot represents one particle of the medium. We see that as the pulse moves to the right the particle only moves up and down.

1.123.1 Motion of a Particle of the Medium

First we consider the motion of a particle of the medium when a pulse moves through the medium. For the explanation we will zoom into the medium so that we are looking at the atoms of the medium. These atoms are connected to each other as shown in Figure 16.19.

Figure 16.19: Particles in a medium.

When a pulse moves through the medium, the particles in the medium **only** move up and down. We can see this in Figure 16.20 which shows the motion of a single particle as a pulse moves through the medium.

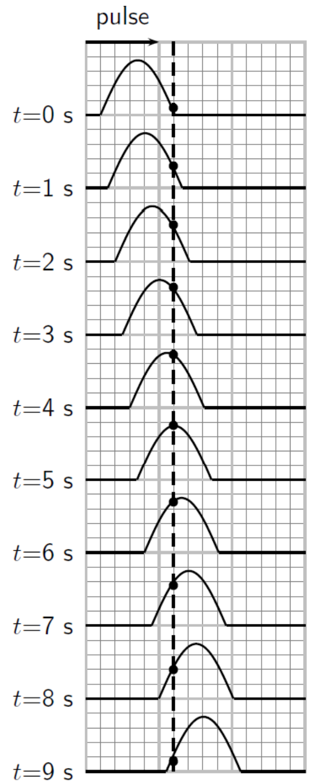


Figure 16.20: Positions of a pulse in a rope at different times. The pulse moves to the right as shown by the arrow. You can also see the motion of a point in the medium through which the pulse is travelling. Each block is 1 cm.

TIP: A particle in the medium **only** moves up and down when a transverse pulse moves through the medium. The pulse moves from left to right (or right to left). The motion of the particle is perpendicular to the motion of a transverse pulse.

If you consider the motion of the particle as a function of time, you can draw a graph of position vs. time and velocity vs. time.

Investigation : Drawing a position-time graph

1. Study Figure 16.20 and complete the following table:

time (s)	0	1	2	3	4	5	6	7	8	9
position (cm)										

Table 16.2

2. Use your table to draw a graph of position vs. time for a particle in a medium.

The position vs. time graph for a particle in a medium when a pulse passes through the medium is shown in Figure 16.21

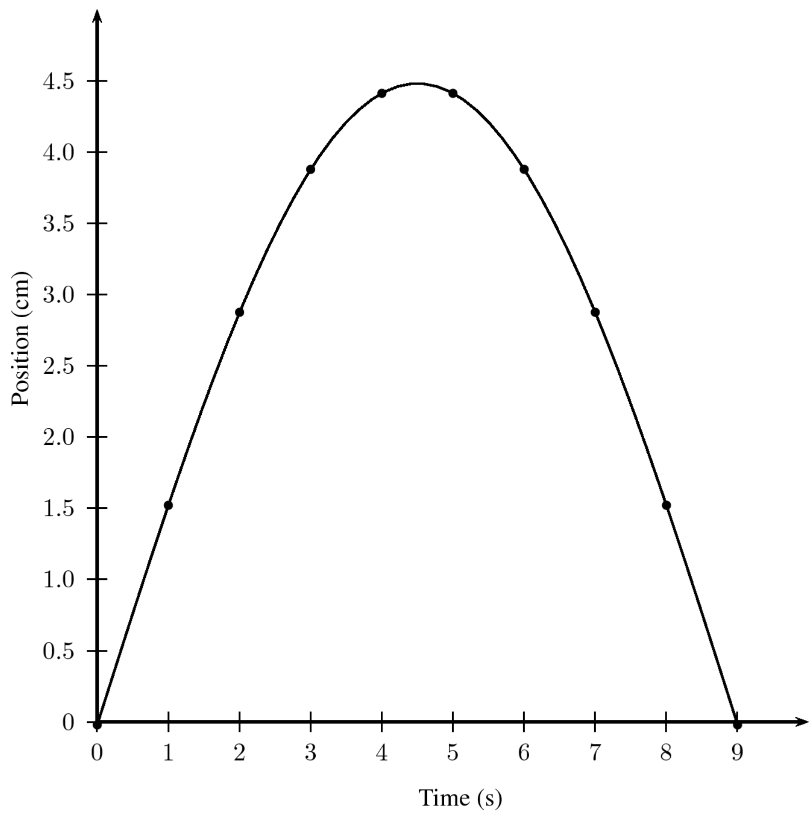


Figure 16.21: Position against Time graph of a particle in the medium through which a transverse pulse is travelling.

Investigation : Drawing a velocity-time graph

1. Study Figure 16.21 and complete the following table:

time (s)	0	1	2	3	4	5	6	7	8	9
velocity (cm.s^{-1})										

Table 16.3

2. Use your table to draw a graph of velocity vs time for a particle in a medium.

The velocity vs. time graph far a particle in a medium when a pulse passes through the medium is shown in Figure 16.22.

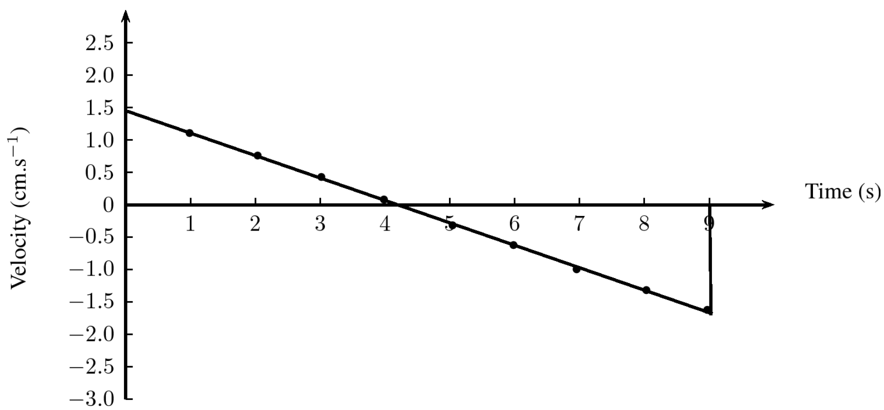


Figure 16.22: Velocity against Time graph of a particle in the medium through which a transverse pulse is travelling.

1.123.2 Motion of the Pulse

The motion of the pulse is much simpler than the motion of a particle in the medium.

TIP: A point on a transverse pulse, eg. the peak, **only** moves in the direction of the motion of the pulse.

Exercise 16.3: Transverse pulse through a medium

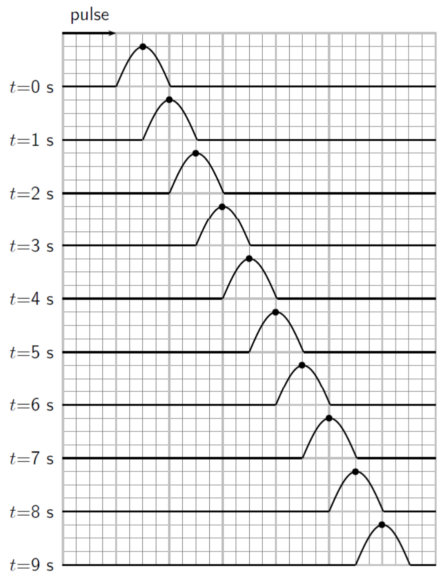


Figure 16.23: Position of the peak of a pulse at different times (since we know the shape of the pulse does not change we can look at only one point on the pulse to keep track of its position, the peak for example). The pulse moves to the right as shown by the arrow. Each square is $0,5\,\mathrm{cm}$.

Given the series of snapshots of a transverse pulse moving through a medium, depicted in Figure 16.23, do the following:

- draw up a table of time, position and velocity,
- plot a position vs. time graph,
- plot a velocity vs. time graph.

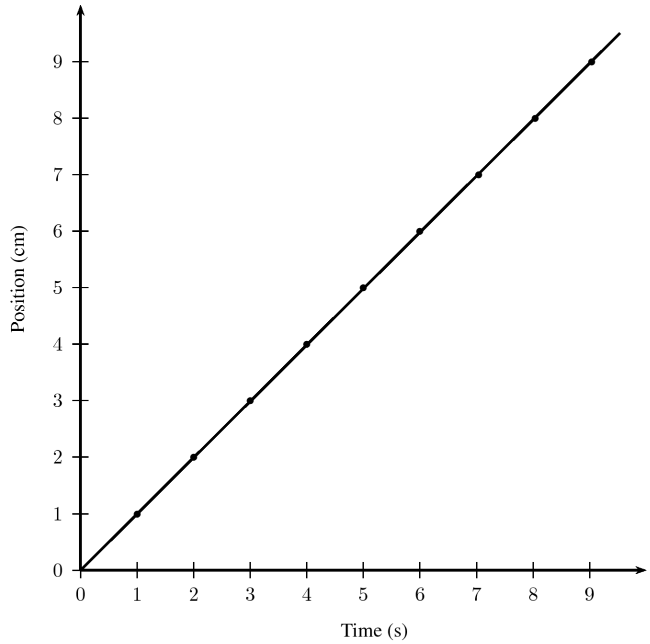
Solution to Exercise

Step 1. Figure 16.23 shows the motion of a pulse through a medium and a dot to indicate the same position on the pulse. If we follow the dot, we can draw a graph of position vs time for a pulse. At $t = 0\,\mathrm{s}$ the dot is at $0\,\mathrm{cm}$. At $t = 1\,\mathrm{s}$ the dot is $1\,\mathrm{cm}$ away from its original postion. At $t = 2\,\mathrm{s}$ the dot is $2\,\mathrm{cm}$ away from its original postion, and so on.

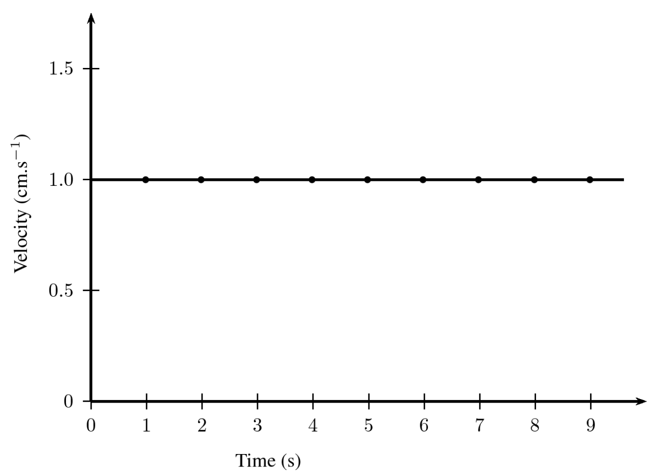
Step 2.

time (s)	0	1	2	3	4	5	6	7	8	9
position (cm)	0	1	2	3	4	5	6	7	8	9
velocity (cm.s^{-1})	1	1	1	1	1	1	1	1	1	1

Table 16.4



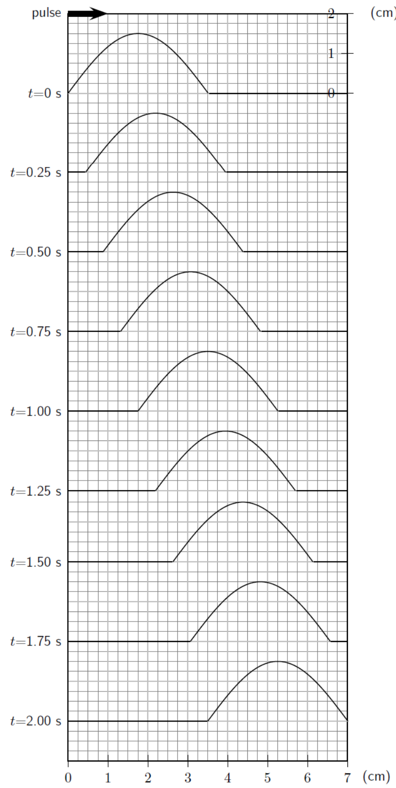
Step 3.



Step 4.

Travelling Pulse

1. A pulse is passed through a rope and the following pictures were obtained for each time interval:



a. Complete the following table for a particle in the medium:

time (s)	0,00	0,25	0,50	0,75	1,00	1,25	1,50	1,75	2,00
position (mm)									
velocity (mm.s^{-1})									

Table 16.5

b. Draw a position vs. time graph for the motion of the particle at $3\,\mathrm{cm}$.
c. Draw a velocity vs. time graph for the motion of the particle at $3\,\mathrm{cm}$.
d. Draw a position vs. time graph for the motion of the pulse through the rope.
e. Draw a velocity vs. time graph for the motion of the pulse through the rope.

Find the answers with the shortcodes:

(1.) l1s

1.124 Transmission and Reflection of a Pulse at a Boundary

(section shortcode: P10033)

What happens when a pulse travelling in one medium finds that medium is joined to another?

1.124.1 Investigation : Two ropes

Find two different ropes and tie both ropes together. Hold the joined ropes horizontally and create a pulse by flicking the rope up and down. What happens to the pulse when it encounters the join?

When a pulse is transmitted from one medium to another, like from a thin rope to a thicker one, the nature of the pulse will change where it meets the boundary of the two media (i.e. where the two ropes are joined). Part of the pulse will be reflected and part of it will be transmitted. Figure 16.27 shows the general case of a pulse meeting a boundary. The **incident pulse** is the one that arrives at the boundary. The **reflected pulse** is the one that moves back, away from the boundary. The **transmitted pulse** is the one that moves into the new medium, away from the boundary. The speed of the pulse depends on the mass of the rope; the pulse is faster in the thinner rope and slower in the thick rope. When the speed of the pulse increases, the pulse length will increase. If the speed decreases, the pulse length will decrease.

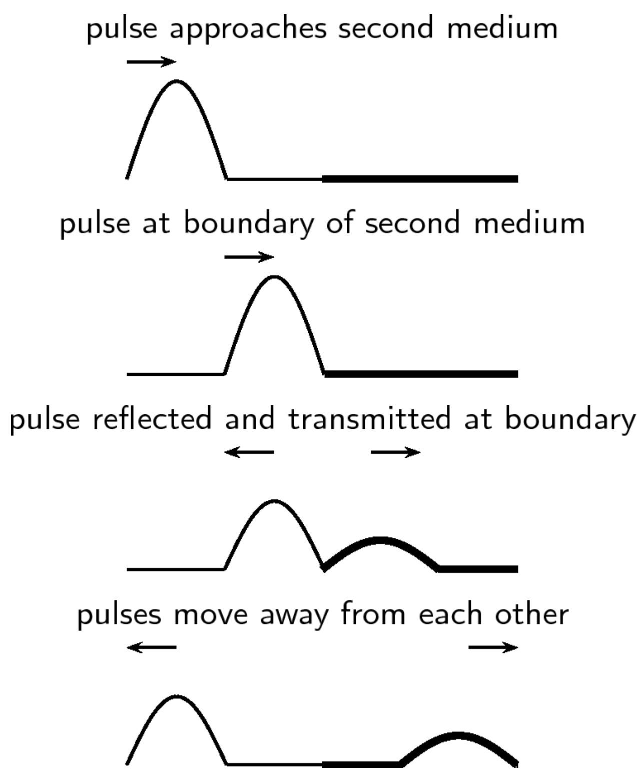


Figure 16.27: Reflection and transmission of a pulse at the boundary between two media.

Consider a pulse moving from a thin rope to a thick rope. As the pulse crosses the boundary, the speed of the pulse will decrease as it moves into the thicker rope. The pulse will move slower, so the pulse length will decrease. The pulse will be reflected and inverted in the thin rope. The reflected pulse will have the same length and speed but will have a smaller amplitude. This is illustrated in Figure 16.28.

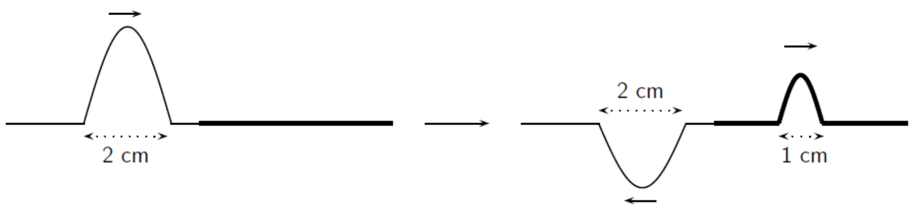


Figure 16.28: Reflection and transmission of a pulse at the boundary between two media.

When a pulse moves from a thick rope to a thin rope, the opposite will happen. As the pulse crosses the boundary, the speed of the pulse will increase as it moves into the thinner rope. The pulse in the thin rope will move faster, so the pulse length will increase. The pulse will be reflected but not inverted in the thick rope. The reflected pulse will have the same length and speed but will have a smaller amplitude. This is illustrated in Figure 16.29

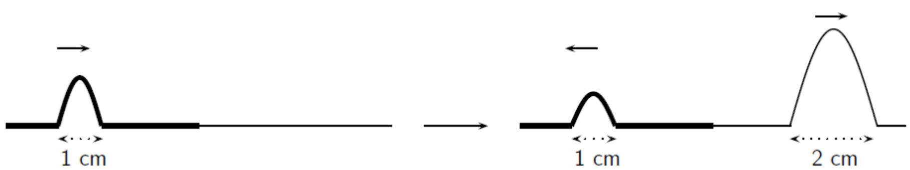


Figure 16.29: Reflection and transmission of a pulse at the boundary between two media.

1.124.2 Pulses at a Boundary I

1. Fill in the blanks or select the correct answer: A pulse in a heavy rope is traveling towards the boundary with a thin piece of string.
   - **a.** The reflected pulse in the heavy rope ***will/will not*** be inverted because __________.
   - **b.** The speed of the transmitted pulse will be ***greater than/less than/the same as*** the speed of the incident pulse.
   - **c.** The speed of the reflected pulse will be ***greater than/less than/the same as*** the speed of the incident pulse.
   - **d.** The pulse length of the transmitted pulse will be ***greater than/less than/the same as*** the pulse length of the incident pulse.
   - **e.** The frequency of the transmitted pulse will be ***greater than/less than/the same as*** the frequency of the incident pulse.
2. A pulse in a light string is traveling towards the boundary with a heavy rope.
   - **a.** The reflected pulse in the light rope ***will/will not*** be inverted because __________.
   - **b.** The speed of the transmitted pulse will be ***greater than/less than/the same as*** the speed of the incident pulse.

c. The speed of the reflected pulse will be ***greater than/less than/the same as*** the speed of the incident pulse.

d. The pulse length of the transmitted pulse will be ***greater than/less than/the same as*** the pulse length of the incident pulse.

Find the answers with the shortcodes:

(1.) I1H (2.) I16

1.125 Reflection of a Pulse from Fixed and Free Ends (not in CAPS - included for completeness)

(section shortcode: P10034)

Let us now consider what happens to a pulse when it reaches the end of a medium. The medium can be fixed, like a rope tied to a wall, or it can be free, like a rope tied loosely to a pole.

1.125.1 Reflection of a Pulse from a Fixed End

Investigation : Reflection of a Pulse from a Fixed End

Tie a rope to a wall or some other object that cannot move. Create a pulse in the rope by flicking one end up and down. Observe what happens to the pulse when it reaches the wall.

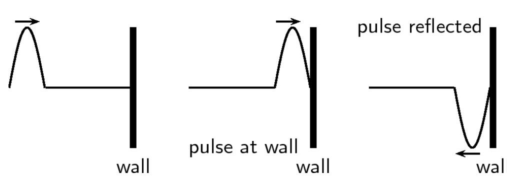


Figure 16.30: Reflection of a pulse from a fixed end.

When the end of the medium is fixed, for example a rope tied to a wall, a pulse reflects from the fixed end, but the pulse is inverted (i.e. it is upside-down). This is shown in Figure 16.30.

1.125.2 Reflection of a Pulse from a Free End

Investigation : Reflection of a Pulse from a Free End

Tie a rope to a pole in such a way that the rope can move up and down the pole. Create a pulse in the rope by flicking one end up and down. Observe what happens to the pulse when it reaches the pole.

When the end of the medium is free, for example a rope tied loosely to a pole, a pulse reflects from the free end, but the pulse is not inverted. This is shown in Figure 16.31. We draw the free end as a ring around the pole. The ring will move up and down the pole, while the pulse is reflected away from the pole.

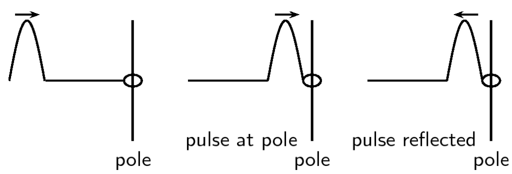


Figure 16.31: Reflection of a pulse from a free end.

TIP: The fixed and free ends that were discussed in this section are examples of *boundary conditions*. You will see more of boundary conditions as you progress in the Physics syllabus.

Pulses at a Boundary II

1. A rope is tied to a tree and a single pulse is generated. What happens to the pulse as it reaches the tree? Draw a diagram to explain what happens.
2. A rope is tied to a ring that is loosely fitted around a pole. A single pulse is sent along the rope. What will happen to the pulse as it reaches the pole? Draw a diagram to explain your answer.

The following simulation will help you understand the previous examples. Choose pulse from the options (either manual, oscillate or pulse). Then click on pulse and see what happens. Change from a fixed to a free end and see what happens. Try varying the width, amplitude, damping and tension.

Phet simulation for transverse pulses www (Simulation: P10035)

www Find the answers with the shortcodes:

(1.) I1F (2.) I1L

1.126 Summary

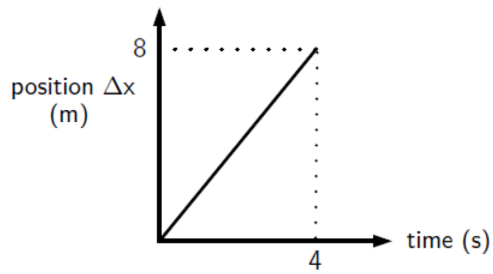
(section shortcode: P10036)

- A medium is the substance or material in which a wave will move
- A pulse is a single disturbance that moves through a medium
- The amplitude of a pules is a measurement of how far the medium is displaced from rest
- Pulse speed is the distance a pulse travels per unit time
- Constructive interference is when two pulses meet and result in a bigger pulse
- Destructive interference is when two pulses meet and and result in a smaller pulse
- We can draw graphs to show the motion of a particle in the medium or to show the motion of a pulse through the medium
- When a pulse moves from a thin rope to a thick rope, the speed and pulse length decrease. The pulse will be reflected and inverted in the thin rope. The reflected pulse has the same length and speed, but a different amplitude
- When a pulse moves from a thick rope to a thin rope, the speed and pulse length increase. The pulse will be reflected in the thick rope. The reflected pulse has the same length and speed, but a different amplitude
- A pulse reaching a free end will be reflected but not inverted. A pulse reaching a fixed end will be reflected and inverted

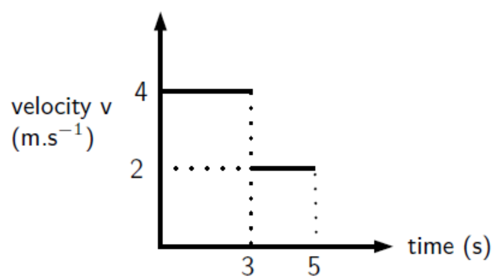
1.127 Exercises - Transverse Pulses

(section shortcode: P10037)

1. A heavy rope is flicked upwards, creating a single pulse in the rope. Make a drawing of the rope and indicate the following in your drawing:
   a. The direction of motion of the pulse
   b. Amplitude
   c. Pulse length
   d. Position of rest
2. A pulse has a speed of $2,5\,\mathrm{m\cdot s^{-1}}$. How far will it have travelled in $6\,\mathrm{s}$?
3. A pulse covers a distance of $75\,\mathrm{cm}$ in $2,5\,\mathrm{s}$. What is the speed of the pulse?
4. How long does it take a pulse to cover a distance of $200\,\mathrm{mm}$ if its speed is $4\,\mathrm{m\cdot s^{-1}}$?
5. The following position-time graph for a pulse in a slinky spring is given. Draw an accurate sketch graph of the velocity of the pulse against time.



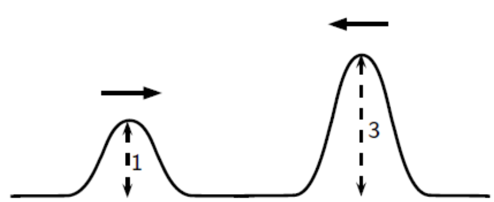
6. The following velocity-time graph for a particle in a medium is given. Draw an accurate sketch graph of the position of the particle vs. time.



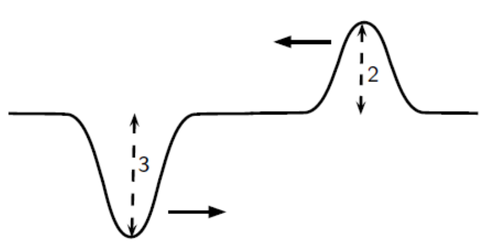
7. Describe what happens to a pulse in a slinky spring when:
   a. the slinky spring is tied to a wall.
   b. the slinky spring is loose, i.e. not tied to a wall.
   (Draw diagrams to explain your answers.)

8. The following diagrams each show two approaching pulses. Redraw the diagrams to show what type of interference takes place, and label the type of interference.

a.



b.



9. Two pulses, A and B, of identical shape and amplitude are simultaneously generated in two identical wires of equal mass and length. Wire A is, however, pulled tighter than wire B. Which pulse will arrive at the other end first, or will they both arrive at the same time?

Find the answers with the shortcodes:

(1.) lrl (2.) lri (3.) lr3 (4.) lrO (5.) lrc (6.) lrx
(7.) lra (8.) lrC (9.) lr1

Transverse Waves

1.128 Introduction

(section shortcode: P10038)

Waves occur frequently in nature. The most obvious examples are waves in water, on a dam, in the ocean, or in a bucket. We are most interested in the properties that waves have. All waves have the same properties, so if we study waves in water, then we can transfer our knowledge to predict how other examples of waves will behave.

1.129 What is a *transverse wave*?

(section shortcode: P10039)

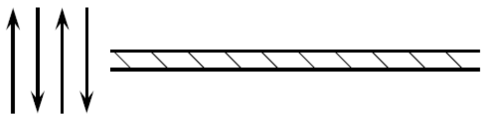
We have studied pulses in Transverse Pulses (Chapter 16), and know that a pulse is a single disturbance that travels through a medium. A *wave* is a periodic, continuous disturbance that consists of a *train* or *succession* of pulses.

Definition: Wave
A *wave* is a periodic, continuous disturbance that consists of a *train* of pulses.

Definition: Transverse wave
A *transverse wave* is a wave where the movement of the particles of the medium is perpendicular (at a right angle) to the direction of propagation of the wave.

1.129.1 Investigation : Transverse Waves

Take a rope or slinky spring. Have two people hold the rope or spring stretched out horizontally. Flick the one end of the rope up and down **continuously** to create a *train of pulses*.



1. Describe what happens to the rope.
2. Draw a diagram of what the rope looks like while the pulses travel along it.
3. In which direction do the pulses travel?
4. Tie a ribbon to the middle of the rope. This indicates a particle in the rope.

5. Flick the rope continuously. Watch the ribbon carefully as the pulses travel through the rope. What happens to the ribbon?
6. Draw a picture to show the motion of the ribbon. Draw the ribbon as a dot and use arrows to indicate how it moves.

In the Activity, you have created waves. The medium through which these waves propagated was the rope, which is obviously made up of a very large number of particles (atoms). From the activity, you would have noticed that the wave travelled from left to right, but the particles (the ribbon) moved only up and down.

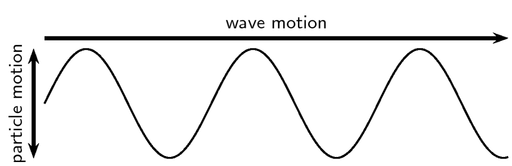


Figure 17.3: A transverse wave, showing the direction of motion of the wave perpendicular to the direction in which the particles move.

When the particles of a medium move at right angles to the direction of propagation of a wave, the wave is called *transverse*. For waves, there is no net displacement of the particles (they return to their equilibrium position), but there is a net displacement of the wave. There are thus two different motions: the motion of the particles of the medium and the motion of the wave.

The following simulation will help you understand more about waves. Select the oscillate option and then observe what happens.

Phet simulation for Transverse Waves (Simulation: P10040)

1.129.2 Peaks and Troughs

Waves have moving *peaks* (or *crests*) and *troughs*. A peak is the highest point the medium rises to and a trough is the lowest point the medium sinks to.

Peaks and troughs on a transverse wave are shown in Figure 17.5.

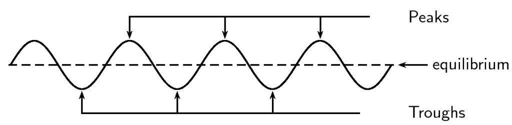


Figure 17.5: Peaks and troughs in a transverse wave.

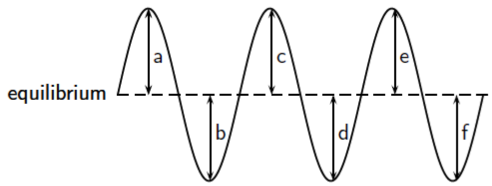
Definition: Peaks and troughs
A *peak* is a point on the wave where the displacement of the medium is at a maximum. A point on the wave is a *trough* if the displacement of the medium at that point is at a minimum.

1.129.3 Amplitude and Wavelength

There are a few properties that we saw with pulses that also apply to waves. These are amplitude and wavelength (we called this pulse length).

Definition: Amplitude
The *amplitude* is the maximum displacement of a particle from its equilibrium position.

Investigation : Amplitude



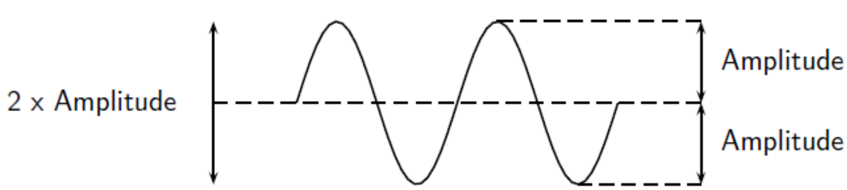
Fill in the table below by measuring the distance between the equilibrium and each peak and troughs in the wave above. Use your ruler to measure the distances.

Peak/Trough	Measurement (cm)
a	
b	
c	
d	
e	
f	

Table 17.1

1. What can you say about your results?
2. Are the distances between the equilibrium position and each peak equal?
3. Are the distances between the equilibrium position and each trough equal?
4. Is the distance between the equilibrium position and peak equal to the distance between equilibrium and trough?

As we have seen in the activity on amplitude, the distance between the peak and the equilibrium position is equal to the distance between the trough and the equilibrium position. This distance is known as the *amplitude* of the wave, and is the characteristic height of wave, above or below the equilibrium position. Normally the symbol A is used to represent the amplitude of a wave. The SI unit of amplitude is the metre (m).

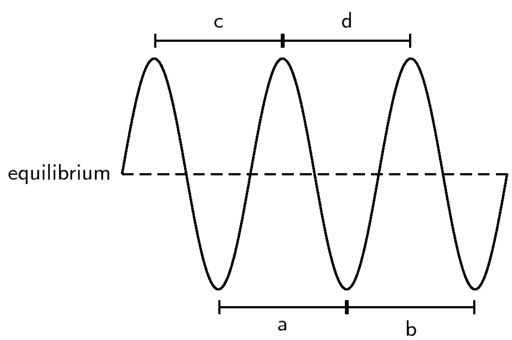


Exercise 17.1: Amplitude of Sea Waves If the peak of a wave measures $2\ \mathrm{m}$ above the still water mark in the harbour, what is the amplitude of the wave?

Solution to Exercise

Step 1. The definition of the amplitude is the height of a peak above the equilibrium position. The still water mark is the height of the water at equilibrium and the peak is $2\ \mathrm{m}$ above this, so the amplitude is $2\ \mathrm{m}$.

Investigation : Wavelength



Fill in the table below by measuring the distance between peaks and troughs in the wave above.

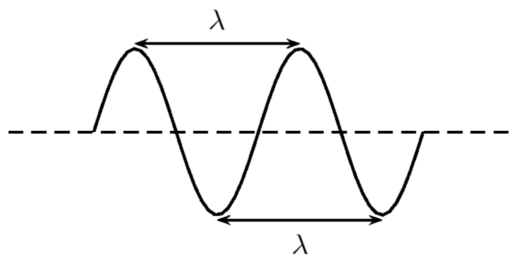
	Distance(cm)
a	
b	
c	
d	

Table 17.2

1. What can you say about your results?
2. Are the distances between peaks equal?
3. Are the distances between troughs equal?
4. Is the distance between peaks equal to the distance between troughs?

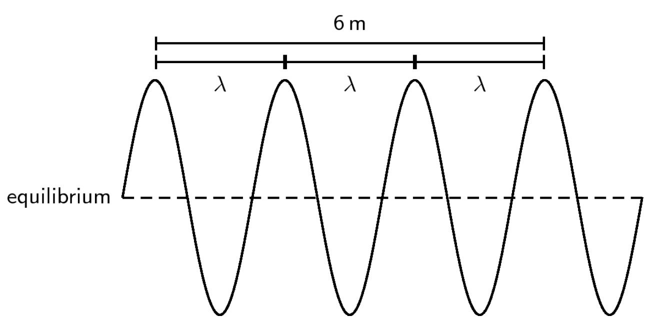
As we have seen in the activity on wavelength, the distance between two *adjacent* peaks is the same no matter which two adjacent peaks you choose. There is a fixed distance between the peaks. Similarly, we have seen that there is a fixed distance between the troughs, no matter which two troughs you look at. More importantly, the distance between two adjacent peaks is the same as the distance between two adjacent troughs. This distance is called the *wavelength* of the wave.

The symbol for the wavelength is λ (the Greek letter *lambda*) and wavelength is measured in metres (m).



Exercise 17.2: Wavelength The total distance between 4 consecutive peaks of a transverse wave is $6\ \mathrm{m}$. What is the wavelength of the wave?

Solution to Exercise



Step 1.

Step 2. From the sketch we see that 4 consecutive peaks is equivalent to 3 wavelengths.

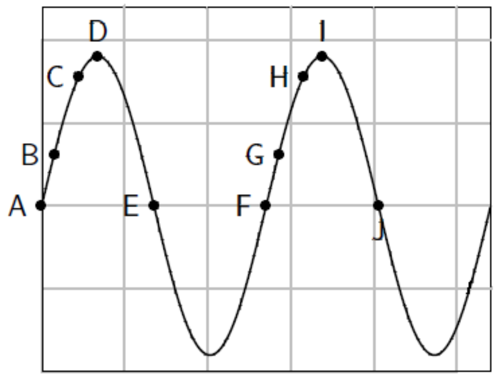
Step 3. Therefore, the wavelength of the wave is:

$$
\begin{aligned}
3\lambda &= 6\ \mathrm{m} \\
\lambda &= \frac{6\ \mathrm{m}}{3} \\
&= 2\ \mathrm{m}
\end{aligned}
\tag{17.1}
$$

1.129.4 Points in Phase

Investigation : Points in Phase

Fill in the table by measuring the distance between the indicated points.



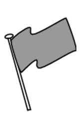
Points	Distance (cm)
A to F	
B to G	
C to H	
D to I	
E to J	

Table 17.3

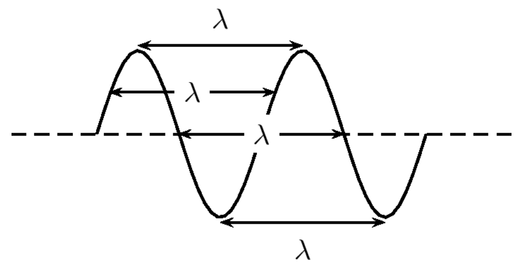
What do you find?

In the activity the distance between the indicated points was the same. These points are then said to be *in phase*. Two points in phase are separate by an integer (0,1,2,3,...) number of complete wave cycles. They do not have to be peaks or troughs, but they must be separated by a complete number of wavelengths.

We then have an alternate definition of the wavelength as the distance between any two adjacent points which are *in phase*.



Definition: Wavelength of wave
The wavelength of a wave is the distance between any two adjacent points that are in phase.



Points that are not in phase, those that are not separated by a complete number of wavelengths, are called *out of phase*. Examples of points like these would be A and C, or D and E, or B and H in the Activity.

1.129.5 Period and Frequency

Imagine you are sitting next to a pond and you watch the waves going past you. First one peak arrives, then a trough, and then another peak. Suppose you measure the time taken between one peak arriving and then the next. This time will be the same for any two successive peaks passing you. We call this time the *period*, and it is a characteristic of the wave.

The symbol T is used to represent the period. The period is measured in seconds (s).

Definition: Period (T)
The period (T) is the time taken for two successive peaks (or troughs) to pass a fixed point.

Imagine the pond again. Just as a peak passes you, you start your stopwatch and count each peak going past. After 1 second you stop the clock and stop counting. The number of peaks that you have counted in the 1 second is the *frequency* of the wave.

Definition: Frequency
The frequency is the number of successive peaks (or troughs) passing a given point in 1 second.

The frequency and the period are related to each other. As the period is the time taken for 1 peak to pass, then the number of peaks passing the point in 1 second is $\frac{1}{T}$. But this is the frequency. So

$$f = \frac{1}{T} \tag{17.2}$$

or alternatively,

$$T = \frac{1}{f} \tag{17.3}$$

For example, if the time between two consecutive peaks passing a fixed point is $\frac{1}{2}\,\mathrm{s}$, then the period of the wave is $\frac{1}{2}\,\mathrm{s}$. Therefore, the frequency of the wave is:

$$\begin{aligned} f &= \frac{1}{T} \\ &= \frac{1}{\frac{1}{2}\,\mathrm{s}} \\ &= 2\,\mathrm{s}^{-1} \end{aligned} \tag{17.4}$$

The unit of frequency is the Hertz (Hz) or s^{-1}.

Exercise 17.3: Period and Frequency What is the period of a wave of frequency $10\,\mathrm{Hz}$?

Solution to Exercise

Step 1. We are required to calculate the period of a $10\,\mathrm{Hz}$ wave.
Step 2. We know that:

$$T = \frac{1}{f} \tag{17.5}$$

Step 3.

$$\begin{aligned} T &= \frac{1}{f} \\ &= \frac{1}{10\,\mathrm{Hz}} \\ &= 0,1\,\mathrm{s} \end{aligned} \tag{17.6}$$

Step 4. The period of a $10\,\mathrm{Hz}$ wave is $0,1\,\mathrm{s}$.

1.129.6 Speed of a Transverse Wave

In Motion in One Dimension, we saw that speed was defined as

$$\text{speed} = \frac{\text{distance traveled}}{\text{time taken}} \tag{17.7}$$

The distance between two successive peaks is 1 wavelength, λ. Thus in a time of 1 period, the wave will travel 1 wavelength in distance. Thus the speed of the wave, v, is:

$$v = \frac{\text{distance traveled}}{\text{time taken}} = \frac{\lambda}{T} \tag{17.8}$$

However, $f = \frac{1}{T}$. Therefore, we can also write:

$$\begin{aligned} v &= \frac{\lambda}{T} \\ &= \lambda \cdot \frac{1}{T} \\ &= \lambda \cdot f \end{aligned} \tag{17.9}$$

We call this equation the *wave equation*. To summarise, we have that $v = \lambda \cdot f$ where

- $v =$ speed in $\mathrm{m \cdot s^{-1}}$
- $\lambda =$ wavelength in m
- $f =$ frequency in Hz

Exercise 17.4: Speed of a Transverse Wave 1 When a particular string is vibrated at a frequency of $10\,\mathrm{Hz}$, a transverse wave of wavelength $0,25\,\mathrm{m}$ is produced. Determine the speed of the wave as it travels along the string.

Solution to Exercise

Step 1.
- frequency of wave: $f = 10\mathrm{Hz}$
- wavelength of wave: $\lambda = 0,25\mathrm{m}$

We are required to calculate the speed of the wave as it travels along the string. All quantities are in SI units.

Step 2. We know that the speed of a wave is:

$$v = f \cdot \lambda \tag{17.10}$$

and we are given all the necessary quantities.

Step 3.

$$\begin{aligned} v &= f \cdot \lambda \\ &= (10\,\mathrm{Hz})\,(0,25\,\mathrm{m}) \\ &= 2,5\,\mathrm{m \cdot s^{-1}} \end{aligned} \tag{17.11}$$

Step 4. The wave travels at $2,5\,\mathrm{m \cdot s^{-1}}$ along the string.

Exercise 17.5: Speed of a Transverse Wave 2 A cork on the surface of a swimming pool bobs up and down once every second on some ripples. The ripples have a wavelength of $20\,\mathrm{cm}$. If the cork is $2\,\mathrm{m}$ from the edge of the pool, how long does it take a ripple passing the cork to reach the edge?

Solution to Exercise

Step 1. We are given:

- frequency of wave: $f = 1\,\mathrm{Hz}$
- wavelength of wave: $\lambda = 20\,\mathrm{cm}$
- distance of cork from edge of pool: $D = 2\,\mathrm{m}$

We are required to determine the time it takes for a ripple to travel between the cork and the edge of the pool.
The wavelength is not in SI units and should be converted.

Step 2. The time taken for the ripple to reach the edge of the pool is obtained from:

$$t = \frac{D}{v} \qquad \left(\text{from } v = \frac{D}{t}\right) \tag{17.12}$$

We know that

$$v = f \cdot \lambda \tag{17.13}$$

Therefore,

$$t = \frac{D}{f \cdot \lambda} \tag{17.14}$$

Step 3.

$$20\,\mathrm{cm} = 0,2\,\mathrm{m} \tag{17.15}$$

Step 4.

$$\begin{aligned} t &= \frac{D}{f \cdot \lambda} \\ &= \frac{2\ \mathrm{m}}{(1\ \mathrm{Hz})(0{,}2\ \mathrm{m})} \\ &= 10\,\mathrm{s} \end{aligned} \tag{17.16}$$

Step 5. A ripple passing the leaf will take $10\,\mathrm{s}$ to reach the edge of the pool.

The following video provides a summary of the concepts covered so far.

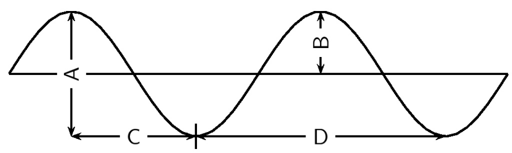
Khan academy video on waves - 1 (Video: P10041)

Waves

1. When the particles of a medium move perpendicular to the direction of the wave motion, the wave is called a .... wave.

2. A transverse wave is moving downwards. In what direction do the particles in the medium move?

3. Consider the diagram below and answer the questions that follow:

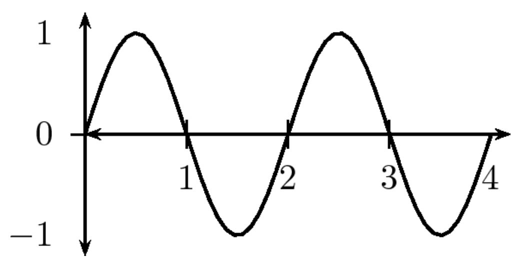


a. the wavelength of the wave is shown by letter __________.
b. the amplitude of the wave is shown by letter __________.

4. Draw 2 wavelengths of the following transverse waves on the same graph paper. Label all important values.

a. Wave 1: Amplitude = 1 cm, wavelength = 3 cm
b. Wave 2: Peak to trough distance (vertical) = 3 cm, peak to peak distance (horizontal) = 5 cm

5. You are given the transverse wave below.



Draw the following:

a. A wave with twice the amplitude of the given wave.
b. A wave with half the amplitude of the given wave.
c. A wave travelling at the same speed with twice the frequency of the given wave.
d. A wave travelling at the same speed with half the frequency of the given wave.
e. A wave with twice the wavelength of the given wave.
f. A wave with half the wavelength of the given wave.
g. A wave travelling at the same speed with twice the period of the given wave.
h. A wave travelling at the same speed with half the period of the given wave.

6. A transverse wave travelling at the same speed with an amplitude of 5 cm has a frequency of 15 Hz. The horizontal distance from a crest to the nearest trough is measured to be 2,5 cm. Find the

a. period of the wave.
b. speed of the wave.

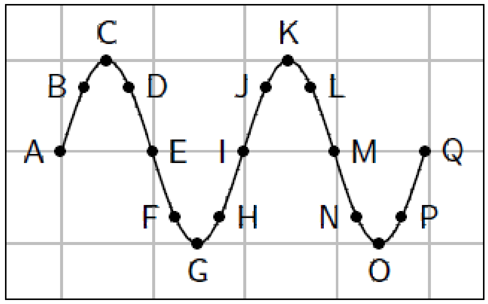
7. A fly flaps its wings back and forth 200 times each second. Calculate the period of a wing flap.

8. As the period of a wave increases, the frequency ***increases/decreases/does not change.***

9. Calculate the frequency of rotation of the second hand on a clock.

10. Microwave ovens produce radiation with a frequency of 2 450 MHz (1 MHz = 10^6 Hz) and a wavelength of 0,122 m. What is the wave speed of the radiation?

11. Study the following diagram and answer the questions:



- **a.** Identify two sets of points that are in phase.
- **b.** Identify two sets of points that are out of phase.
- **c.** Identify any two points that would indicate a wavelength.

12. Tom is fishing from a pier and notices that four wave crests pass by in 8 s and estimates the distance between two successive crests is 4 m. The timing starts with the first crest and ends with the fourth. Calculate the speed of the wave.

Find the answers with the shortcodes:

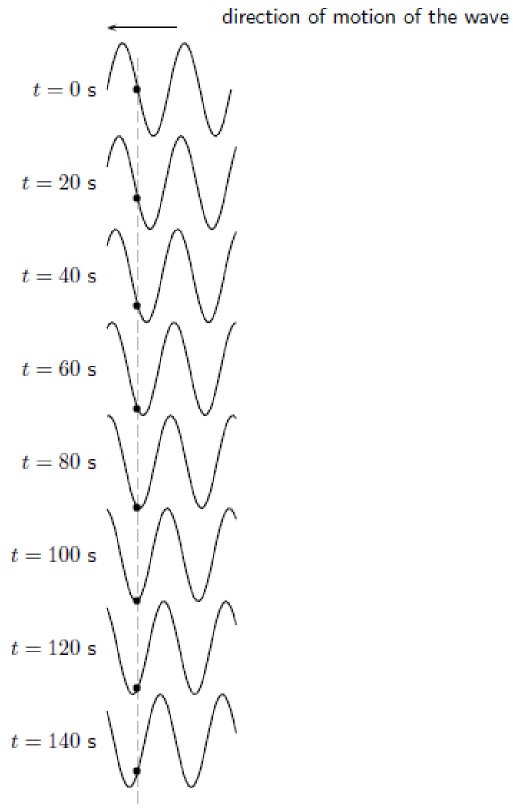
(1.) liq	(2.) li4	(3.) li2	(4.) lr8	(5.) lr9	(6.) lrX
(7.) lrl	(8.) lr5	(9.) lrN	(10.) lrR	(11.) lrn	(12.) lrQ

1.130 Graphs of Particle Motion (Not in CAPS - Included for Interest)

(section shortcode: P10042)

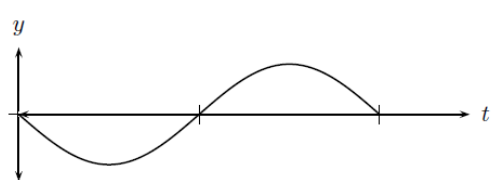
In Transverse Pulses, we saw that when a pulse moves through a medium, there are two different motions: the motion of the particles of the medium and the motion of the pulse. These two motions are at right angles to each other when the pulse is transverse. Since a transverse wave is a series of transverse pulses, the particle in the medium and the wave move in exactly the same way as for the pulse.

When a transverse wave moves horizontally through the medium, the particles in the medium **only** move up and down. We can see this in the figure below, which shows the motion of a single particle as a transverse wave moves through the medium.

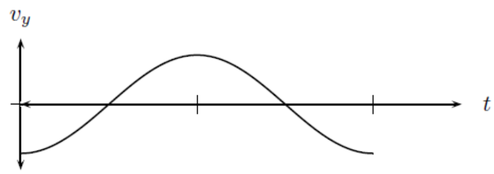


TIP: A particle in the medium **only** moves up and down when a transverse wave moves horizontally through the medium.

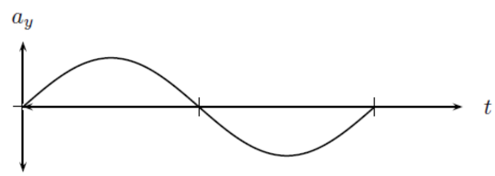
As in Transverse Pulses (Chapter 16), we can draw a graph of the particles' position as a function of time. For the wave shown in the above figure, we can draw the graph shown below.



The graph of the particle's velocity as a function of time is obtained by taking the gradient of the position vs. time graph. The graph of velocity vs. time for the position vs. time graph above, is shown in the graph below.

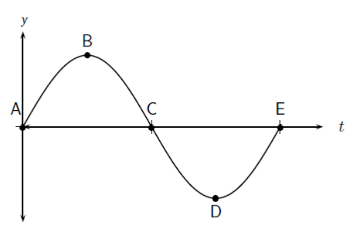


The graph of the particle's acceleration as a function of time is obtained by taking the gradient of the velocity vs. time graph. The graph of acceleration vs. time for the position vs. time graph shown above, is shown below.



As for motion in one dimension, these graphs can be used to describe the motion of the particle in the medium. This is illustrated in the worked examples below.

Exercise 17.6: Graphs of particle motion 1 The following graph shows the position of a particle of a wave as a function of time.

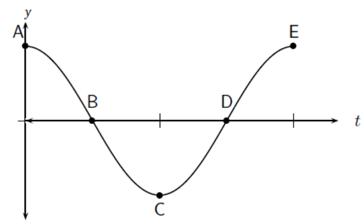


1. Draw the corresponding velocity vs. time graph for the particle.
2. Draw the corresponding acceleration vs. time graph for the particle.

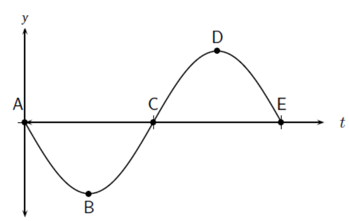
Solution to Exercise

Step 1. The y vs. t graph is given. The v_y vs. t and a_y vs. t graphs are required.

Step 2. To find the velocity of the particle we need to find the gradient of the y vs. t graph at each time. At point A the gradient is a maximum and positive. At point B the gradient is zero. At point C the gradient is a maximum, but negative. At point D the gradient is zero. At point E the gradient is a maximum and positive again.



Step 3. To find the acceleration of the particle we need to find the gradient of the v_y vs. t graph at each time. At point A the gradient is zero. At point B the gradient is negative and a maximum. At point C the gradient is zero. At point D the gradient is positive and a maximum. At point E the gradient is zero.



1.130.1 Mathematical Description of Waves

If you look carefully at the pictures of waves you will notice that they look very much like *sine* or *cosine* functions. This is correct. Waves can be described by trigonometric functions that are functions of time or of position. Depending on which case we are dealing with the function will be a function of t or x. For example, a function of position would be:

$$y\left(x\right) = Asin\left(360^\circ\frac{x}{\lambda} + \phi\right) \tag{17.17}$$

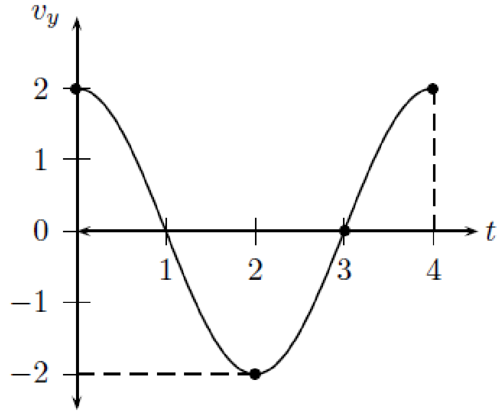
where A is the amplitude, λ the wavelength and ϕ is a *phase shift*. The phase shift accounts for the fact that the wave at $x = 0$ does not start at the equilibrium position. A function of time would be:

$$y\left(t\right) = Asin\left(360^\circ\frac{t}{T} + \phi\right) \tag{17.18}$$

where T is the period of the wave. Descriptions of the wave incorporate the amplitude, wavelength, frequency or period and a phase shift.

1.130.2 Graphs of Particle Motion

1. The following velocity vs. time graph for a particle in a wave is given.



   a. Draw the corresponding position vs. time graph for the particle.
   b. Draw the corresponding acceleration vs. time graph for the particle.

Find the answers with the shortcodes:

(1.) lrU

1.131 Standing Waves and Boundary Conditions (Not in CAPS - Included for Interest)

(section shortcode: P10043)

1.131.1 Reflection of a Transverse Wave from a Fixed End

We have seen that when a pulse meets a fixed endpoint, the pulse is reflected, but it is inverted. Since a transverse wave is a series of pulses, a transverse wave meeting a fixed endpoint is also reflected and the reflected wave is inverted. That means that the peaks and troughs are swapped around.

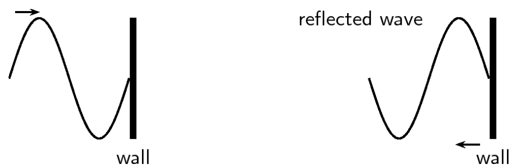


Figure 17.25: Reflection of a transverse wave from a fixed end.

1.131.2 Reflection of a Transverse Wave from a Free End

If transverse waves are reflected from an end, which is free to move, the waves sent down the string are reflected but do not suffer a phase shift as shown in Figure 17.26.

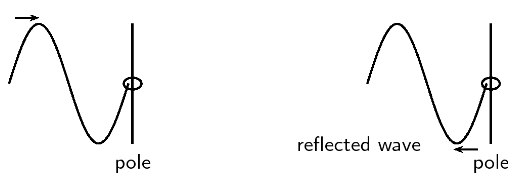


Figure 17.26: Reflection of a transverse wave from a free end.

1.131.3 Standing Waves

What happens when a reflected transverse wave meets an incident transverse wave? When two waves move in opposite directions, through each other, interference takes place. If the two waves have the same frequency and wavelength then *standing waves* are generated.

Standing waves are so-called because they appear to be standing still.

Investigation : Creating Standing Waves

Tie a rope to a fixed object such that the tied end does not move. Continuously move the free end up and down to generate firstly transverse waves and later standing waves.

We can now look closely how standing waves are formed. Figure 17.27 shows a reflected wave meeting an incident wave.

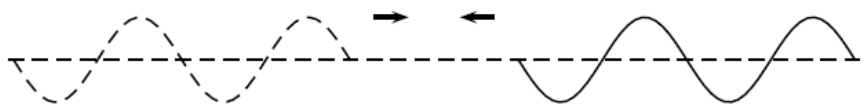


Figure 17.27: A reflected wave (solid line) approaches the incident wave (dashed line).

When they touch, both waves have an amplitude of zero:

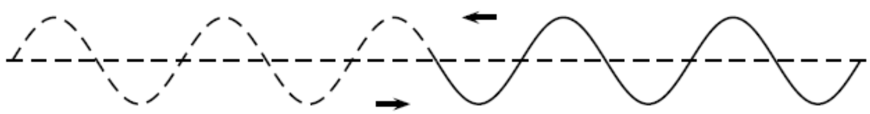


Figure 17.28: A reflected wave (solid line) meets the incident wave (dashed line).

If we wait for a short time the ends of the two waves move past each other and the waves overlap. To find the resultant wave, we add the two together.

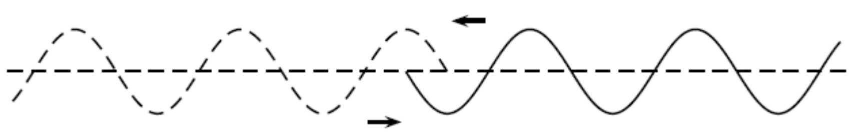
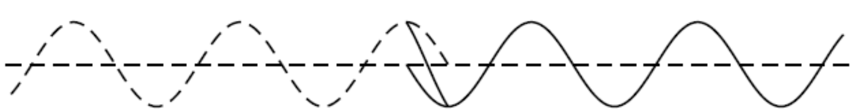


Figure 17.29: A reflected wave (solid line) overlaps slightly with the incident wave (dashed line).

In this picture, we show the two waves as dotted lines and the sum of the two in the overlap region is shown as a solid line:

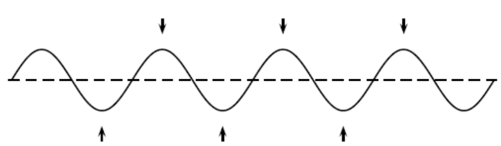


The important thing to note in this case is that there are some points where the two waves always destructively interfere to zero. If we let the two waves move a little further we get the picture below:

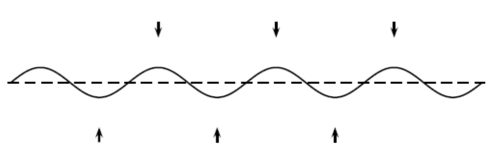
Again we have to add the two waves together in the overlap region to see what the sum of the waves looks like.

In this case the two waves have moved half a cycle past each other but because they are completely out of phase they cancel out completely.

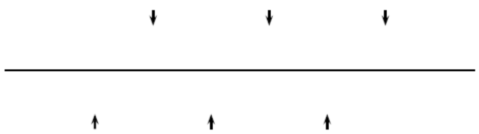
When the waves have moved past each other so that they are overlapping for a large region the situation looks like a wave oscillating in place. The following sequence of diagrams show what the resulting wave will look like. To make it clearer, the arrows at the top of the picture show peaks where maximum positive constructive interference is taking place. The arrows at the bottom of the picture show places where maximum negative interference is taking place.



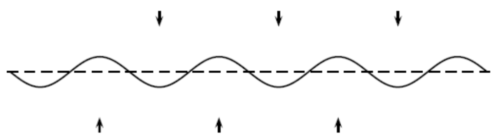
As time goes by the peaks become smaller and the troughs become shallower but they do not move.



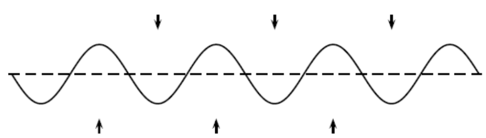
For an instant the entire region will look completely flat.



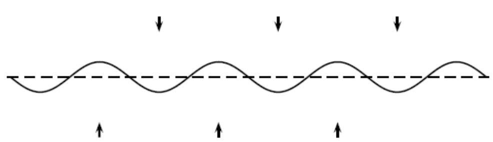
The various points continue their motion in the same manner.



Eventually the picture looks like the complete reflection through the x-axis of what we started with:



Then all the points begin to move back. Each point on the line is oscillating up and down with a different amplitude.



If we look at the overall result, we get a standing wave.

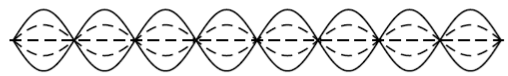


Figure 17.39: A standing wave

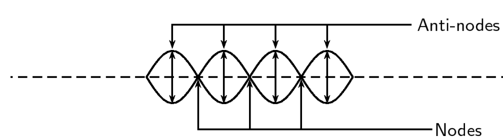
If we superimpose the two cases where the peaks were at a maximum and the case where the same waves were at a minimum we can see the lines that the points oscillate between. We call this the *envelope* of the standing wave as it contains all the oscillations of the individual points. To make the concept of the envelope clearer let us draw arrows describing the motion of points along the line.

Every point in the medium containing a standing wave oscillates up and down and the amplitude of the oscillations depends on the location of the point. It is convenient to draw the envelope for the oscillations to describe the motion. We cannot draw the up and down arrows for every single point!

NOTE: Standing waves can be a problem in for example indoor concerts where the dimensions of the concert venue coincide with particular wavelengths. Standing waves can appear as 'feedback', which would occur if the standing wave was picked up by the microphones on stage and amplified.

1.131.4 Nodes and Anti-nodes

A node is a point on a wave where no displacement takes place at any time. In a standing wave, a node is a place where two waves cancel out completely as the two waves destructively interfere in the same place. A fixed end of a rope is a node. An anti-node is a point on a wave where maximum displacement takes place. In a standing wave, an anti-node is a place where the two waves constructively interfere. Anti-nodes occur midway between nodes. A free end of a rope is an anti-node.



Definition: Node
A node is a point on a standing wave where no displacement takes place at any time. A fixed end of a rope is a node.

Definition: Anti-Node
An anti-node is a point on standing a wave where maximum displacement takes place. A free end of a rope is an anti-node.

TIP: The distance between two anti-nodes is only $\frac{1}{2}\lambda$ because it is the distance from a peak to a trough in one of the waves forming the standing wave. It is the same as the distance between two adjacent nodes. This will be important when we work out the allowed wavelengths in tubes later. We can take this further because half-way between any two anti-nodes is a node. Then the distance from the node to the anti-node is half the distance between two anti-nodes. This is half of half a wavelength which is one quarter of a wavelength, $\frac{1}{4}\lambda$.

1.131.5 Wavelengths of Standing Waves with Fixed and Free Ends

There are many applications which make use of the properties of waves and the use of fixed and free ends. Most musical instruments rely on the basic picture that we have presented to create specific sounds, either through standing pressure waves or standing vibratory waves in strings.

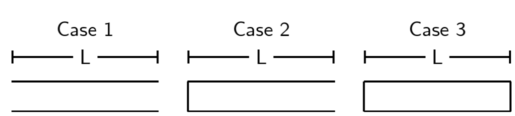
The key is to understand that a standing wave must be created in the medium that is oscillating. There are restrictions as to what wavelengths can form standing waves in a medium.

For example, if we consider a rope that can move in a pipe such that it can have

- both ends free to move (Case 1)
- one end free and one end fixed (Case 2)
- both ends fixed (Case 3).

Each of these cases is slightly different because the free or fixed end determines whether a node or anti-node will form when a standing wave is created in the rope. These are the main restrictions when we determine the wavelengths of potential standing waves. These restrictions are known as *boundary conditions* and **must** be met.

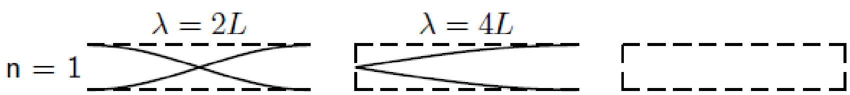
In the diagram below you can see the three different cases. It is possible to create standing waves with different frequencies and wavelengths as long as the end criteria are met.



The longer the wavelength the less the number of anti-nodes in the standing waves. We cannot have a standing wave with no anti-nodes because then there would be no oscillations. We use n to number the anti-nodes. If all of the tubes have a length L and we know the end constraints we can find the wavelength, λ, for a specific number of anti-nodes.

One Node

Let's work out the longest wavelength we can have in each tube, i.e. the case for $n = 1$.



Case 1: In the first tube, both ends must be anti-nodes, so we must place one node in the middle of the tube. We know the distance from one anti-node to another is $\frac{1}{2}\lambda$ and we also know this distance is L. So we can equate the two and solve for the wavelength:

$$\begin{aligned} \frac{1}{2}\lambda &= L \\ \lambda &= 2L \end{aligned} \tag{17.19}$$

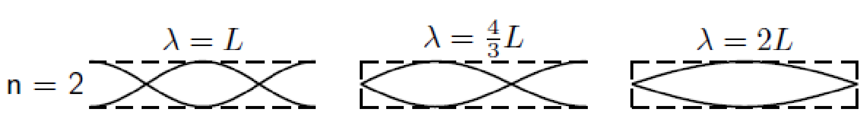
Case 2: In the second tube, one end must be a node and the other must be an anti-node. Since we are looking at the case with one node, we are forced to have it at the end. We know the distance from one node to another is $\frac{1}{2}\lambda$ but we only have half this distance contained in the tube. So :

$$\begin{aligned} \frac{1}{2}\left(\frac{1}{2}\lambda\right) &= L \\ \lambda &= 4L \end{aligned} \tag{17.20}$$

Case 3: Here both ends are closed and so we must have two nodes so it is impossible to construct a case with only one node.

Two Nodes

Next we determine which wavelengths could be formed if we had two nodes. Remember that we are dividing the tube up into smaller and smaller segments by having more nodes so we expect the wavelengths to get shorter.



Case 1: Both ends are open and so they must be anti-nodes. We can have two nodes inside the tube only if we have one anti-node contained inside the tube and one on each end. This means we have 3 anti-nodes in the tube. The distance between any two anti-nodes is half a wavelength. This means there is half wavelength between the left side and the middle and another half wavelength between the middle and the right side so there must be one wavelength inside the tube. The safest thing to do is work out how many half wavelengths there are and equate this to the length of the tube L and then solve for λ.

$$\begin{aligned} 2\left(\tfrac{1}{2}\lambda\right) &= L \\ \lambda &= L \end{aligned} \tag{17.21}$$

Case 2: We want to have two nodes inside the tube. The left end must be a node and the right end must be an anti-node. We can have one node inside the tube as drawn above. Again we can count the number of distances between adjacent nodes or anti-nodes. If we start from the left end we have one half wavelength between the end and the node inside the tube. The distance from the node inside the tube to the right end which is an anti-node is half of the distance to another node. So it is half of half a wavelength. Together these add up to the length of the tube:

$$\begin{aligned} \tfrac{1}{2}\lambda + \tfrac{1}{2}\left(\tfrac{1}{2}\lambda\right) &= L \\ \tfrac{2}{4}\lambda + \tfrac{1}{4}\lambda &= L \\ \tfrac{3}{4}\lambda &= L \\ \lambda &= \tfrac{4}{3}L \end{aligned} \tag{17.22}$$

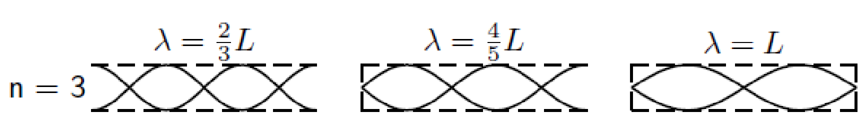
Case 3: In this case both ends have to be nodes. This means that the length of the tube is one half wavelength: So we can equate the two and solve for the wavelength:

$$\begin{aligned} \tfrac{1}{2}\lambda &= L \\ \lambda &= 2L \end{aligned} \tag{17.23}$$

TIP: If you ever calculate a longer wavelength for more nodes you have made a mistake. Remember to check if your answers make sense!

Three Nodes

To see the complete pattern for all cases we need to check what the next step for case 3 is when we have an additional node. Below is the diagram for the case where $n = 3$.



Case 1: Both ends are open and so they must be anti-nodes. We can have three nodes inside the tube only if we have two anti-nodes contained inside the tube and one on each end. This means we have 4 anti-nodes in the tube. The distance between any two anti-nodes is half a wavelength. This means there is half wavelength

between every adjacent pair of anti-nodes. We count how many gaps there are between adjacent anti-nodes to determine how many half wavelengths there are and equate this to the length of the tube L and then solve for λ.

$$\begin{aligned} 3\left(\tfrac{1}{2}\lambda\right) &= L \\ \lambda &= \tfrac{2}{3}L \end{aligned} \tag{17.24}$$

Case 2: We want to have three nodes inside the tube. The left end must be a node and the right end must be an anti-node, so there will be two nodes between the ends of the tube. Again we can count the number of distances between adjacent nodes or anti-nodes, together these add up to the length of the tube. Remember that the distance between the node and an adjacent anti-node is only half the distance between adjacent nodes. So starting from the left end we count 3 nodes, so 2 half wavelength intervals and then only a node to anti-node distance:

$$\begin{aligned} 2\left(\tfrac{1}{2}\lambda\right) + \tfrac{1}{2}\left(\tfrac{1}{2}\lambda\right) &= L \\ \lambda + \tfrac{1}{4}\lambda &= L \\ \tfrac{5}{4}\lambda &= L \\ \lambda &= \tfrac{4}{5}L \end{aligned} \tag{17.25}$$

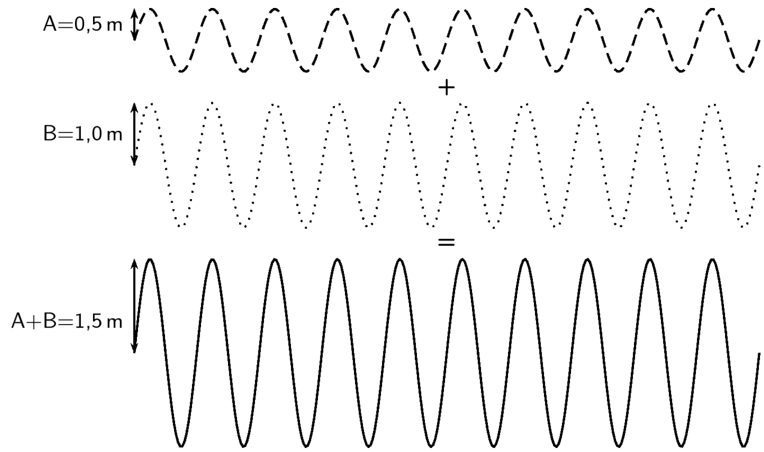
Case 3: In this case both ends have to be nodes. With one node in between there are two sets of adjacent nodes. This means that the length of the tube consists of two half wavelength sections:

$$\begin{aligned} 2\left(\tfrac{1}{2}\lambda\right) &= L \\ \lambda &= L \end{aligned} \tag{17.26}$$

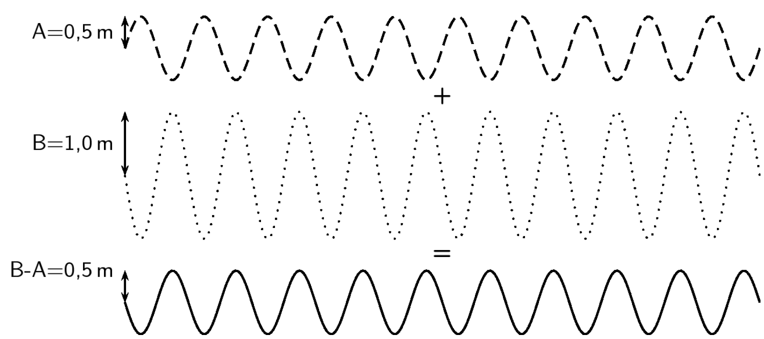
1.131.6 Superposition and Interference

If two waves meet interesting things can happen. Waves are basically collective motion of particles. So when two waves meet they both try to impose their collective motion on the particles. This can have quite different results.

If two identical (same wavelength, amplitude and frequency) waves are both trying to form a peak then they are able to achieve the sum of their efforts. The resulting motion will be a peak which has a height which is the sum of the heights of the two waves. If two waves are both trying to form a trough in the same place then a deeper trough is formed, the depth of which is the sum of the depths of the two waves. Now in this case, the two waves have been trying to do the same thing, and so add together constructively. This is called *constructive interference*.

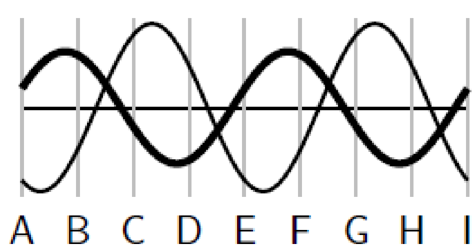


If one wave is trying to form a peak and the other is trying to form a trough, then they are competing to do different things. In this case, they can cancel out. The amplitude of the resulting wave will depend on the amplitudes of the two waves that are interfering. If the depth of the trough is the same as the height of the peak nothing will happen. If the height of the peak is bigger than the depth of the trough, a smaller peak will appear. And if the trough is deeper then a less deep trough will appear. This is *destructive interference*.



Superposition and Interference

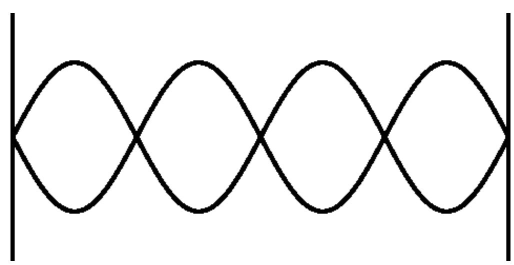
1. For each labelled point, indicate whether constructive or destructive interference takes place at that point.



Position	Constructive/Destructive
A	
B	
C	
D	
E	
F	
G	
H	
I	

2. A ride at the local amusement park is called "Standing on Standing Waves". Which position (a node or an antinode) on the ride would give the greatest thrill?

3. How many nodes and how many anti-nodes appear in the standing wave below?



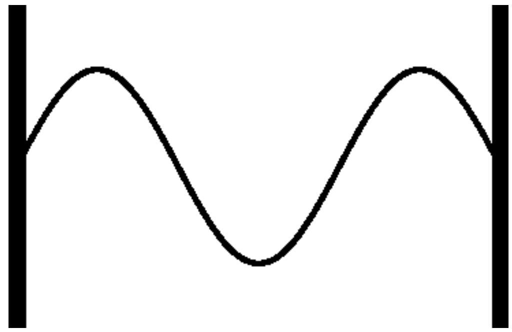
4. For a standing wave on a string, you are given three statements:

A. you can have any λ and any f as long as the relationship, $v = \lambda \cdot f$ is satisfied.
B. only certain wavelengths and frequencies are allowed
C. the wave velocity is only dependent on the medium

Which of the statements are true:

a. A and C only
b. B and C only
c. A, B, and C
d. none of the above

5. Consider the diagram below of a standing wave on a string 9 m long that is tied at both ends. The wave velocity in the string is $16\,\mathrm{m \cdot s^{-1}}$. What is the wavelength?



Find the answers with the shortcodes:

(1.) lrP (2.) lrE (3.) lrm (4.) lry (5.) lrG

1.132 Summary

(section shortcode: P10044)

1. A wave is formed when a continuous number of pulses are transmitted through a medium.
2. A peak is the highest point a particle in the medium rises to.
3. A trough is the lowest point a particle in the medium sinks to.
4. In a transverse wave, the particles move perpendicular to the motion of the wave.
5. The amplitude is the maximum distance from equilibrium position to a peak (or trough), or the maximum displacement of a particle in a wave from its position of rest.
6. The wavelength (λ) is the distance between any two adjacent points on a wave that are in phase. It is measured in metres.
7. The period (T) of a wave is the time it takes a wavelength to pass a fixed point. It is measured in seconds (s).
8. The frequency (f) of a wave is how many waves pass a point in a second. It is measured in hertz (Hz) or s^{-1}.
9. Frequency: $f = \frac{1}{T}$
10. Period: $T = \frac{1}{f}$
11. Speed: $v = f\lambda$ or $v = \frac{\lambda}{T}$.
12. When a wave is reflected from a fixed end, the resulting wave will move back through the medium, but will be inverted. When a wave is reflected from a free end, the waves are reflected, but not inverted.

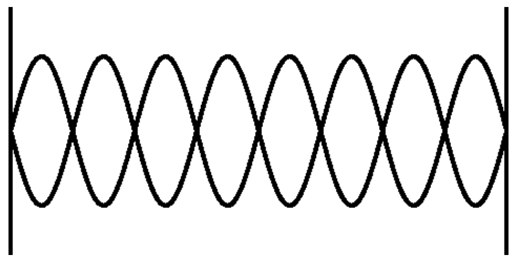
1.133 Exercises

(section shortcode: P10045)

1. A standing wave is formed when:
   a. a wave refracts due to changes in the properties of the medium
   b. a wave reflects off a canyon wall and is heard shortly after it is formed

  - **c.** a wave refracts and reflects due to changes in the medium
  - **d.** two identical waves moving different directions along the same medium interfere

2. How many nodes and anti-nodes are shown in the diagram?



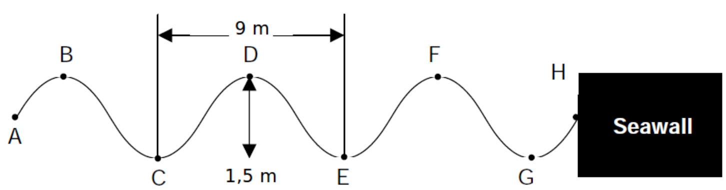
3. Draw a transverse wave that is reflected from a fixed end.

4. Draw a transverse wave that is reflected from a free end.

5. A wave travels along a string at a speed of $1,5\mathrm{m}\cdot\mathrm{s}^{-1}$. If the frequency of the source of the wave is 7,5 Hz, calculate:

  - **a.** the wavelength of the wave
  - **b.** the period of the wave

6. Water waves crash against a seawall around the harbour. Eight waves hit the seawall in 5 s. The distance between successive troughs is 9 m. The height of the waveform trough to crest is 1,5 m.



  - **a.** How many complete waves are indicated in the sketch?
  - **b.** Write down the letters that indicate any TWO points that are:
    - **i.** in phase
    - **ii.** out of phase
    - **iii.** Represent ONE wavelength.
  - **c.** Calculate the amplitude of the wave.
  - **d.** Show that the period of the wave is 0,67 s.
  - **e.** Calculate the frequency of the waves.
  - **f.** Calculate the velocity of the waves.

Find the answers with the shortcodes:

(1.) IrV (2.) Irp (3.) Ird (4.) Irv (5.) Irw (6.) Irf

Longitudinal Waves

1.134 Introduction

(section shortcode: P10046)

We have already studied transverse pulses and waves. In this chapter we look at another type of wave called *longitudinal* waves. In transverse waves, the motion of the particles in the medium were perpendicular to the direction of the wave. In longitudinal waves, the particles in the medium move *parallel* (in the *same* direction as) to the motion of the wave. Examples of transverse waves are water waves or light waves. An example of a longitudinal wave is a sound wave.

1.135 What is a *longitudinal wave*?

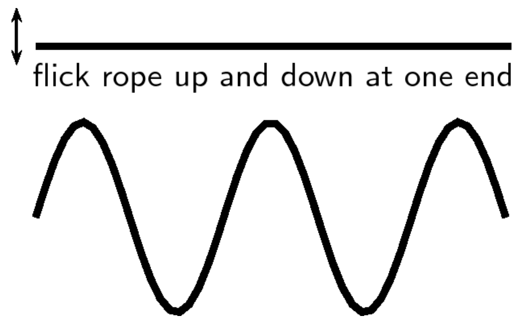
(section shortcode: P10047)

Definition: Longitudinal waves
A longitudinal wave is a wave where the particles in the medium move parallel to the direction of propagation of the wave.

When we studied transverse waves we looked at two different motions: the motion of the particles of the medium and the motion of the wave itself. We will do the same for longitudinal waves.

The question is how do we construct such a wave?

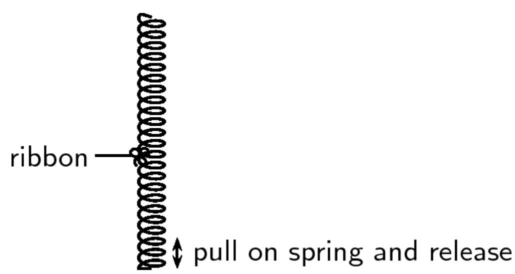
To create a transverse wave, we flick the end of for example a rope up and down. The particles move up and down and return to their equilibrium position. The wave moves from left to right and will be displaced.



A longitudinal wave is seen best in a spring that is hung from a ceiling. Do the following investigation to find out more about longitudinal waves.

1.135.1 Investigation : Investigating longitudinal waves

1. Take a spring and hang it from the ceiling. Pull the free end of the spring and release it. Observe what happens.



2. In which direction does the disturbance move?
3. What happens when the disturbance reaches the ceiling?
4. Tie a ribbon to the middle of the spring. Watch carefully what happens to the ribbon when the free end of the spring is pulled and released. Describe the motion of the ribbon.

From the investigation you will have noticed that the disturbance moves parallel to the direction in which the spring was pulled. The spring was pulled down and the wave moved up and down. The ribbon in the investigation represents one particle in the medium. The particles in the medium move in the same direction as the wave. The ribbon moves from rest upwards, then back to its original position, then down and then back to its original position.

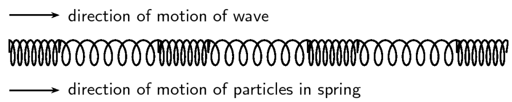


Figure 18.3: Longitudinal wave through a spring

1.136 Characteristics of Longitudinal Waves

(section shortcode: P10048)

As in the case of transverse waves the following properties can be defined for longitudinal waves: wavelength, amplitude, period, frequency and wave speed. However instead of peaks and troughs, longitudinal waves have *compressions* and *rarefactions*.

Definition: Compression
A **compression** is a region in a longitudinal wave where the particles are closest together.

Definition: Rarefaction
A **rarefaction** is a region in a longitudinal wave where the particles are furthest apart.

1.136.1 Compression and Rarefaction

As seen in Figure 18.4, there are regions where the medium is compressed and other regions where the medium is spread out in a longitudinal wave.

The region where the medium is compressed is known as a **compression** and the region where the medium is spread out is known as a **rarefaction**.

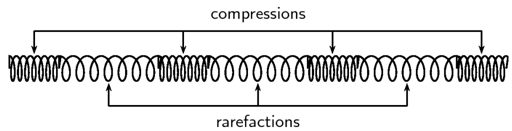
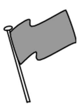


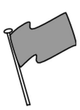
Figure 18.4: Compressions and rarefactions on a longitudinal wave

1.136.2 Wavelength and Amplitude



Definition: Wavelength
The **wavelength** in a longitudinal wave is the distance between two consecutive points that are in phase.

The wavelength in a longitudinal wave refers to the distance between two consecutive compressions or between two consecutive rarefactions.



Definition: Amplitude
The **amplitude** is the maximum displacement from equilibrium. For a longitudinal wave which is a pressure wave this would be the maximum increase (or decrease) in pressure from the equilibrium pressure that is cause when a peak (or trough) passes a point.

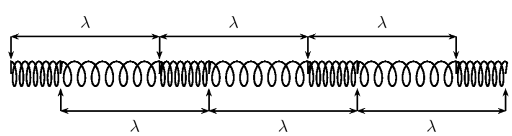


Figure 18.5: Wavelength on a longitudinal wave

The amplitude is the distance from the equilibrium position of the medium to a compression or a rarefaction.

1.136.3 Period and Frequency

Definition: Period
The **period** of a wave is the time taken by the wave to move one wavelength.

Definition: Frequency
The **frequency** of a wave is the number of wavelengths per second.

The *period* of a longitudinal wave is the time taken by the wave to move one wavelength. As for transverse waves, the symbol T is used to represent period and period is measured in seconds (s).

The *frequency* f of a wave is the number of wavelengths per second. Using this definition and the fact that the period is the time taken for 1 wavelength, we can define:

$$f = \frac{1}{T} \tag{18.1}$$

or alternately,

$$T = \frac{1}{f} \tag{18.2}$$

1.136.4 Speed of a Longitudinal Wave

The speed of a longitudinal wave is defined as:

$$v = f \cdot \lambda \tag{18.3}$$

where

- $v = \text{speed in m} \cdot \text{s}^{-1}$
- $f = \text{frequency in Hz}$
- $\lambda = \text{wavelength in m}$

Exercise 18.1: Speed of longitudinal waves The musical note “A” is a sound wave. The note has a frequency of 440 Hz and a wavelength of 0,784 m. Calculate the speed of the musical note.

Solution to Exercise

Step 1.

$$\begin{aligned} f &= 440\ \mathrm{Hz} \\ \lambda &= 0,784\ \mathrm{m} \end{aligned} \tag{18.4}$$

We need to calculate the speed of the musical note "A".

Step 2. We are given the frequency and wavelength of the note. We can therefore use:

$$v = f \cdot \lambda \tag{18.5}$$

Step 3.

$$\begin{aligned} v &= f \cdot \lambda \\ &= (440\,\mathrm{Hz})\,(0,784\,\mathrm{m}) \\ &= 345\,\mathrm{m} \cdot \mathrm{s}^{-1} \end{aligned} \tag{18.6}$$

Step 4. The musical note "A" travels at $345\,\mathrm{m} \cdot \mathrm{s}^{-1}$.

Exercise 18.2: Speed of longitudinal waves A longitudinal wave travels into a medium in which its speed increases. How does this affect its... (write only *increases, decreases, stays the same*).

1. period?
2. wavelength?

Solution to Exercise

Step 1. We need to determine how the period and wavelength of a longitudinal wave change when its speed increases.

Step 2. We need to find the link between period, wavelength and wave speed.

Step 3. We know that the frequency of a longitudinal wave is dependent on the frequency of the vibrations that lead to the creation of the longitudinal wave. Therefore, the frequency is always unchanged, irrespective of any changes in speed. Since the period is the inverse of the frequency, the period remains the same.

Step 4. The frequency remains unchanged. According to the wave equation

$$v = f\lambda \tag{18.7}$$

if f remains the same and v increases, then λ, the wavelength, must also increase.

1.137 Sound Waves

(section shortcode: P10049)

Sound waves coming from a tuning fork are caused by the vibrations of the tuning fork which push against the air particles in front of it. As the air particles are pushed together a compression is formed. The particles behind the compression move further apart causing a rarefaction. As the particles continue to push against each other, the sound wave travels through the air. Due to this motion of the particles, there is a constant variation in the pressure in the air. Sound waves are therefore pressure waves. This means that in media where the particles are closer together, sound waves will travel quicker.

Sound waves travel faster through liquids, like water, than through the air because water is denser than air (the particles are closer together). Sound waves travel faster in solids than in liquids.

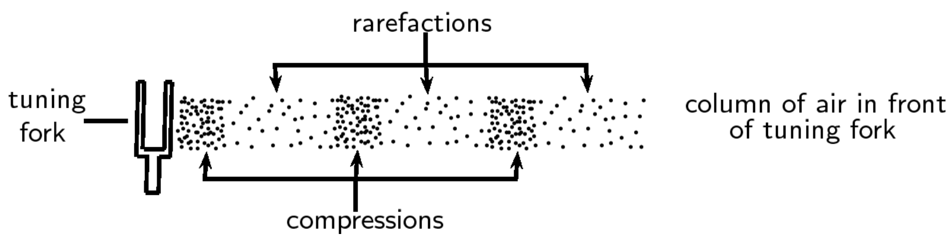


Figure 18.6: Sound waves are pressure waves and need a medium through which to travel.

TIP: A sound wave is different from a light wave.

- A sound wave is produced by an oscillating object while a light wave is not.

Also, because a sound wave is a mechanical wave (i.e. that it needs a medium) it is not capable of traveling through a vacuum, whereas a light wave can travel through a vacuum.

TIP: A sound wave is a pressure wave. This means that regions of high pressure (compressions) and low pressure (rarefactions) are created as the sound source vibrates. These compressions and rarefactions arise because the source vibrates longitudinally and the longitudinal motion of air produces pressure fluctuations.

Sound will be studied in more detail in Sound (Chapter 19).

1.138 Seismic Waves - (Not Included in CAPS - Included for Interest)

(section shortcode: P10050)

Seismic waves are waves from vibrations in the Earth (core, mantle, oceans). Seismic waves also occur on other planets, for example the moon and can be natural (due to earthquakes, volcanic eruptions or meteor strikes) or man-made (due to explosions or anything that hits the earth hard). Seismic P-waves (P for pressure) are longitudinal waves which can travel through solid and liquid.

1.139 Graphs of Particle Position, Displacement, Velocity and Acceleration (Not Included in CAPS - Included for Completeness)

(section shortcode: P10051)

When a longitudinal wave moves through the medium, the particles in the medium **only** move back and forth relative to the direction of motion of the wave. We can see this in Figure 18.7 which shows the motion of the particles in a medium as a longitudinal wave moves through the medium.

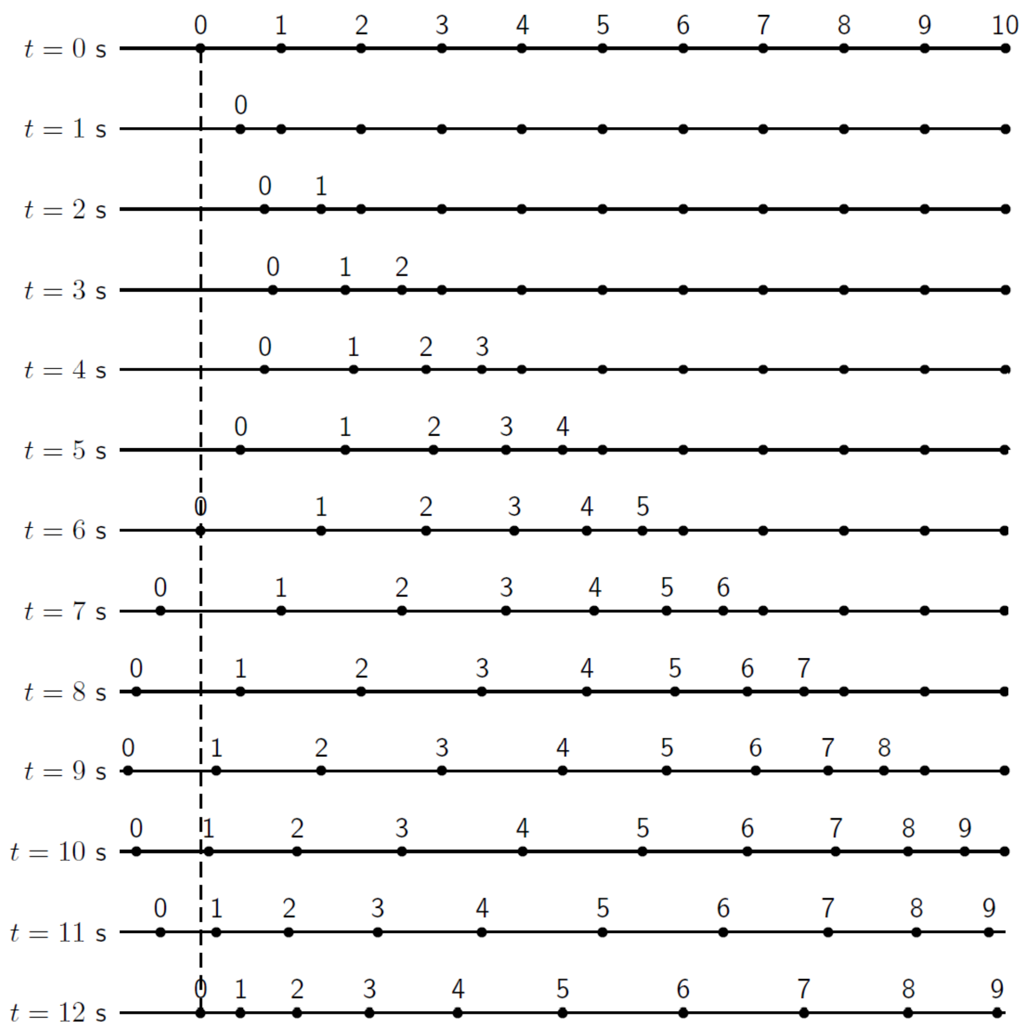


Figure 18.7: Positions of particles in a medium at different times as a longitudinal wave moves through it. The wave moves to the right. The dashed line shows the equilibrium position of particle 0.

TIP: A particle in the medium **only** moves back and forth when a longitudinal wave moves through the medium.

We can draw a graph of the particle's change in position from its starting point as a function of time. For the wave shown in Figure 18.7, we can draw the graph shown in Figure 18.8 for particle 0. The graph for each of the other particles will be identical.

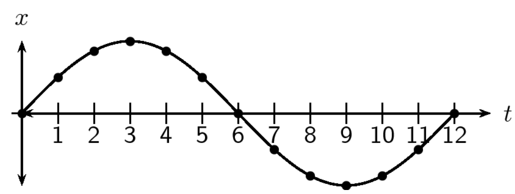


Figure 18.8: Graph of particle displacement as a function of time for the longitudinal wave shown in Figure 18.7.

The graph of the particle's velocity as a function of time is obtained by taking the gradient of the position vs. time graph. The graph of velocity vs. time for the position vs. time graph shown in Figure 18.8 is shown is Figure 18.9.

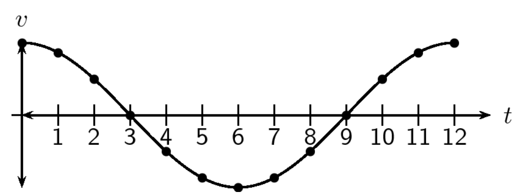


Figure 18.9: Graph of velocity as a function of time.

The graph of the particle's acceleration as a function of time is obtained by taking the gradient of the velocity vs. time graph. The graph of acceleration vs. time for the position vs. time graph shown in Figure 18.8 is shown is Figure 18.10.

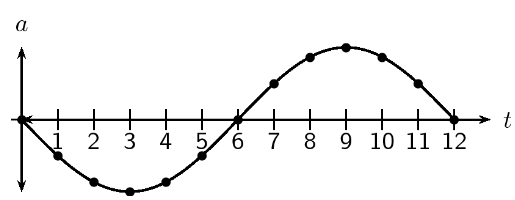


Figure 18.10: Graph of acceleration as a function of time.

1.140 Summary - Longitudinal Waves

(section shortcode: P10052)

1. A longitudinal wave is a wave where the particles in the medium move parallel to the direction in which the wave is travelling.
2. Longitudinal waves consist of areas of higher pressure, where the particles in the medium are closest together (compressions) and areas of lower pressure, where the particles in the medium are furthest apart (rarefactions).
3. The wavelength of a longitudinal wave is the distance between two consecutive compressions, or two consecutive rarefactions.
4. The relationship between the period (T) and frequency (f) is given by

$$T = \frac{1}{f} \text{ or } f = \frac{1}{T} \tag{18.8}$$

5. The relationship between wave speed (v), frequency (f) and wavelength (λ) is given by

$$v = f\lambda \tag{18.9}$$

6. Graphs of position vs time, velocity vs time and acceleration vs time can be drawn and are summarised in figures
7. Sound waves are examples of longitudinal waves. The speed of sound depends on the medium, temperature and pressure. Sound waves travel faster in solids than in liquids, and faster in liquids than in gases. Sound waves also travel faster at higher temperatures and higher pressures.

1.141 Exercises - Longitudinal Waves

(section shortcode: P10053)

1. Which of the following is not a longitudinal wave?
   - **a.** seismic P-wave
   - **b.** light
   - **c.** sound
   - **d.** ultrasound
2. Which of the following media can sound not travel through?
   - **a.** solid
   - **b.** liquid
   - **c.** gas
   - **d.** vacuum
3. Select a word from Column B that best fits the description in Column A:

Column A	Column B
waves in the air caused by vibrations	longitudinal waves
waves that move in one direction, but medium moves in another	frequency
waves and medium that move in the same direction	white noise
the distance between consecutive points of a wave which are in phase	amplitude
how often a single wavelength goes by	sound waves
half the difference between high points and low points of waves	standing waves
the distance a wave covers per time interval	transverse waves
the time taken for one wavelength to pass a point	wavelength
	music
	sounds
	wave speed

Table 18.1

4. A longitudinal wave has a crest to crest distance of 10 m. It takes the wave 5 s to pass a point.

- **a.** What is the wavelength of the longitudinal wave?
- **b.** What is the speed of the wave?

5. A flute produces a musical sound travelling at a speed of $320\,\mathrm{m \cdot s^{-1}}$. The frequency of the note is 256 Hz. Calculate:

- **a.** the period of the note
- **b.** the wavelength of the note

6. A person shouts at a cliff and hears an echo from the cliff 1 s later. If the speed of sound is $344\,\mathrm{m \cdot s^{-1}}$, how far away is the cliff?

7. A wave travels from one medium to another and the speed of the wave decreases. What will the effect be on the ... (write only *increases, decreases* or *remains the same*)

- **a.** wavelength?
- **b.** period?

Find the answers with the shortcodes:

(1.) l2d (2.) l2v (3.) l2w (4.) l2f (5.) l2G (6.) l27
(7.) l2A

Sound

1.142 Introduction

(section shortcode: P10054)

Now that we have studied the basics of longitudinal waves, we are ready to study sound waves in detail.

Have you ever thought about how amazing your sense of hearing is? It is actually pretty remarkable. There are many types of sounds: a car horn, a laughing baby, a barking dog, and somehow your brain can sort it all out. Though it seems complicated, it is rather simple to understand once you learn a very simple fact. Sound is a wave. So you can use everything you know about waves to explain sound.

1.143 Characteristics of a Sound Wave

(section shortcode: P10055)

Since sound is a wave, we can relate the properties of sound to the properties of a wave. The basic properties of sound are: pitch, loudness and tone.

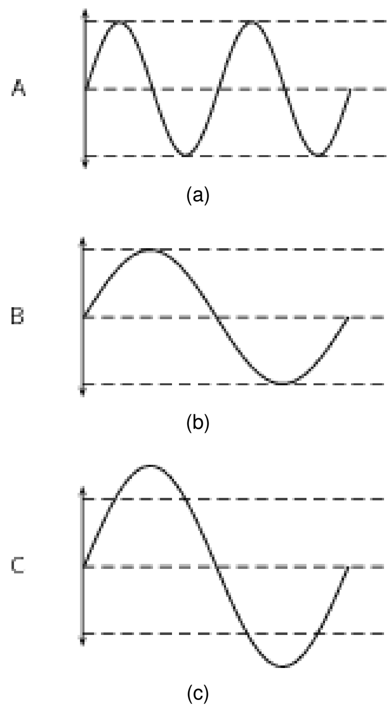


Figure 19.1: Pitch and loudness of sound. Sound B has a *lower* pitch (lower frequency) than Sound A and is *softer* (smaller amplitude) than Sound C.

1.143.1 Pitch

The frequency of a sound wave is what your ear understands as pitch. A higher frequency sound has a higher pitch, and a lower frequency sound has a lower pitch. In Figure 19.1 sound A has a higher pitch than sound B. For instance, the chirp of a bird would have a high pitch, but the roar of a lion would have a low pitch.

The human ear can detect a wide range of frequencies. Frequencies from 20 to 20 000 Hz are audible to the human ear. Any sound with a frequency below 20 Hz is known as an **infrasound** and any sound with a frequency above 20 000 Hz is known as an **ultrasound**.

Table 19.1 lists the hearing ranges of some common animals compared to humans.

	lower frequency (Hz)	upper frequency (Hz)
Humans	20	20 000
Dogs	50	45 000
Cats	45	85 000
Bats	20	120 000
Dolphins	0,25	200 000
Elephants	5	10 000

Table 19.1: Range of frequencies

Investigation : Range of Wavelengths

Using the information given in Table 19.1, calculate the lower and upper wavelengths that each species can hear. Assume the speed of sound in air is $344\,\mathrm{m}\cdot\mathrm{s}^{-1}$.

1.143.2 Loudness

The amplitude of a sound wave determines its loudness or volume. A larger amplitude means a louder sound, and a smaller amplitude means a softer sound. In Figure 19.1 sound C is louder than sound B. The vibration of a source sets the amplitude of a wave. It transmits energy into the medium through its vibration. More energetic vibration corresponds to larger amplitude. The molecules move back and forth more vigorously.

The loudness of a sound is also determined by the sensitivity of the ear. The human ear is more sensitive to some frequencies than to others. The volume we receive thus depends on both the amplitude of a sound wave and whether its frequency lies in a region where the ear is more or less sensitive.

1.143.3 Tone

Tone is a measure of the quality of the sound wave. For example, the quality of the sound produced in a particular musical instruments depends on which harmonics are superposed and in which proportions. The harmonics are determined by the standing waves that are produced in the instrument. For general interest see Physics of music, which explains the physics of music in greater detail.

The quality (timbre) of the sound heard depends on the pattern of the incoming vibrations, i.e. the *shape* of the sound wave. The more irregular the vibrations, the more jagged is the shape of the sound wave and the harsher is the sound heard.

1.144 Speed of Sound

www (section shortcode: P10056)

The speed of sound depends on the medium the sound is travelling in. Sound travels faster in solids than in liquids, and faster in liquids than in gases. This is because the density of solids is higher than that of liquids which means that the particles are closer together. Sound can be transmitted more easily.

The speed of sound also depends on the temperature of the medium. The hotter the medium is, the faster its particles move and therefore the quicker the sound will travel through the medium. When we heat a substance, the particles in that substance have more kinetic energy and vibrate or move faster. Sound can therefore be transmitted more easily and quickly in hotter substances.

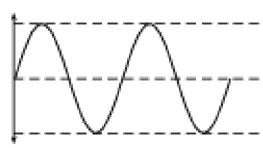
Sound waves are pressure waves. The speed of sound will therefore be influenced by the pressure of the medium through which it is travelling. At sea level the air pressure is higher than high up on a mountain. Sound will travel faster at sea level where the air pressure is higher than it would at places high above sea level.

Definition: Speed of sound
The speed of sound in air, at sea level, at a temperature of $21°\mathrm{C}$ and under normal atmospheric conditions, is $344\,\mathrm{m}\cdot\mathrm{s}^{-1}$.

1.144.1 Sound frequency and amplitude

Study the following diagram representing a musical note. Redraw the diagram for a note

1. with a higher pitch
2. that is louder
3. that is softer



Find the answers with the shortcodes:

(1.) l2o

1.145 Physics of the Ear and Hearing

(section shortcode: P10057)

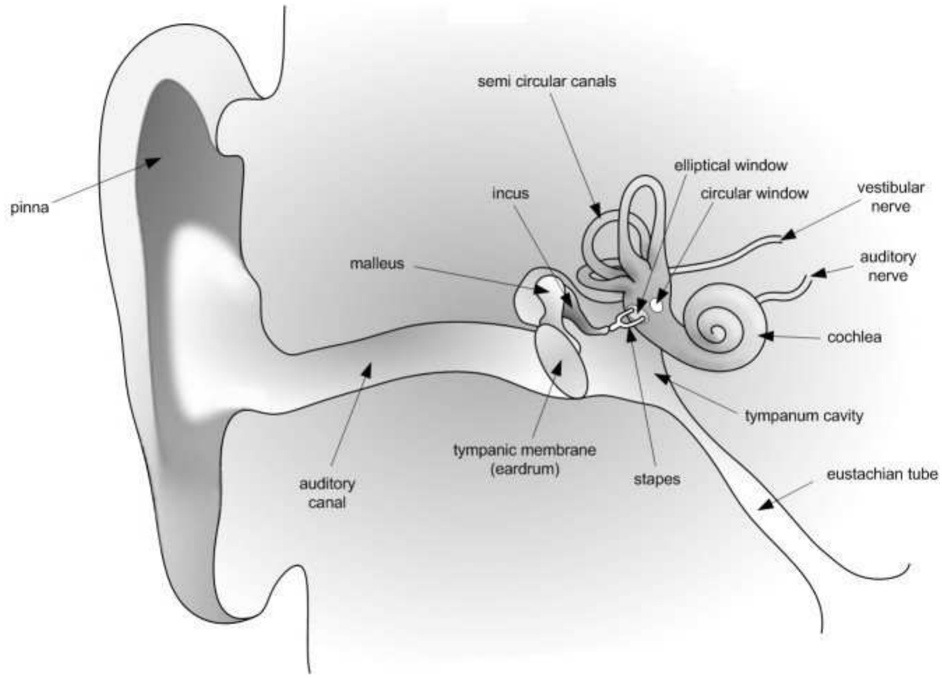


Figure 19.3: Diagram of the human ear.

The human ear is divided into three main sections: the outer, middle, and inner ear. Let's follow the journey of a sound wave from the pinna (outermost part) to the auditory nerve (innermost part) which transmits a signal to the brain. The pinna is the part of the ear we typically think of when we refer to the ear. Its main function is to collect and focus an incident sound wave. The wave then travels through the ear canal until it meets the eardrum. The pressure fluctuations of the sound wave make the eardrum vibrate. The three very small bones of the middle ear, the malleus (hammer), the incus (anvil), and the stapes (stirrup), transmit the signal through to the elliptical window. The elliptical window is the beginning of the inner ear. From the elliptical window the sound waves are transmitted through the liquid in the inner ear and interpreted as sounds by the brain. The inner ear, made of the semicircular canals, the cochlea, and the auditory nerve, is filled with fluid. The fluid allows the body to detect quick movements and maintain balance. The snail-shaped cochlea is covered in nerve cells. There are more than 25 000 hairlike nerve cells. Different nerve cells vibrate with different frequencies. When a nerve cell vibrates, it releases electrical impulses to the auditory nerve. The impulses are sent to the brain through the auditory nerve and understood as sound.

1.146 Ultrasound

(section shortcode: P10058)

Ultrasound is sound with a frequency that is higher than 20 kHz. Some animals, such as dogs, dolphins, and bats, have an upper limit that is greater than that of the human ear and can hear ultrasound.

Application	Lowest Frequency (kHz)	Highest Frequency (kHz)
Cleaning (e.g. Jewelery)	20	40
Material testing for flaws	50	500
Welding of plastics	15	40
Tumour ablation	250	2000

Table 19.2: Different uses of ultrasound and the frequencies applicable.

The most common use of ultrasound is to create images, and has industrial and medical applications. The use of ultrasound to create images is based on the reflection and transmission of a wave at a boundary. When an ultrasound wave travels inside an object that is made up of different materials such as the human body, each time it encounters a boundary, e.g. between bone and muscle, or muscle and fat, part of the wave is reflected and part of it is transmitted. The reflected rays are detected and used to construct an image of the object.

Ultrasound in medicine can visualise muscle and soft tissue, making them useful for scanning the organs, and is commonly used during pregnancy. Ultrasound is a safe, non-invasive method of looking inside the human body.

Ultrasound sources may be used to generate local heating in biological tissue, with applications in physical therapy and cancer treatment. Focussed ultrasound sources may be used to break up kidney stones.

Ultrasonic cleaners, sometimes called supersonic cleaners, are used at frequencies from 20-40 kHz for jewellery, lenses and other optical parts, watches, dental instruments, surgical instruments and industrial parts. These cleaners consist of containers with a fluid in which the object to be cleaned is placed. Ultrasonic waves are then sent into the fluid. The main mechanism for cleaning action in an ultrasonic cleaner is actually the energy released from the collapse of millions of microscopic bubbles occurring in the liquid of the cleaner.

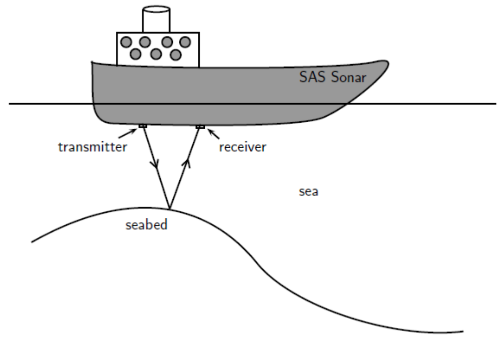
Interesting Fact

NOTE: Ultrasound generator/speaker systems are sold with claims that they frighten away rodents and insects, but there is no scientific evidence that the devices work; controlled tests have shown that rodents quickly learn that the speakers are harmless.

In echo-sounding the reflections from ultrasound pulses that are bounced off objects (for example the bottom of the sea, fish etc.) are picked up. The reflections are timed and since their speed is known, the distance to the object can be found. This information can be built into a picture of the object that reflects the ultrasound pulses.

1.147 SONAR

(section shortcode: P10059)



Ships on the ocean make use of the reflecting properties of sound waves to determine the depth of the ocean. A sound wave is transmitted and bounces off the seabed. Because the speed of sound is known and the time lapse between sending and receiving the sound can be measured, the distance from the ship to the bottom of the ocean can be determined, This is called sonar, which stands from **So**und **N**avigation **A**nd **R**anging.

1.147.1 Echolocation

Animals like dolphins and bats make use of sounds waves to find their way. Just like ships on the ocean, bats use sonar to navigate. Ultrasound waves that are sent out are reflected off the objects around the animal. Bats, or dolphins, then use the reflected sounds to form a “picture” of their surroundings. This is called echolocation.

Exercise 19.1: SONAR A ship sends a signal to the bottom of the ocean to determine the depth of the ocean. The speed of sound in sea water is $1450\,\mathrm{m.s^{-1}}$. If the signal is received 1,5 seconds later, how deep is the ocean at that point?

Solution to Exercise

Step 1.

$$\begin{aligned} s &= 1450\ \mathrm{m.s^{-1}} \\ t &= 1,5\ \mathrm{s\ there\ and\ back} \\ \therefore t &= 0,75\ \mathrm{s\ one\ way} \\ D &= ? \end{aligned} \tag{19.1}$$

Step 2.

$$\begin{aligned} \text{Distance} &= \text{speed} \times \text{time} \\ D &= s \times t \\ &= 1450\,\mathrm{m.s^{-1}} \times 0,75s \\ &= 1087,5\ \mathrm{m} \end{aligned} \tag{19.2}$$

1.148 Intensity of Sound (Not Included in CAPS - Advanced)

(section shortcode: P10060)

Interesting Fact

IMPORTANT: This section is more advanced than required and is best revisited for interest only when you are comfortable with concepts like power and logarithms.

Intensity is one indicator of amplitude. Intensity is the energy transmitted over a unit of area each second.

1.148.1 Intensity

Intensity is defined as:

$$\text{Intensity} = \frac{\text{energy}}{\text{time} \times \text{area}} = \frac{\text{power}}{\text{area}} \tag{19.3}$$

By the definition of intensity, we can see that the units of intensity are

$$\frac{\text{Joules}}{\text{s} \cdot \text{m}^2} = \frac{\text{Watts}}{\text{m}^2} \tag{19.4}$$

The unit of intensity is the **decibel** (symbol: dB). This reduces to an SI equivalent of $\text{W} \cdot \text{m}^{-2}$.

The average threshold of hearing is $10^{-12}\ \text{W} \cdot \text{m}^{-2}$. Below this intensity, the sound is too soft for the ear to hear. The threshold of pain is $1.0\ \text{W} \cdot \text{m}^{-2}$. Above this intensity a sound is so loud it becomes uncomfortable for the ear.

Notice that there is a factor of 10^{12} between the thresholds of hearing and pain. This is one reason we define the decibel (dB) scale.

1.148.2 dB Scale

The intensity in dB of a sound of intensity I, is given by:

$$\beta = 10\ log\frac{I}{I_o} \qquad I_o = 10^{-12}\ \text{W} \cdot \text{m}^{-2} \tag{19.5}$$

In this way we can compress the whole hearing intensity scale into a range from 0 dB to 120 dB.

Source	**Intensity** (dB)	**Times greater than hearing threshold**
Rocket Launch	180	10^{18}
Jet Plane	140	10^{14}
Threshold of Pain	120	10^{12}
Rock Band	110	10^{11}
Subway Train	90	10^{9}
Factory	80	10^{8}
City Traffic	70	10^{7}
Normal Conversation	60	10^{6}
Library	40	10^{4}
Whisper	20	10^{2}
Threshold of hearing	0	0

Table 19.3: Examples of sound intensities.

Notice that there are sounds which exceed the threshold of pain. Exposure to these sounds can cause immediate damage to hearing. In fact, exposure to sounds from 80 dB and above can damage hearing over time. Measures can be taken to avoid damage, such as wearing earplugs or ear muffs. Limiting exposure time and increasing distance between you and the source are also important steps for protecting your hearing.

1.148.3 Discussion : Importance of Safety Equipment

Working in groups of 5, discuss the importance of safety equipment such as ear protectors for workers in loud environments, e.g. those who use jack hammers or direct aeroplanes to their parking bays. Write up your conclusions in a one page report. Some prior research into the importance of safety equipment might be necessary to complete this group discussion.

1.149 Summary

www (section shortcode: P10061)

1. Sound waves are longitudinal waves
2. The **frequency** of a sound is an indication of how high or low the *pitch* of the sound is.
3. The human ear can hear frequencies from 20 to 20 000 Hz. **Infrasound** waves have frequencies lower than 20 Hz. **Ultrasound** waves have frequencies higher than 20 000 Hz.
4. The **amplitude** of a sound determines its *loudness* or volume.
5. The **tone** is a measure of the *quality* of a sound wave.
6. The speed of sound in air is around $340\,\mathrm{m\cdot s^{-1}}$. It is dependent on the temperature, height above sea level and the phase of the medium through which it is travelling.

7. Sound travels faster when the medium is hot.
8. Sound travels faster in a solid than a liquid and faster in a liquid than in a gas.
9. Sound travels faster at sea level where the air pressure is higher.
10. The intensity of a sound is the energy transmitted over a certain area. Intensity is a measure of frequency.
11. Ultrasound can be used to form pictures of things we cannot see, like unborn babies or tumors.
12. Echolocation is used by animals such as dolphins and bats to “see” their surroundings by using ultrasound.

13. Ships use sonar to determine how deep the ocean is or to locate shoals of fish.

1.150 Exercises

(section shortcode: P10062)

1. Choose a word from column B that best describes the concept in column A.

Column A	**Column B**
pitch of sound	amplitude
loudness of sound	frequency
quality of sound	speed
	waveform

Table 19.4

2. A tuning fork, a violin string and a loudspeaker are producing sounds. This is because they are all in a state of:
   a. compression
   b. rarefaction
   c. rotation
   d. tension
   e. vibration
3. What would a drummer do to make the sound of a drum give a note of lower pitch?
   a. hit the drum harder
   b. hit the drum less hard
   c. hit the drum near the edge
   d. loosen the drum skin
   e. tighten the drum skin
4. What is the approximate range of audible frequencies for a healthy human?
   a. 0.2 Hz $\rightarrow$ 200 Hz
   b. 2 Hz $\rightarrow$ 2 000 Hz
   c. 20 Hz $\rightarrow$ 20 000 Hz
   d. 200 Hz $\rightarrow$ 200 000 Hz
   e. 2 000 Hz $\rightarrow$ 2 000 000 Hz
5. X and Y are different wave motions. In air, X travels much faster than Y but has a much shorter wavelength. Which types of wave motion could X and Y be?

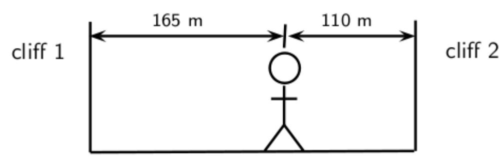
	X	Y
A	microwaves	red light
B	radio	infra red
C	red light	sound
D	sound	ultraviolet
E	ultraviolet	radio

Table 19.5

6. Astronauts are in a spaceship orbiting the moon. They see an explosion on the surface of the moon. Why can they not hear the explosion?

a. explosions do not occur in space
b. sound cannot travel through a vacuum
c. sound is reflected away from the spaceship
d. sound travels too quickly in space to affect the ear drum
e. the spaceship would be moving at a supersonic speed

7. A man stands between two cliffs as shown in the diagram and claps his hands once.



Assuming that the velocity of sound is $330\,\mathrm{m \cdot s^{-1}}$, what will be the time interval between the two loudest echoes?

a. $\frac{2}{3}\,\mathrm{s}$
b. $\frac{1}{6}\,\mathrm{s}$
c. $\frac{5}{6}\,\mathrm{s}$
d. 1 s
e. $\frac{1}{3}\,\mathrm{s}$

8. A dolphin emits an ultrasonic wave with frequency of 0,15 MHz. The speed of the ultrasonic wave in water is $1500\,\mathrm{m \cdot s^{-1}}$. What is the wavelength of this wave in water?

a. 0,1 mm
b. 1 cm
c. 10 cm
d. 10 m
e. 100 m

9. The amplitude and frequency of a sound wave are both increased. How are the loudness and pitch of the sound affected?

	loudness	pitch
A	increased	raised
B	increased	unchanged
C	increased	lowered
D	decreased	raised
E	decreased	lowered

Table 19.6

10. A jet fighter travels slower than the speed of sound. Its speed is said to be:
- **a.** Mach 1
- **b.** supersonic
- **c.** subsonic
- **d.** hypersonic
- **e.** infrasonic

11. A sound wave is different from a light wave in that a sound wave is:
- **a.** produced by a vibrating object and a light wave is not.
- **b.** not capable of traveling through a vacuum.
- **c.** not capable of diffracting and a light wave is.
- **d.** capable of existing with a variety of frequencies and a light wave has a single frequency.

12. At the same temperature, sound waves have the fastest speed in:
- **a.** rock
- **b.** milk
- **c.** oxygen
- **d.** sand

13. Two sound waves are traveling through a container of nitrogen gas. The first wave has a wavelength of 1,5 m, while the second wave has a wavelength of 4,5 m. The velocity of the second wave must be:
- **a.** $\frac{1}{9}$ the velocity of the first wave.
- **b.** $\frac{1}{3}$ the velocity of the first wave.
- **c.** the same as the velocity of the first wave.
- **d.** three times larger than the velocity of the first wave.
- **e.** nine times larger than the velocity of the first wave.

14. Sound travels at a speed of 340 m·s^{-1}. A straw is 0,25 m long. The standing wave set up in such a straw with one end closed has a wavelength of 1,0 m. The standing wave set up in such a straw with both ends open has a wavelength of 0,50 m.
- **a.** calculate the frequency of the sound created when you blow across the straw with the bottom end closed.
- **b.** calculate the frequency of the sound created when you blow across the straw with the bottom end open.

15. A lightning storm creates both lightning and thunder. You see the lightning almost immediately since light travels at $3 \times 10^8\,\mathrm{m \cdot s^{-1}}$. After seeing the lightning, you count 5 s and then you hear the thunder. Calculate the distance to the location of the storm.

16. A person is yelling from a second story window to another person standing at the garden gate, 50 m away. If the speed of sound is 344 m·s^{-1}, how long does it take the sound to reach the person standing at the gate?

17. A piece of equipment has a warning label on it that says, "Caution! This instrument produces 140 decibels." What safety precaution should you take before you turn on the instrument?

18. What property of sound is a measure of the amount of energy carried by a sound wave?

19. Person 1 speaks to person 2. Explain how the sound is created by person 1 and how it is possible for person 2 to hear the conversation.

20. Sound cannot travel in space. Discuss what other modes of communication astronauts can use when they are outside the space shuttle?

21. An automatic focus camera uses an ultrasonic sound wave to focus on objects. The camera sends out sound waves which are reflected off distant objects and return to the camera. A sensor detects the time it takes for the waves to return and then determines the distance an object is from the camera. If a sound wave (speed = 344 $m \cdot s^{-1}$) returns to the camera 0,150 s after leaving the camera, how far away is the object?

22. Calculate the frequency (in Hz) and wavelength of the annoying sound made by a mosquito when it beats its wings at the average rate of 600 wing beats per second. Assume the speed of the sound waves is 344 $m \cdot s^{-1}$.

23. How does halving the frequency of a wave source affect the speed of the waves?

24. Humans can detect frequencies as high as 20 000 Hz. Assuming the speed of sound in air is 344 $m \cdot s^{-1}$, calculate the wavelength of the sound corresponding to the upper range of audible hearing.

25. An elephant trumpets at 10 Hz. Assuming the speed of sound in air is 344 $m \cdot s^{-1}$, calculate the wavelength of this infrasonic sound wave made by the elephant.

26. A ship sends a signal out to determine the depth of the ocean. The signal returns 2,5 seconds later. If sound travels at 1450 $m.s^{-1}$ in sea water, how deep is the ocean at that point?

Find the answers with the shortcodes:

(1.) l4Y	(2.) l41	(3.) l4C	(4.) l4a	(5.) l4x	(6.) l4c
(7.) l4O	(8.) l43	(9.) l4i	(10.) l4l	(11.) l4q	(12.) lgh
(13.) lgS	(14.) lgJ	(15.) lgu	(16.) lgz	(17.) lgt	(18.) lge
(19.) lgM	(20.) lgL	(21.) lgF	(22.) lg6	(23.) lgH	(24.) lgs
(25.) lgo	(26.) lgA				

Electromagnetic Radiation

1.151 Introduction

(section shortcode: P10063)

This chapter will focus on the electromagnetic (EM) radiation. Electromagnetic radiation is a self-propagating wave in space with electric and magnetic components. These components oscillate at right angles to each other and to the direction of propagation, and are in phase with each other. Electromagnetic radiation is classified into types according to the frequency of the wave: these types include, in order of increasing frequency, radio waves, microwaves, infrared radiation, visible light, ultraviolet radiation, X-rays and gamma rays.

1.152 Particle/Wave Nature of Electromagnetic Radiation

(section shortcode: P10064)

If you watch a colony of ants walking up the wall, they look like a thin continuous black line. But as you look closer, you see that the line is made up of thousands of separated black ants.

Light and all other types of electromagnetic radiation seems like a continuous wave at first, but when one performs experiments with light, one can notice that light can have both wave and particle like properties. Just like the individual ants, the light can also be made up of individual bundles of energy, or quanta of light.

Light has both wave-like and particle-like properties (wave–particle duality), but only shows one or the other, depending on the kind of experiment we perform. A wave-type experiment shows the wave nature, and a particle-type experiment shows particle nature. One **cannot** test the wave and the particle nature at the same time. A particle of light is called a photon.

Definition: Photon
A photon is a quantum (energy packet) of light.

The particle nature of light can be demonstrated by the interaction of photons with matter. One way in which light interacts with matter is via the photoelectric effect, which will be studied in detail in Chapter 20.

1.152.1 Particle/wave nature of electromagnetic radiation

1. Give examples of the behaviour of EM radiation which can best be explained using a wave model.
2. Give examples of the behaviour of EM radiation which can best be explained using a particle model.

Find the answers with the shortcodes:

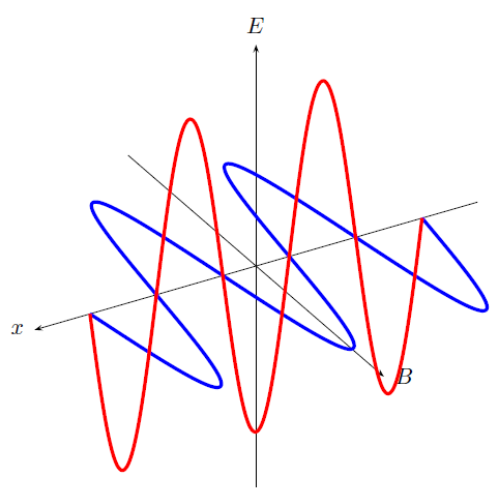
(1.) l22 (2.) l2T

1.153 The wave nature of electromagnetic radiation

(section shortcode: P10065)

Accelerating charges emit electromagnetic waves. We have seen that a changing electric field generates a magnetic field and a changing magnetic field generates an electric field. This is the principle behind the propagation of electromagnetic waves, because electromagnetic waves, unlike sound waves, do not need a medium to travel through. EM waves propagate when an electric field oscillating in one plane produces a magnetic field oscillating in a plane at right angles to it, which produces an oscillating electric field, and so on. The propagation of electromagnetic waves can be described as *mutual induction*.

These mutually regenerating fields travel through empty space at a constant speed of $3 \times 10^8\ \mathrm{m \cdot s^{-1}}$, represented by c.



1.154 Electromagnetic spectrum

(section shortcode: P10066)

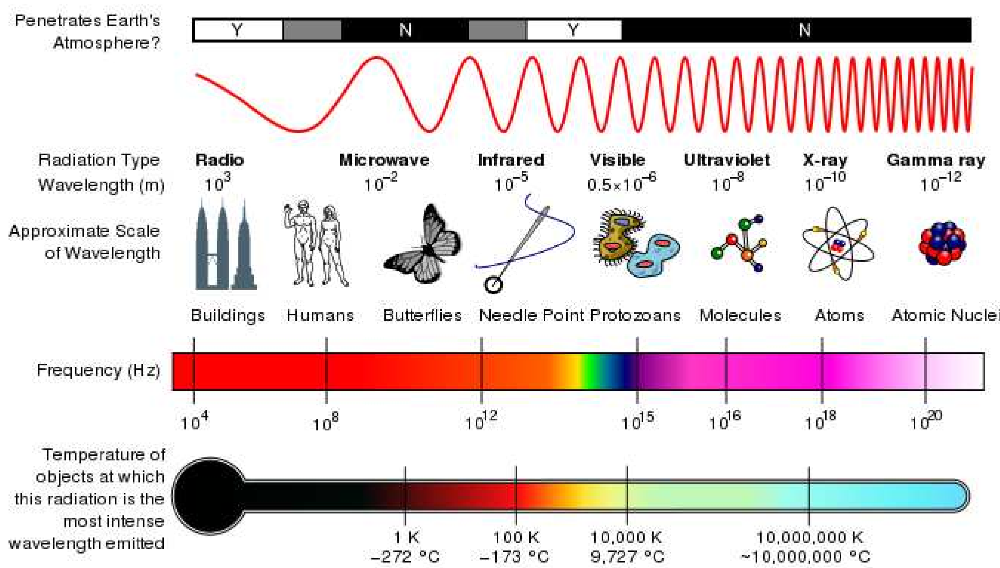


Figure 20.2: The electromagnetic spectrum as a function of frequency. The different types according to wavelength are shown as well as everyday comparisons.

Observe the things around you, your friend sitting next to you, a large tree across the field. How is it that you are able to see these things? What is it that is leaving your friend's arm and entering your eye so that you can see his arm? It is light. The light originally comes from the sun, or possibly a light bulb or burning fire. In physics, light is given the more technical term electromagnetic radiation, which includes all forms of light, not just the form which you can see with your eyes.

Electromagnetic radiation allows us to observe the world around us. It is this radiation which reflects off of the objects around you and into your eye. The radiation your eye is sensitive to is only a small fraction of the total radiation emitted in the physical universe. All of the different fractions taped together make up the electromagnetic spectrum.

1.154.1 Dispersion

When white light is split into its component colours by a prism, you are looking at a portion of the electromagnetic spectrum.

The wavelength of a particular electromagnetic radiation will depend on how it was created.

1.154.2 Wave Nature of EM Radiation

1. List one source of electromagnetic waves. Hint: consider the spectrum diagram and look at the names we give to different wavelengths.
2. Explain how an EM wave propagates, with the aid of a diagram.
3. What is the speed of light? What symbol is used to refer to the speed of light? Does the speed of light change?
4. Do EM waves need a medium to travel through?

The radiation can take on any wavelength, which means that the spectrum is continuous. Physicists broke down this continuous band into sections. Each section is defined by how the radiation is created, not the wavelength of the radiation. But each category is continuous within the min and max wavelength of that category, meaning there are no wavelengths excluded within some range.

The spectrum is in order of wavelength, with the shortest wavelength at one end and the longest wavelength at the other. The spectrum is then broken down into categories as detailed in Table 20.1.

Category	Range of Wavelengths (nm)	Range of Frequencies (Hz)
gamma rays	< 1	$> 3 \times 10^{19}$
X-rays	1-10	3×10^{17}-3×10^{19}
ultraviolet light	10-400	$7,5 \times 10^{14}$-3×10^{17}
visible light	400-700	$4,3 \times 10^{14}$-$7,5 \times 10^{14}$
infrared	700-10^5	3×10^{12}-$4,3 \times 10^{19}$
microwave	$10^5 - 10^8$	3×10^{9}-3×10^{12}
radio waves	$> 10^8$	$< 3 \times 10^{9}$

Table 20.1: Electromagnetic spectrum

Since an electromagnetic wave is still a wave, the following equation that you learnt in Grade 10 still applies:

$$c = f \cdot \lambda \qquad (20.1)$$

Exercise 20.1: EM spectrum I Calculate the frequency of red light with a wavelength of $4,2 \times 10^{-7}$ m

Solution to Exercise

Step 1. We use the formula: $c = f\lambda$ to calculate frequency. The speed of light is a constant 3×10^8m/s.

$$\begin{aligned} c &= f\lambda \\ 3 \times 10^8 &= f \times 4,2 \times 10^{-7} \\ f &= 7,14 \times 10^{14}\text{Hz} \end{aligned} \tag{20.2}$$

Exercise 20.2: EM spectrum II Ultraviolet radiation has a wavelength of $200\ \text{nm}$. What is the frequency of the radiation?

Solution to Exercise

Step 1. Recall that all radiation travels at the speed of light (c) in vacuum. Since the question does not specify through what type of material the Ultraviolet radiation is traveling, one can assume that it is traveling through a vacuum. We can identify two properties of the radiation - $wavelength\ (200\ \text{nm})$ and speed (c).

Step 2.

$$\begin{aligned} c &= f\lambda \\ 3 \times 10^8 &= f \times 200 \times 10^{-9} \\ f &= 1.5 \times 10^{15}\ \text{Hz} \end{aligned} \tag{20.3}$$

Examples of some uses of electromagnetic waves are shown in Table 20.2.

Category	**Uses**
gamma rays	used to kill the bacteria in marshmallows and to sterilise medical equipment
X-rays	used to image bone structures
ultraviolet light	bees can see into the ultraviolet because flowers stand out more clearly at this frequency
visible light	used by humans to observe the world
infrared	night vision, heat sensors, laser metal cutting
microwave	microwave ovens, radar
radio waves	radio, television broadcasts

Table 20.2: Uses of EM waves

In theory the spectrum is infinite, although realistically we can only observe wavelengths from a few hundred kilometers to those of gamma rays due to experimental limitations.

Humans experience electromagnetic waves differently depending on their wavelength. Our eyes are sensitive to visible light while our skin is sensitive to infrared, and many wavelengths we do not detect at all.

Find the answers with the shortcodes:

(1.) l2c (2.) l2x (3.) l2a (4.) l2C

1.154.3 EM Radiation

1. Arrange the following types of EM radiation in order of increasing frequency: infrared, X-rays, ultraviolet, visible, gamma.
2. Calculate the frequency of an EM wave with a wavelength of 400 nm.
3. Give an example of the use of each type of EM radiation, i.e. gamma rays, X-rays, ultraviolet light, visible light, infrared, microwave and radio and TV waves.

Find the answers with the shortcodes:

(1.) l23 (2.) l2O (3.) l2i

1.155 The particle nature of electromagnetic radiation

(section shortcode: P10067)

When we talk of electromagnetic radiation as a particle, we refer to photons, which are packets of energy. The energy of the photon is related to the wavelength of electromagnetic radiation according to:

Definition: Planck's constant
Planck's constant is a physical constant named after Max Planck.
$h = 6,626 \times 10^{-34}$ J · s

The energy of a photon can be calculated using the formula: $E = hf$ or $E = h\frac{c}{\lambda}$. Where E is the energy of the photon in joules (J), h is planck's constant, c is the speed of light, f is the frequency in hertz (Hz) and λ is the wavelength in metres (m).

Exercise 20.3: Calculating the energy of a photon I Calculate the energy of a photon with a frequency of 3×10^{18} Hz

Solution to Exercise

Step 1.

$$\begin{aligned} E &= hf \\ &= 6,6 \times 10^{-34} \times 3 \times 10^{18} \\ &= 2 \times 10^{-15}\,\mathrm{J} \end{aligned} \tag{20.4}$$

Exercise 20.4: Calculating the energy of a photon II What is the energy of an ultraviolet photon with a wavelength of 200 nm?

Solution to Exercise

Step 1. We are required to calculate the energy associated with a photon of ultraviolet light with a wavelength of 200 nm.
We can use:

$$E = h\frac{c}{\lambda} \tag{20.5}$$

Step 2.

$$\begin{aligned} E &= h\tfrac{c}{\lambda} \\ &= \left(6,626 \times 10^{-34}\right) \tfrac{3 \times 10^{8}}{200 \times 10^{-9}} \\ &= 9,939 \times 10^{-10}\,\mathrm{J} \end{aligned} \tag{20.6}$$

1.155.1 Exercise - particle nature of EM waves

1. How is the energy of a photon related to its frequency and wavelength?
2. Calculate the energy of a photon of EM radiation with a frequency of 10^{12} Hz.
3. Determine the energy of a photon of EM radiation with a wavelength of 600 nm.

Find the answers with the shortcodes:

(1.) l2H (2.) l26 (3.) l2F

1.156 Penetrating ability of electromagnetic radiation

(section shortcode: P10068)

Different kinds of electromagnetic radiation have different penetrabilities. For example, if we take the human body as the object. Infrared light is emitted by the human body. Visible light is reflected off the surface of the human body, ultra-violet light (from sunlight) damages the skin, but X-rays are able to penetrate the skin and bone and allow for pictures of the inside of the human body to be taken.

If we compare the energy of visible light to the energy of X-rays, we find that X-rays have a much higher energy. Usually, kinds of electromagnetic radiation with higher energy have higher penetrabilities than those with low energies.

Certain kinds of electromagnetic radiation such as ultra-violet radiation, X-rays and gamma rays are very dangerous. Radiation such as these are called ionising radiation. Ionising radiation transfers energy as it passes through matter, breaking molecular bonds and creating ions.

Excessive exposure to radiation, including sunlight, X-rays and all nuclear radiation, can cause destruction of biological tissue.

1.156.1 Ultraviolet(UV) radiation and the skin

UVA and UVB are different ranges of frequencies for ultraviolet (UV) light. UVA and UVB can damage collagen fibres which results in the speeding up skin aging. In general, UVA is the least harmful, but it can contribute to the aging of skin, DNA damage and possibly skin cancer. It penetrates deeply and does not cause sunburn. Because it does not cause reddening of the skin (erythema) it cannot be measured in the SPF testing. There is no good clinical measurement of the blocking of UVA radiation, but it is important that sunscreen block both UVA and UVB.

UVB light can cause skin cancer. The radiation excites DNA molecules in skin cells, resulting in possible mutations, which can cause cancer. This cancer connection is one reason for concern about ozone depletion and the ozone hole.

As a defense against UV radiation, the body tans when exposed to moderate (depending on skin type) levels of radiation by releasing the brown pigment melanin. This helps to block UV penetration and prevent damage to the vulnerable skin tissues deeper down. Suntan lotion, often referred to as sunblock or sunscreen, partly blocks UV and is widely available. Most of these products contain an SPF rating that describes the amount of protection given. This protection, however, applies only to UVB rays responsible for sunburn and not to UVA rays that penetrate more deeply into the skin and may also be responsible for causing cancer and wrinkles. Some sunscreen lotion now includes compounds such as titanium dioxide which helps protect against UVA rays. Other UVA blocking compounds found in sunscreen include zinc oxide and avobenzone.

What makes a good sunscreen?

- UVB protection: Padimate O, Homosalate, Octisalate (octyl salicylate), Octinoxate (octyl methoxycinnamate)
- UVA protection: Avobenzone
- UVA/UVB protection: Octocrylene, titanium dioxide, zinc oxide, Mexoryl (ecamsule)

Another means to block UV is by wearing sun protective clothing. This is clothing that has a UPF rating that describes the protection given against both UVA and UVB.

1.156.2 Ultraviolet radiation and the eyes

High intensities of UVB light are hazardous to the eyes, and exposure can cause welder's flash (photo keratitis or arc eye) and may lead to cataracts, pterygium and pinguecula formation.

Protective eyewear is beneficial to those who are working with or those who might be exposed to ultraviolet radiation, particularly short wave UV. Given that light may reach the eye from the sides, full coverage eye protection is usually warranted if there is an increased risk of exposure, as in high altitude mountaineering. Mountaineers are exposed to higher than ordinary levels of UV radiation, both because there is less atmospheric filtering and because of reflection from snow and ice.

Ordinary, untreated eyeglasses give some protection. Most plastic lenses give more protection than glass lenses. Some plastic lens materials, such as polycarbonate, block most UV. There are protective treatments available for eyeglass lenses that need it which will give better protection. But even a treatment that completely blocks UV will not protect the eye from light that arrives around the lens. To convince yourself of the potential dangers of stray UV light, cover your lenses with something opaque, like aluminum foil, stand next to a bright light, and consider how much light you see, despite the complete blockage of the lenses. Most contact lenses help to protect the retina by absorbing UV radiation.

1.156.3 X-rays

While x-rays are used significantly in medicine, prolonged exposure to X-rays can lead to cell damage and cancer.

For example, a mammogram is an x-ray of the human breast to detect breast cancer, but if a woman starts having regular mammograms when she is too young, her chances of getting breast cancer increases.

1.156.4 Gamma-rays

Due to the high energy of gamma-rays, they are able to cause serious damage when absorbed by living cells.

Gamma-rays are not stopped by the skin and can induce DNA alteration by interfering with the genetic material of the cell. DNA double-strand breaks are generally accepted to be the most biologically significant lesion by which ionising radiation causes cancer and hereditary disease.

A study done on Russian nuclear workers exposed to external whole-body gamma-radiation at high cumulative doses shows a link between radiation exposure and death from leukaemia, lung, liver, skeletal and other solid cancers.

Cellphones and electromagnetic radiation

Cellphone radiation and health concerns have been raised, especially following the enormous increase in the use of wireless mobile telephony throughout the world. This is because mobile phones use electromagnetic waves in the microwave range. These concerns have induced a large body of research. Concerns about effects on health have also been raised regarding other digital wireless systems, such as data communication networks. In 2009 the World Health Organisation announced that they have found a link between brain cancer and cellphones.

Cellphone users are recommended to minimise radiation, by for example:

1. Use hands-free to decrease the radiation to the head.
2. Keep the mobile phone away from the body.
3. Do not telephone in a car without an external antenna.

1.156.5 Exercise - Penetrating ability of EM radiation

1. Indicate the penetrating ability of the different kinds of EM radiation and relate it to energy of the radiation.
2. Describe the dangers of gamma rays, X-rays and the damaging effect of ultra-violet radiation on skin.

Find the answers with the shortcodes:

(1.) l2l (2.) l2q

1.157 Summary

(section shortcode: P10069)

1. Electromagnetic radiation has both a wave and a particle nature.
2. Electromagnetic waves travel at a speed of $3 \times 10^8\ m \cdot s^{-1}$ in a vaccum.
3. The Electromagnetic spectrum consists of the follwing types of radiation: radio waves, microwaves, infrared, visible, ultraviolet, X-rays, gamma-rays.
4. Gamma-rays have the most energy and are the most penetrating, while radio waves have the lowest energy and are the least penetrating.

1.158 End of chapter exercise

(section shortcode: P10070)

1. What is the energy of a photon of EM radiation with a frequency of 3×10^8 Hz?
2. What is the energy of a photon of light with a wavelength of 660 nm?
3. List the main types of electromagnetic radiation in order of increasing wavelength.
4. List the main uses of:
   a. radio waves

  b. infrared
  c. gamma rays
  d. X-rays

5. Explain why we need to protect ourselves from ultraviolet radiation from the Sun.

6. List some advantages and disadvantages of using X-rays.

7. What precautions should we take when using cell phones?

8. Write a short essay on a type of electromagnetic waves. You should look at uses, advantages and disadvantages of your chosen radiation.

9. Explain why some types of electromagnetic radiation are more penetrating than others.

www Find the answers with the shortcodes:

(1.) l4J (2.) l4u (3.) l2r (4.) l21 (5.) l2Y (6.) l4h
(7.) l4S (8.) l24 (9.) l2g

Magnetism

1.159 Introduction

(section shortcode: P10071)

Magnetism is a force that certain kinds of objects, which are called 'magnetic' objects, can exert on each other without physically touching. A magnetic object is surrounded by a magnetic 'field' that gets weaker as one moves further away from the object. A second object can feel a magnetic force from the first object because it feels the magnetic field of the first object.

Humans have known about magnetism for many thousands of years. For example, *lodestone* is a magnetised form of the iron oxide mineral *magnetite*. It has the property of attracting iron objects. It is referred to in old European and Asian historical records; from around 800 BCE in Europe and around 2 600 BCE in Asia.

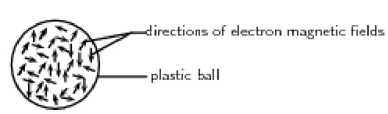
NOTE: The root of the English word *magnet* is from the Greek word *magnes*, probably from Magnesia in Asia Minor, once an important source of lodestone.

1.160 Magnetic fields

(section shortcode: P10072)

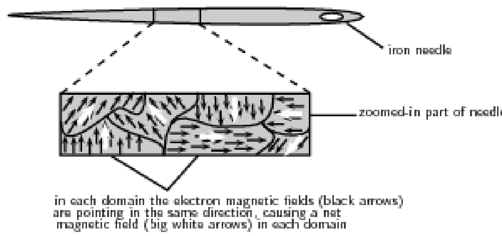
A magnetic field is a region in space where a magnet or object made of magnetic material will experience a non-contact force.

Electrons inside any object have magnetic fields associated with them. In most materials these fields point in all directions, so the net magnetic field is zero. For example, in the plastic ball below, the directions of the magnetic fields of the electrons (shown by the arrows) are pointing in different directions and cancel each other out. Therefore the plastic ball is not magnetic and has no magnetic field.

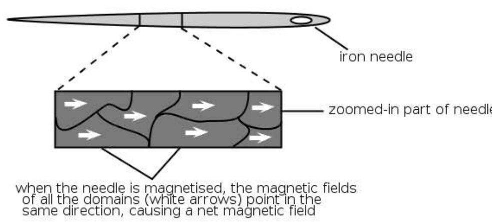


The electron magnetic fields point in all directions so there is no net (total) magnetic field for the whole ball

In some materials (e.g. iron), called **ferromagnetic** materials, there are regions called *domains*, where the electrons' magnetic fields line up with each other. All the atoms in each domain are grouped together so that the magnetic fields from their electrons point the same way. The picture shows a piece of an iron needle zoomed in to show the domains with the electric fields lined up inside them.



In permanent magnets, many domains are lined up, resulting in a *net magnetic field*. Objects made from ferromagnetic materials can be magnetised, for example by rubbing a magnet along the object in one direction. This causes the magnetic fields of most, or all, of the domains to line up in one direction. As a result the object as a whole will have a net magnetic field. It is *magnetic*. Once a ferromagnetic object has been magnetised, it can stay magnetic without another magnet being nearby (i.e. without being in another magnetic field). In the picture below, the needle has been magnetised because the magnetic fields in all the domains are pointing in the same direction.



1.160.1 Investigation : Ferromagnetic materials and magnetisation

1. Find 2 paper clips. Put the paper clips close together and observe what happens.
   a. What happens to the paper clips?
   b. Are the paper clips magnetic?
2. Now take a permanent bar magnet and rub it once along 1 of the paper clips. Remove the magnet and put the paper clip which was touched by the magnet close to the other paper clip and observe what happens. Does the untouched paper clip feel a force on it? If so, is the force attractive or repulsive?
3. Rub the same paper clip a few more times with the bar magnet, in the same direction as before. Put the paper clip close to the other one and observe what happens.
   a. Is there any difference to what happened in step 2?
   b. If there is a difference, what is the reason for it?
   c. Is the paper clip which was rubbed repeatedly by the magnet now magnetised?
   d. What is the difference between the two paper clips at the level of their atoms and electrons?

4. Now, find a *metal* knitting needle, or a metal ruler, or other metal object. Rub the bar magnet along the knitting needle a few times in the same direction. Now put the knitting needle close to the paper clips and observe what happens.
   a. Does the knitting needle attract the paper clips?
   b. What does this tell you about the material of the knitting needle? Is it ferromagnetic?
5. Repeat this experiment with objects made from other materials. Which materials appear to be ferromagnetic and which are not? Put your answers in a table.

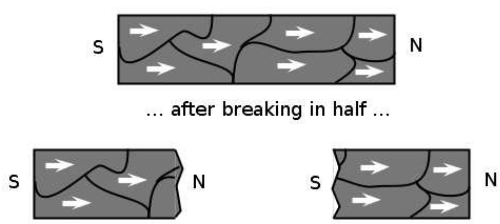
A ferromagnetic material is a substance that shows spontaneous magnetisation.

1.161 Permanent magnets

(section shortcode: P10073)

1.161.1 The poles of permanent magnets

Because the domains in a permanent magnet all line up in a particular direction, the magnet has a pair of opposite poles, called **north** (usually shortened to **N**) and **south** (usually shortened to **S**). Even if the magnet is cut into tiny pieces, each piece will still have *both* a N and a S pole. These magnetic poles *always* occur in pairs. In nature, we never find a north magnetic pole or south magnetic pole on its own.



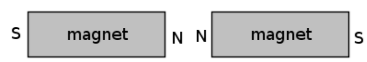
Magnetic fields are *different* from gravitational and electric fields. In nature, positive and negative electric charges can be found on their own, but you *never* find just a north magnetic pole or south magnetic pole on its own. On the very small scale, zooming in to the size of atoms, magnetic fields are caused by moving charges (i.e. the negatively charged electrons).

1.161.2 Magnetic attraction and repulsion

Like (identical) poles of magnets repel one another whilst unlike (opposite) poles attract. This means that two N poles or two S poles will push away from each other while a N pole and a S pole will be drawn towards each other.

Definition: Attraction and Repulsion
Like poles of magnets *repel* each other whilst *unlike* poles *attract* each other.

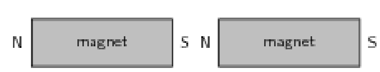
Exercise 21.1: Attraction and Repulsion Do you think the following magnets will repel or be attracted to each other?



Solution to Exercise

Step 1. We are required to determine whether the two magnets will repel each other or be attracted to each other.
Step 2. We are given two magnets with the N pole of one approaching the N pole of the other.
Step 3. Since both poles are the same, the magnets will repel each other.

Exercise 21.2: Attraction and repulsion Do you think the following magnets will repel or be attracted to each other?

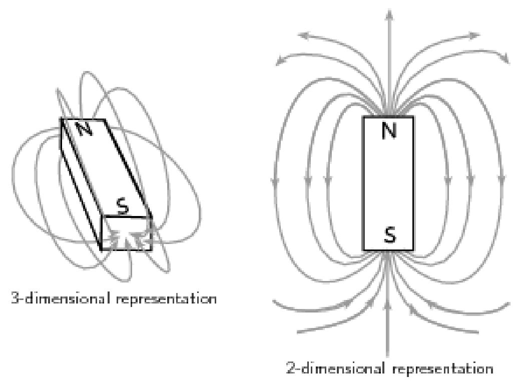


Solution to Exercise

Step 1. We are required to determine whether the two magnets will repel each other or be attracted to each other.
Step 2. We are given two magnets with the N pole of one approaching the S pole of the other.
Step 3. Since both poles are the different, the magnets will be attracted to each other.

1.161.3 Representing magnetic fields

Magnetic fields can be *represented* using **magnetic field lines** starting at the North pole and ending at the South pole. Although the magnetic field of a permanent magnet is everywhere surrounding the magnet (in all three dimensions), we draw only some of the field lines to represent the field (usually only a two-dimensional cross-section is shown in drawings).



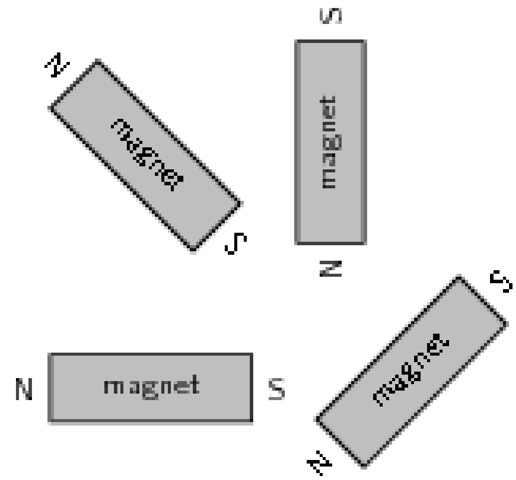
In areas where the magnetic field is strong, the field lines are closer together. Where the field is weaker, the field lines are drawn further apart. The number of field lines drawn crossing a given two-dimensional surface is referred to as the **magnetic flux**. The magnetic flux is used as a measure of the strength of the magnetic field over that surface.

TIP:

1.Field lines *never* cross.
2.Arrows drawn on the field lines indicate the direction of the field.
3.A magnetic field points from the north to the south pole of a magnet.

Investigation : Field around a Bar Magnet

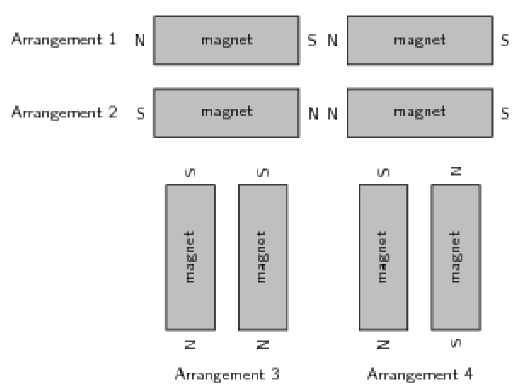
Take a bar magnet and place it on a flat surface. Place a sheet of white paper over the bar magnet and sprinkle some iron filings onto the paper. Give the paper a shake to evenly distribute the iron filings. In your workbook, draw the bar magnet and the pattern formed by the iron filings. Draw the pattern formed when you rotate the bar magnet to a different angle as shown below.



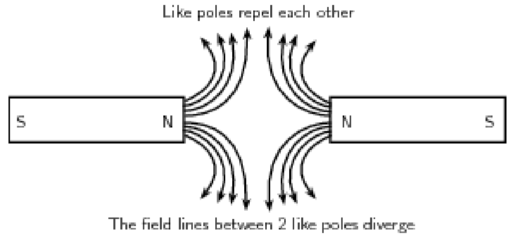
As the activity shows, one can map the magnetic field of a magnet by placing it underneath a piece of paper and sprinkling iron filings on top. The iron filings line themselves up parallel to the magnetic field.

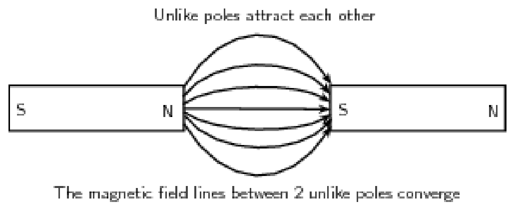
Investigation : Field around a Pair of Bar Magnets

Take two bar magnets and place them a short distance apart such that they are repelling each other. Place a sheet of white paper over the bar magnets and sprinkle some iron filings onto the paper. Give the paper a shake to evenly distribute the iron filings. In your workbook, draw both the bar magnets and the pattern formed by the iron filings. Repeat the procedure for two bar magnets attracting each other and draw what the pattern looks like for this situation. Make a note of the shape of the lines formed by the iron filings, as well as their size and their direction for both arrangements of the bar magnet. What does the pattern look like when you place both bar magnets side by side?



As already said, opposite poles of a magnet attract each other and bringing them together causes their magnetic field lines to *converge* (come together). Like poles of a magnet repel each other and bringing them together causes their magnetic field lines to *diverge* (bend out from each other).





Ferromagnetism and Retentivity

Ferromagnetism is a phenomenon shown by materials like iron, nickel or cobalt. These materials can form permanent magnets. They always magnetise so as to be attracted to a magnet, no matter which magnetic pole is brought toward the unmagnetised iron/nickel/cobalt.

The ability of a ferromagnetic material to retain its magnetisation *after* an external field is removed is called its **retentivity**.

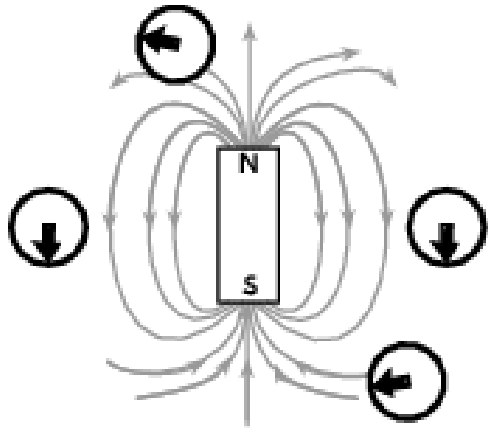
Paramagnetic materials are materials like aluminium or platinum, which become magnetised in an external magnetic field in a similar way to ferromagnetic materials. However, they lose their magnetism when the external magnetic field is removed.

Diamagnetism is shown by materials like copper or bismuth, which become magnetised in a magnetic field with a polarity *opposite* to the external magnetic field. Unlike iron, they are slightly repelled by a magnet.

1.162 The compass and the earth's magnetic field

(section shortcode: P10074)

A **compass** is an instrument which is used to find the direction of a magnetic field. A compass consists of a small metal needle which is magnetised itself and which is free to turn in any direction. Therefore, when in the presence of a magnetic field, the needle is able to line up in the same direction as the field.

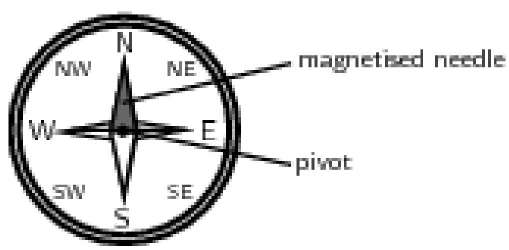


The direction of the compass arrow is the same as the direction of the magnetic field

Interesting Fact

NOTE: Lodestone, a magnetised form of iron-oxide, was found to orientate itself in a north-south direction if left free to rotate by suspension on a string or on a float in water. Lodestone was therefore used as an early navigational compass.

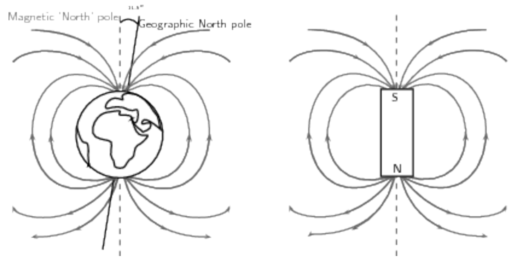
Compasses are mainly used in navigation to find direction on the earth. This works because the earth itself has a magnetic field which is similar to that of a bar magnet (see the picture below). The compass needle aligns with the earth's magnetic field direction and points north-south. Once you know where north is, you can figure out any other direction. A picture of a compass is shown below:



Some animals can detect magnetic fields, which helps them orientate themselves and navigate. Animals which can do this include pigeons, bees, Monarch butterflies, sea turtles and certain fish.

1.162.1 The earth's magnetic field

In the picture below, you can see a representation of the earth's magnetic field which is very similar to the magnetic field of a giant bar magnet like the one on the right of the picture. So the earth has two sets of north poles and south poles: **geographic poles** and **magnetic poles**.



The earth's magnetic field is thought to be caused by flowing liquid metals in the outer core which causes electric currents and a magnetic field. From the picture you can see that the direction of magnetic north and true north are not identical. The **geographic north pole**, which is the point through which the earth's rotation axis goes, is about 11,5^{o} away from the direction of the **magnetic north pole** (which is where a compass will point). However, the magnetic poles shift slightly all the time.

Another interesting thing to note is that if we think of the earth as a big bar magnet, and we know that magnetic field lines always point *from north to south*, then the compass tells us that what we call the *magnetic north pole* is actually the *south pole* of the bar magnet!

Interesting Fact

NOTE: The direction of the earth's magnetic field flips direction about once every 200 000 years! You can picture this as a bar magnet whose north and south pole periodically switch sides. The reason for this is still not fully understood.

Phenomena related to the Earth's magnetic field

The importance of the magnetic field to life on Earth

The earth's magnetic field is very important for humans and other animals on earth because it protects us from being bombarded (hit) by high energy charged particles which are emitted by the Sun. The stream of charged particles (mainly positively charged protons and negatively charged electrons) coming from the sun is called the solar wind. When these particles come close to the Earth, they become trapped by the Earth's magnetic field and cannot shower down to the surface where they can harm living organisms. Astronauts in space are at risk of being irradiated by the solar wind because they are outside the zones where the charged particles are trapped.

The region above Earth's atmosphere in which charged particles are affected by Earth's magnetic field is called the magnetosphere. Relatively often, in addition to the usual solar wind, the Sun may eject a large bubble of material (protons and electrons) with its own magnetic field from its outer atmosphere. Sometimes these bubbles travel towards the Earth where their magnetic fields can join with Earth's magnetic field. When this happens a huge amount of energy is released into the Earth's magnetosphere, causing a geomagnetic storm. These storms cause rapid changes in the Earth's magnetosphere which in turn may affect electric and magnetic systems on the Earth such as power grids, cellphone networks, and other electronic systems.

Aurorae (pronounced Or-roar-ee)

Another effect caused by the Earth's magnetic field is the spectacular Northern and Southern Lights, which are also called the Aurora Borealis and the Aurora Australis respectively. When charged particles from the solar wind reach the Earth's magnetosphere, they spiral along the magnetic field lines towards the North and South poles. If they collide with particles in the Earth's atmosphere, they can cause red or green lights which stretch across a large part of the sky and which is called the aurora. Because this only happens close to the North and South poles, we cannot see the aurorae from South Africa. However, people living in the high Northern latitudes in Canada, Sweden, and Finland, for example, often see the Northern lights.

This simulation shows you the Earth's magnetic field and a compass. www (Simulation: lb9)

1.163 Summary

(section shortcode: P10075)

1. Magnets have two poles - North and South.
2. Some substances can be easily magnetised.
3. Like poles repel each other and unlike poles attract each other.
4. The Earth also has a magnetic field.
5. A compass can be used to find the magnetic north pole and help us find our direction.

This video provides a summary of the work covered in this chapter.

Khan academy video on magnets www (Video: EZURLhttp://www.youtube.com/v/8Y4JSp5U82I&rel=0)

1.164 End of chapter exercises

www (section shortcode: P10077)

1. Describe what is meant by the term *magnetic field*.
2. Use words and pictures to explain why permanent magnets have a magnetic field around them. Refer to *domains* in your explanation.
3. What is a magnet?
4. What happens to the poles of a magnet if it is cut into pieces?
5. What happens when like magnetic poles are brought close together?
6. What happens when unlike magnetic poles are brought close together?

7. Draw the shape of the magnetic field around a bar magnet.
8. Explain how a compass indicates the direction of a magnetic field.
9. Compare the magnetic field of the Earth to the magnetic field of a bar magnet using words and diagrams.
10. Explain the difference between the geographical north pole and the magnetic north pole of the Earth.
11. Give examples of phenomena that are affected by Earth's magnetic field.
12. Draw a diagram showing the magnetic field around the Earth.

www Find the answers with the shortcodes:

(1.) llS (2.) lia (3.) lix (4.) lic (5.) liO (6.) li3
(7.) lii (8.) llu (9.) lil (10.) llh (11.) llJ (12.) llt

Electrostatics

1.165 Introduction

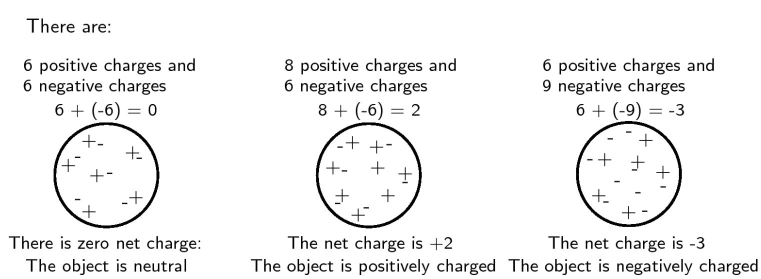
(section shortcode: P10078)

Electrostatics is the study of electric charge which is static (not moving). In this chapter we will look at some of the basic principle of electric charge as well as the principle of conservation of charge.

1.166 Two kinds of charge

(section shortcode: P10079)

All objects surrounding us (including people!) contain large amounts of electric charge. There are two types of electric charge: **positive** charge and **negative** charge. If the same amounts of negative and positive charge are brought together, they neutralise each other and there is *no net charge*. **Neutral** objects are objects which contain equal amouts of positive and negative charges. However, if there is a little bit more of one type of charge than the other on the object then the object is said to be **electrically charged**. The picture below shows what the distribution of charges might look like for a neutral, positively charged and negatively charged object.



1.167 Unit of charge

(section shortcode: P10080)

Charge is measured in units called **coulombs (C)**. A coulomb of charge is a very large charge. In electrostatics we therefore often work with charge in microcoulombs ($1\,\mu\,\mathrm{C} = 1 \times 10^{-6}\,\mathrm{C}$) and nanocoulombs ($1\ \mathrm{nC} = 1 \times 10^{-9}\,\mathrm{C}$).

1.168 Conservation of charge

(section shortcode: P10081)

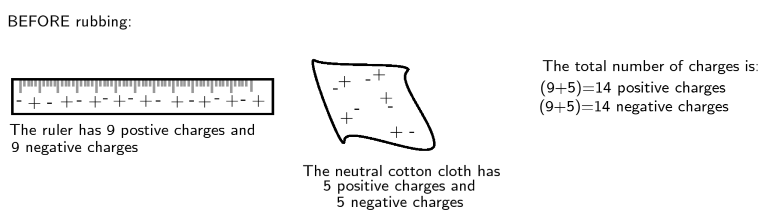
Objects may become charged in many ways, including by contact with or being rubbed by other objects. This means that they can gain extra negative or positive charge. For example, charging happens when you rub your feet against the carpet. When you then touch something metallic or another person, you feel a shock as the excess charge that you have collected is *discharged*.

> **TIP:** Charge, like energy, cannot be created or destroyed. We say that charge is **conserved**.

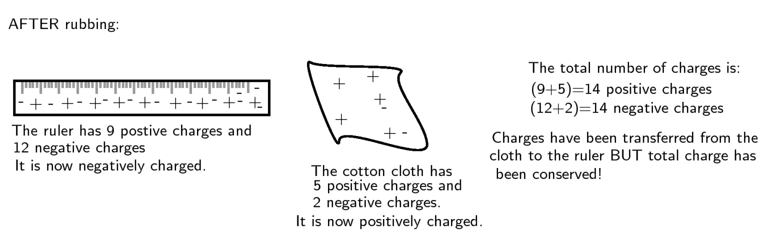
The principle of conservation of charge states that the net charge of an isolated system remains constant during any physical process, e.g. when two charges make contact and are separated again.

When you rub your feet against the carpet, negative charge is transferred to you from the carpet. The carpet will then become positively charged by the *same amount*.

Another example is to take two *neutral* objects such as a plastic ruler and a cotton cloth (handkerchief). To begin, the two objects are neutral (i.e. have the same amounts of positive and negative charge).



Now, if the cotton cloth is used to rub the ruler, negative charge is transferred *from* the cloth *to* the ruler. The ruler is now *negatively* charged (i.e. has an excess of electrons) and the cloth is *positively* charged (i.e. is electron deficient). If you count up all the positive and negative charges at the beginning and the end, there are still the same amount. i.e. total charge has been *conserved*!



Note that in this example the numbers are made up to be easy to calculate. In the real world only a tiny fraction of the charges would move from one object to the other, but the total charge would still be conserved.

The following simulation will help you understand what happens when you rub an object against another object. (Simulation: lbX)

Interesting Fact

NOTE: The process of materials becoming charged when they come into contact with other materials is known as tribo-electric charging. Materials can be arranged in a tribo-electric series according to whether they are more positive or more negative. This tribo-electric series can allow us to determine whether one material is likely to become charged from another material. For example, amber is more negative than wool and so if a piece of wool is rubbed against a piece of amber then the amber will become negatively charged.

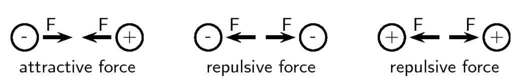
1.169 Force between Charges

(section shortcode: P10082)

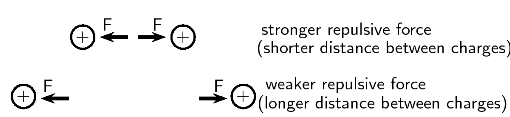
The force exerted by non-moving (static) charges on each other is called the **electrostatic force.** The electrostatic force between:

- **like** charges is **repulsive**
- **opposite** (unlike) charges is **attractive**.

In other words, like charges repel each other while opposite charges attract each other. This is different to the gravitational force which is only attractive.



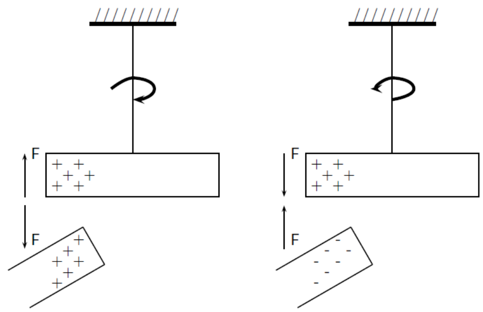
The *closer* together the charges are, the *stronger* the electrostatic force between them.



1.169.1 Experiment : Electrostatic Force

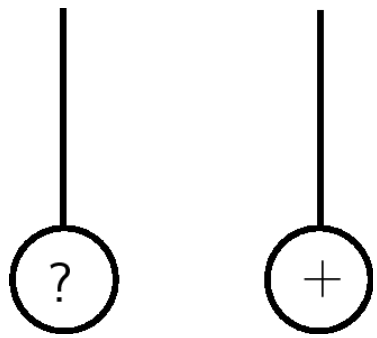
You can easily test that like charges repel and unlike charges attract each other by doing a very simple experiment.

Take a glass rod and rub it with a piece of silk, then hang it from its middle with a piece string so that it is free to move. If you then bring another glass rod which you have also charged in the same way next to it, you will see the rod on the string turn *away* from the rod in your hand i.e. it is **repelled**. If, however, you take a plastic rod, rub it with a piece of fur and then bring it close to the rod on the string, you will see the rod on the string turn *towards* the rod in your hand i.e. it is **attracted**.



This happens because when you rub the glass with silk, tiny amounts of negative charge are transferred from the glass onto the silk, which causes the glass to have less negative charge than positive charge, making it **positively charged**. When you rub the plastic rod with the fur, you transfer tiny amounts of negative charge onto the rod and so it has more negative charge than positive charge on it, making it **negatively charged**.

Exercise 22.1: Application of electrostatic forces Two charged metal spheres hang from strings and are free to move as shown in the picture below. The right hand sphere is positively charged. The charge on the left hand sphere is unknown.



The left sphere is now brought close to the right sphere.

1. If the left hand sphere swings towards the right hand sphere, what can you say about the charge on the left sphere and why?
2. If the left hand sphere swings away from the right hand sphere, what can you say about the charge on the left sphere and why?

Solution to Exercise

Step 1. In the first case, we have a sphere with positive charge which is *attracting* the left charged sphere. We need to find the charge on the left sphere.

Step 2. We are dealing with electrostatic forces between charged objects. Therefore, we know that *like* charges *repel* each other and *opposite* charges *attract* each other.

Step 3.
a. In the first case, the positively charged sphere is attracting the left sphere. Since an electrostatic force between unlike charges is attractive, the left sphere must be *negatively* charged.
b. In the second case, the positively charged sphere repels the left sphere. Like charges repel each other. Therefore, the left sphere must now also be *positively* charged.

Interesting Fact

NOTE: The word 'electron' comes from the Greek word for amber. The ancient Greeks observed that if you rubbed a piece of amber, you could use it to pick up bits of straw.

1.170 Conductors and insulators

(section shortcode: P10083)

All atoms are electrically neutral i.e. they have the same amounts of negative and positive charge inside them. By convention, the electrons carry negative charge and the protons carry positive charge. The basic unit of charge, called the elementary charge, *e*, is the amount of charge carried by one electron.

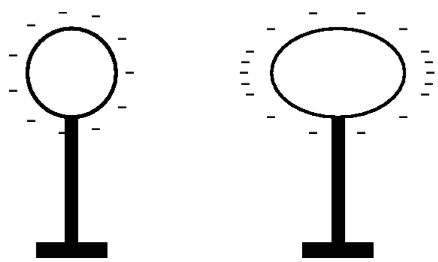
The charge on a single electron is $q_e = 1,6x10^{-19}\,\mathrm{C}$. All other charges in the universe consist of an interger multiple of this charge (i.e. $\mathrm{Q} = \mathrm{nq}_e$). This is known as charge quantisation.

All the matter and materials on earth are made up of atoms. Some materials allow electrons to move relatively freely through them (e.g. most metals, the human body). These materials are called **conductors**.

Other materials do not allow the charge carriers, the electrons, to move through them (e.g. plastic, glass). The electrons are bound to the atoms in the material. These materials are called **non-conductors** or **insulators**.

If an excess of charge is placed on an insulator, it will stay where it is put and there will be a concentration of charge in that area of the object. However, if an excess of charge is placed on a conductor, the like charges will repel each other and spread out over the outside surface of the object. When two conductors are made to touch, the total charge on them is shared between the two. If the two conductors are identical, then each conductor will be left with half of the total charge.

NOTE: The electrostatic force determines the arrangement of charge on the surface of conductors. This is possible because charges can move inside a conductive material. When we place a charge on a spherical conductor the repulsive forces between the individual like charges cause them to spread uniformly over the surface of the sphere. However, for conductors with non-regular shapes, there is a concentration of charge near the point or points of the object. Notice in Figure 22.9 that we show a concentration of charge with more $-$ or + signs, while we represent uniformly spread charges with uniformly spaced $-$ or + signs.



This collection of charge can actually allow charge to leak off the conductor if the point is sharp enough. It is for this reason that buildings often have a lightning rod on the roof to remove any charge the building has collected. This minimises the possibility of the building being struck by lightning. This "spreading out" of charge would not occur if we were to place the charge on an insulator since charge cannot move in insulators.

ASIDE: The basic unit of charge, namely the elementary charge is carried by the electron (equal to 1.602×10^{-19} C!). In a conducting material (e.g. copper), when the atoms bond to form the material, some of the outermost, loosely bound electrons become detached from the individual atoms and so become free to move around. The charge carried by these electrons can move around in the material. In insulators, there are very few, if any, free electrons and so the charge cannot move around in the material.

NOTE: In 1909 Robert Millikan and Harvey Fletcher measured the charge on an electron. This experiment is now known as Millikan's oil drop experiment. Millikan and Fletcher sprayed oil droplets into the space between two charged plates and used what they knew about forces and in particular the electric force to determine the charge on an electron.

Exercise 22.2: Conducting spheres and movement of charge I have 2 charged metal conducting spheres which are identical except for having different charge. Sphere A has a charge of -5 nC and sphere B has a charge of -3 nC. I then bring the spheres together so that they touch each other. Afterwards I move the two spheres apart so that they are no longer touching.

1. What happens to the charge on the two spheres?
2. What is the final charge on each sphere?

Solution to Exercise

Step 1. We have two identical negatively charged conducting spheres which are brought together to touch each other and then taken apart again. We need to explain what happens to the charge on each sphere and what the final charge on each sphere is after they are moved apart.

Step 2. We know that the charge carriers in conductors are free to move around and that charge on a conductor spreads itself out on the surface of the conductor.

Step 3.

a. When the two conducting spheres are brought together to touch, it is as though they become one single big conductor and the total charge of the two spheres spreads out across the whole surface of the touching spheres. When the spheres are moved apart again, each one is left with half of the total original charge.

b. Before the spheres touch, the total charge is: -5 nC + (-3) nC = -8 nC. When they touch they share out the -8 nC across their whole surface. When they are removed from each other, each is left with half of the original charge:

$$-8 \text{ nC } /2 \quad = \quad -4 \text{ nC} \tag{22.1}$$

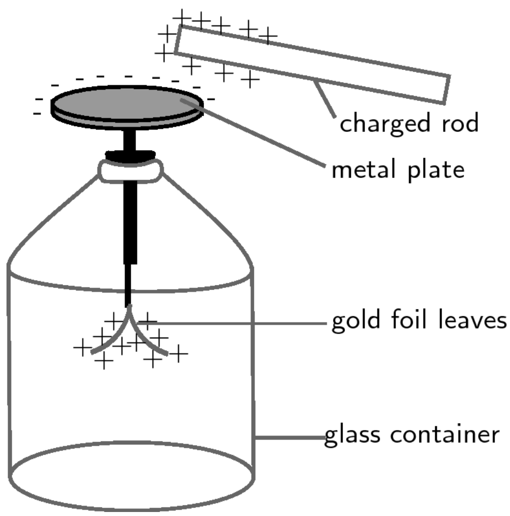
on each sphere.

In the previous example we worked out what happens when two identical conductors are allowed to touch. We noticed that if we take two identically sized conducting spheres on insulating stands and bring them together so that they touch, each sphere will have the same final charge. If the initial charge on the first sphere is Q_1 and the initial charge on the second sphere is Q_2, then the final charge on the two spheres after they have been brought into contact is:

$$Q = \frac{Q_1 + Q_2}{2} \tag{22.2}$$

1.170.1 The electroscope

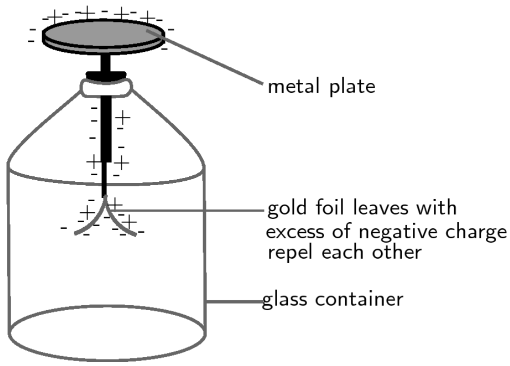
The electroscope is a very sensitive instrument which can be used to detect electric charge. A diagram of a gold leaf electroscope is shown the figure below. The electroscope consists of a glass container with a metal rod inside which has 2 thin pieces of gold foil attached. The other end of the metal rod has a metal plate attached to it outside the glass container.



The electroscope detects charge in the following way: A charged object, like the positively charged rod in the picture, is brought close to (but not touching) the neutral metal plate of the electroscope. This causes negative charge in the gold foil, metal rod, and metal plate, to be attracted to the positive rod. Because the metal (gold is a metal too!) is a conductor, the charge can move freely from the foil up the metal rod and onto the metal plate. There is now more negative charge on the plate and more positive charge on the gold foil leaves. This is called *inducing* a charge on the metal plate. It is important to remember that the electroscope is still neutral (the total positive and negative charges are the same), the charges have just been induced to *move* to different parts of the instrument! The induced positive charge on the gold leaves forces them apart since like charges repel! This is how we can tell that the rod is charged. If the rod is now moved away from the metal plate, the charge in the electroscope will spread itself out evenly again and the leaves will fall down because there will no longer be an induced charge on them.

Grounding

If you were to bring the charged rod close to the uncharged electroscope, and then you touched the metal plate with your finger at the same time, this would cause charge to flow up from the ground (the earth), through your body onto the metal plate. Connecting to the earth so charge flows is called **grounding**. The charge flowing onto the plate is opposite to the charge on the rod, since it is attracted to the charge on the rod. Therefore, for our picture, the charge flowing onto the plate would be negative. Now that charge has been added to the electroscope, it is no longer neutral, but has an excess of negative charge. Now if we move the rod away, the leaves will remain apart because they have an excess of negative charge and they repel each other. If we ground the electroscope again (this time without the charged rod nearby), the excess charge will flow back into the earth, leaving it neutral.



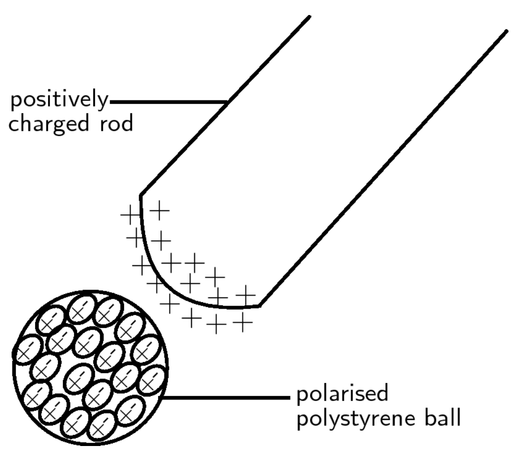
1.171 Attraction between charged and uncharged objects

(section shortcode: P10084)

1.171.1 Polarisation of Insulators

Unlike conductors, the electrons in insulators (non-conductors) are bound to the atoms of the insulator and cannot move around freely through the material. However, a charged object can still exert a force on a neutral insulator due to a phenomenon called **polarisation**.

If a positively charged rod is brought close to a neutral insulator such as polystyrene, it can attract the bound electrons to move round to the side of the atoms which is closest to the rod and cause the positive nuclei to move slightly to the opposite side of the atoms. This process is called *polarisation*. Although it is a very small (microscopic) effect, if there are many atoms and the polarised object is light (e.g. a small polystyrene ball), it can add up to enough force to cause the object to be attracted onto the charged rod. Remember, that the polystyrene is *only* polarised, *not charged.* The polystyrene ball is still neutral since no charge was added or removed from it. The picture shows a not-to-scale view of the polarised atoms in the polystyrene ball:



Some materials are made up of molecules which are already polarised. These are molecules which have a more positive and a more negative side but are still neutral overall. Just as a polarised polystyrene ball can be attracted to a charged rod, these materials are also affected if brought close to a charged object.

Water is an example of a substance which is made of polarised molecules. If a positively charged rod is brought close to a stream of water, the molecules can rotate so that the negative sides all line up towards the rod. The stream of water will then be attracted to the rod since opposite charges attract.

1.172 Summary

(section shortcode: P10085)

1. Objects can be **positively** charged, **negatively** charged or **neutral**.
2. Objects that are neutral have equal numbers of positive and negative charge.
3. Unlike charges are attracted to each other and like charges are repelled from each other.
4. Charge is neither created nor destroyed, it can only be transferred.
5. Charge is measured in coulombs (C).
6. Conductors allow charge to move through them easily.
7. Insulators do not allow charge to move through them easily.

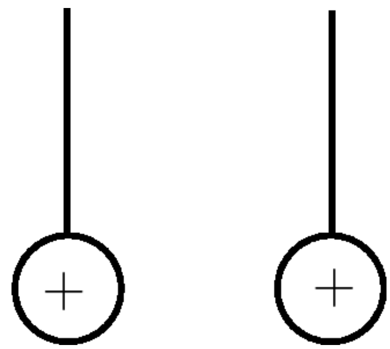
The following presentation is a summary of the work covered in this chapter. Note that the last two slides are not needed for exam purposes, but are included for general interest.

(Presentation: P10086)

1.173 End of chapter exercise

(section shortcode: P10087)

1. What are the two types of charge called?
2. Provide evidence for the existence of two types of charge.
3. Fill in the blanks: The electrostatic force between like charges is __________ while the electrostatic force between opposite charges is __________ .
4. I have two positively charged metal balls placed 2 m apart.
   a. Is the electrostatic force between the balls attractive or repulsive?
   b. If I now move the balls so that they are 1 m apart, what happens to the strength of the electrostatic force between them?
5. I have 2 charged spheres each hanging from string as shown in the picture below.



Choose the correct answer from the options below: The spheres will

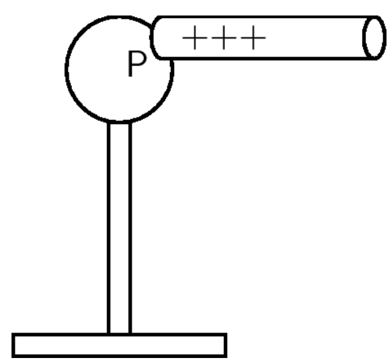
- **a.** swing towards each other due to the attractive electrostatic force between them.
- **b.** swing away from each other due to the attractive electrostatic force between them.
- **c.** swing towards each other due to the repulsive electrostatic force between them.
- **d.** swing away from each other due to the repulsive electrostatic force between them.

6. Describe how objects (insulators) can be charged by contact or rubbing.

7. You are given a perspex ruler and a piece of cloth.

- **a.** How would you charge the perspex ruler?
- **b.** Explain how the ruler becomes charged in terms of charge.
- **c.** How does the charged ruler attract small pieces of paper?

8. (IEB 2005/11 HG) An uncharged hollow metal sphere is placed on an insulating stand. A positively charged rod is brought up to touch the hollow metal sphere at P as shown in the diagram below. It is then moved away from the sphere.



Where is the excess charge distributed on the sphere after the rod has been removed?

- **a.** It is still located at point P where the rod touched the sphere.
- **b.** It is evenly distributed over the outer surface of the hollow sphere.
- **c.** It is evenly distributed over the outer and inner surfaces of the hollow sphere.
- **d.** No charge remains on the hollow sphere.

9. What is the process called where molecules in an uncharged object are caused to align in a particular direction due to an external charge?

10. Explain how an uncharged object can be attracted to a charged object. You should use diagrams to illustrate your answer.

11. Explain how a stream of water can be attracted to a charged rod.

Find the answers with the shortcodes:

(1.) lqs (2.) lqo (3.) lqA (4.) lqG (5.) lqf (6.) lqw
(7.) lqv (8.) lqd (9.) lqp (10.) la2 (11.) lqP

Electric Circuits

1.174 Electric Circuits

(section shortcode: P10089)

People all over the world depend on electricity to provide power for most appliances in the home and at work. For example, fluorescent lights, electric heating and cooking (on electric stoves), all depend on electricity to work. To realise just how big an impact electricity has on our daily lives, just think about what happens when there is a power failure or load shedding.

1.174.1 Discussion : Uses of electricity

With a partner, take the following topics and, for each topic, write down at least 5 items/appliances/machines which need electricity to work. Try not to use the same item more than once.

- At home
- At school
- At the hospital
- In the city

Once you have finished making your lists, compare with the lists of other people in your class. (Save your lists somewhere safe for later because there will be another activity for which you'll need them.)

When you start comparing, you should notice that there are many different items which we use in our daily lives which rely on electricity to work!

> TIP: **Safety Warning:** We believe in experimenting and learning about physics at every opportunity, BUT playing with electricity and electrical appliances can be **EXTREMELY DANGEROUS**! Do not try to build homemade circuits alone. Make sure you have someone with you who knows if what you are doing is safe. Normal electrical outlets are dangerous. Treat electricity with respect in your everyday life. Do not touch exposed wires and do not approach downed power lines.

1.174.2 Closed circuits

In the following activity we will investigate what is needed to cause charge to flow in an electric circuit.

Experiment : Closed circuits

Aim: To determine what is required to make electrical charges flow. In this experiment, we will use a lightbulb to check whether electrical charge is flowing in the circuit or not. If charge is flowing, the lightbulb should glow. On the other hand, if no charge is flowing, the lightbulb will not glow.

Apparatus: You will need a small lightbulb which is attached to a metal conductor (e.g. a bulb from a school electrical kit), some connecting wires and a battery.

Method: Take the apparatus items and try to connect them in a way that you cause the light bulb to glow (i.e. charge flows in the circuit).

Questions:

1. Once you have arranged your circuit elements to make the lightbulb glow, draw your circuit.
2. What can you say about how the battery is connected? (i.e. does it have one or two connecting leads attached? Where are they attached?)
3. What can you say about how the light bulb is connected in your circuit? (i.e. does it connect to one or two connecting leads, and where are they attached?)
4. Are there any items in your circuit which are not attached to something? In other words, are there any gaps in your circuit?

Write down your conclusion about what is needed to make an electric circuit work and charge to flow.

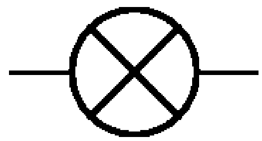
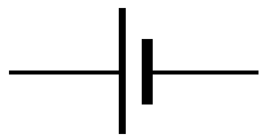
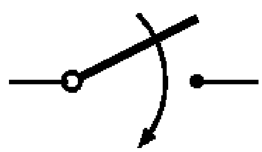
In the experiment above, you will have seen that the light bulb only glows when there is a *closed* circuit i.e. there are no gaps in the circuit and all the circuit elements are connected in a *closed loop*. Therefore, in order for charges to flow, a closed circuit and an energy source (in this case the battery) are needed. (Note: you do not have to have a lightbulb in the circuit! We used this as a check that charge was flowing.)

Definition: Electric circuit
An electric circuit is a closed path (with no breaks or gaps) along which electrical charges (electrons) flow powered by an energy source.

1.174.3 Representing electric circuits

Components of electrical circuits

Some common elements (components) which can be found in electrical circuits include light bulbs, batteries, connecting leads, switches, resistors, voltmeters and ammeters. You will learn more about these items in later sections, but it is important to know what their symbols are and how to represent them in circuit diagrams. Below is a table with the items and their symbols:

Component	Symbol	Usage
light bulb		glows when charge moves through it
battery	 **Figure 23.2:** The longer line shows the positive (+) side of the battery, the shorter line shows the (-) side of the battery.	provides energy for charge to move - conventional current flow from positive to negative through a circuit
switch		allows a circuit to be open or closed
resistor		resists the flow of charge
	OR	
		

continued on next page

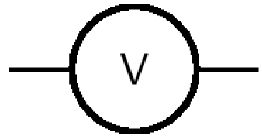
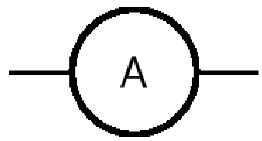
voltmeter		measures potential difference
ammeter		measures current in a circuit
connecting lead		connects circuit elements together

Table 23.1

Circuit diagrams

Definition: Representing circuits
A **physical circuit** is the electric circuit you create with real components.
A **circuit diagram** is a drawing which uses symbols to represent the different components in the physical circuit.

We use circuit diagrams to represent circuits because they are much simpler and more general than drawing the physical circuit because they only show the workings of the electrical components. You can see this in the two pictures below. The first picture shows the *physical circuit* for an electric torch. You can see the light bulb, the batteries, the switch and the outside plastic casing of the torch. The picture is actually a *cross-section* of the torch so that we can see inside it.

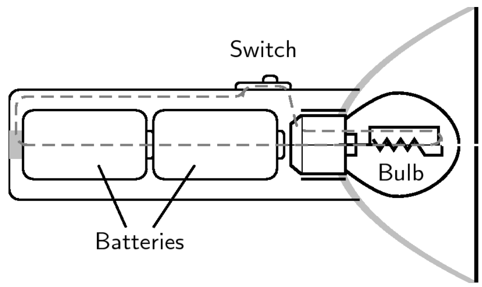


Figure 23.9: Physical components of an electric torch. The dotted line shows the path of the electrical circuit.

Below is the *circuit diagram* for the electric torch. Now the light bulb is represented by its symbol, as are the batteries, the switch and the connecting wires. It is not necessary to show the plastic casing of the torch since it has nothing to do with the electric workings of the torch. You can see that the circuit diagram is much simpler than the physical circuit drawing!

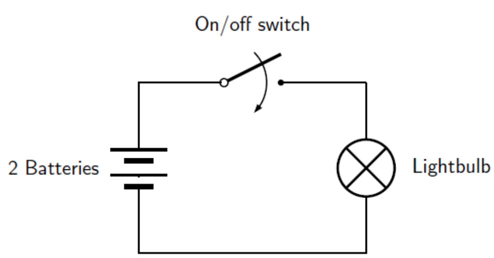
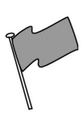


Figure 23.10: Circuit diagram of an electric torch.

Series and parallel circuits

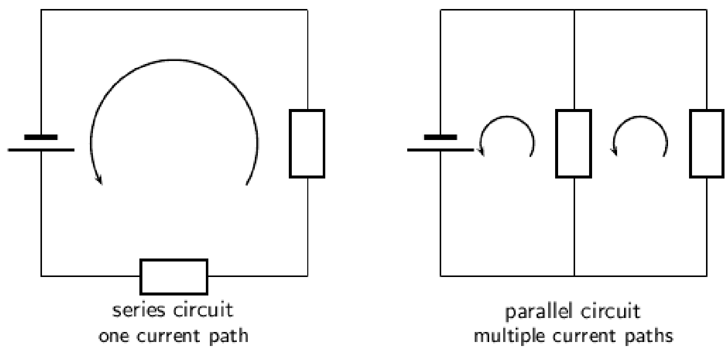
There are two ways to connect electrical components in a circuit: **in series** or **in parallel**.



Definition: Series circuit
In a series circuit, the charge flowing from the battery can only flow along a **single** path to return to the battery.

Definition: Parallel circuit
In a parallel circuit, the charge flowing from the battery can flow along **multiple** paths to return to the battery.

The picture below shows a circuit with two resistors connected *in series* on the left and a circuit with two resistors connected *in parallel* on the right. In the series circiut, the charge path from the battery goes through every component before returning to the battery. In the parallel circuit, there is more than one path for the charge to flow from the battery through one of the components and back to the battery.



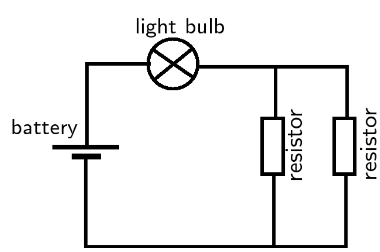
This simulation allows you to experiment with building circuits. www (Simulation: lbK)

Exercise 23.1: Drawing circuits I Draw the circuit diagram for a circuit which has the following components:

1. 1 battery
2. 1 lightbulb connected in series
3. 2 resistors connected in parallel

Solution to Exercise

Step 1.

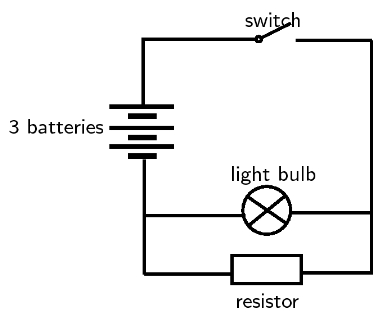


Exercise 23.2: Drawing circuits II Draw the circuit diagram for a circuit which has the following components:

1. 3 batteries in series
2. 1 lightbulb connected in parallel with 1 resistor
3. a switch in series with the batteries

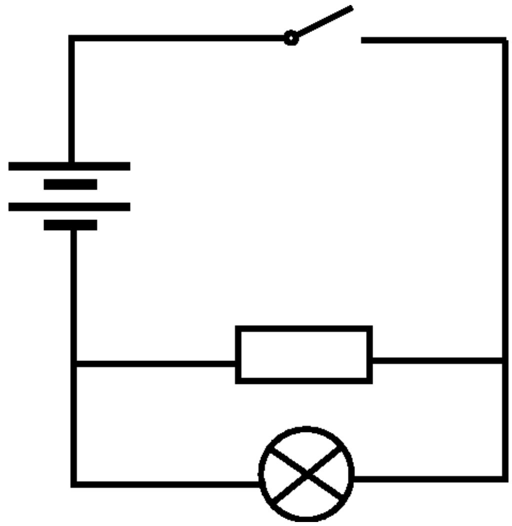
Solution to Exercise

Step 1.



Circuits

1. Using physical components, set up the physical circuit which is described by the circuit diagram below and then draw the physical circuit:



2. Using physical components, set up a closed circuit which has one battery and a light bulb in series with a resistor.
   a. Draw the physical circuit.
   b. Draw the resulting circuit diagram.
   c. How do you know that you have built a closed circuit? (What happens to the light bulb?)
   d. If you add one more resistor to your circuit (also in series), what do you notice? (What happens to the light from the light bulb?)
   e. Draw the new circuit diagram which includes the second resistor.
3. Draw the circuit diagram for the following circuit: 2 batteries and a switch in series, and 1 lightbulb which is in parallel with two resistors.
   a. Now use physical components to set up the circuit.
   b. What happens when you close the switch? What does does this mean about the circuit?
   c. Draw the physical circuit.

Find the answers with the shortcodes:

(1.) lqZ (2.) lqB (3.) lbY

Discussion : Alternative Energy

At the moment, most electric power is produced by burning fossil fuels such as coal and oil. In South Africa, our main source of electric power is coal-burning power stations. (We also have one nuclear power plant called Koeberg in the Western Cape). However, burning fossil fuels releases large amounts of pollution into the earth's atmosphere and contributes to global warming. Also, the earth's fossil fuel reserves (especially oil) are starting to run low. For these reasons, people all across the world are working to find *alternative*/other sources of energy and on ways to *conserve*/save energy. Other sources of energy include wind power, solar power (from the sun), and hydro-electric power (from water, e.g. dammed rivers) among others.

With a partner, take out the lists you made earlier of the item/appliances/machines which used electricity in the following environments. For each item, try to think of an *alternative* AND a way to *conserve* or save power.

For example, if you had a flourescent light as an item used in the home, then:

- Alternative: use candles at suppertime to reduce electricity consumption
- Conservation: turn off lights when not in a room, or during the day.

Topics:

- At home
- At school
- At the hospital
- In the city

Once you have finished making your lists, compare with the lists of other people in your class.

1.175 Potential Difference

(section shortcode: P10090)

1.175.1 Potential Difference

When a circuit is connected and complete, charge can move through the circuit. Charge will not move unless there is something to make it move. Think of it as though charge is at rest and something has to push it along. This means that work needs to be done to make charge move. A force acts on the charges, doing work, to make them move. The force is provided by the battery in the circuit.

We call the moving charge "current" and we will talk about this later.

The amount of work done to move a charge from one point to another point in a circuit is called the potential difference between those two points. You can think of it as a difference in the potential energy of the charge due to its position in the circuit. The difference in potential energy, called potential difference, is equal to the amount of work done to move the charge between the two points.Just like with gravity, when you raise an object above the ground it has gravitational potential energy due to its position, the same goes for a charge in a circuit and electrical potential energy. Potential difference is measured *between* or *across* two points. We do not say potential difference *through* something.

Definition: Potential Difference
Electrical potential difference is the difference in electrical potential energy per unit charge between two points. The unit of potential difference is the volt[a] (V). The potential difference of a battery is the voltage measured across it when current is flowing through it.

[a]named after the Italian physicist Alessandro Volta (1745–1827)

The unit of potential difference is the volt (V), which is the same as 1 joule per coulomb, the amount of work done per unit charge. Electrical potential difference is also called voltage.

1.175.2 Potential difference and emf

We use an instrument called a voltmeter to measure the potential difference between two points in a circuit. It must be connected across the two points, *in parallel* to that portion of the circuit as shown in the diagram below.

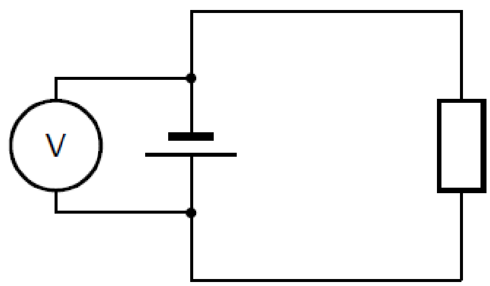


Figure 23.16: A voltmeter should be connected in parallel in a circuit.

EMF

When you use a voltmeter to measure the potential difference across (or between) the terminals of a battery, when no current is flowing through the battery, you are measuring the **electromotive force (emf)** of the battery. This is how much potential energy the battery has to make charges move through the circuit. It is a measure of how much chemical potential energy can be transferred to electrical energy in the battery. This driving potential energy is equal to the total potential energy drops in the circuit. This means that the voltage across the battery is equal to the sum of the voltages in the circuit.

Definition: emf
The emf (electromotive force) is the voltage measured across the terminals of a battery when *no current is flowing* through the battery.

You have now learnt that the emf is the voltage measured across the terminals of a battery when there is no current flowing through it and that the potential difference of a battery is the voltage measured across it when there is current flowing through it. So how do these two quantities compare with each other?

Experiment : Investigate the emf and potential difference of a battery

Aim: To measure the emf and potential difference across a battery in a circuit

Apparatus: A battery, connecting wires, a light bulb, a switch, a voltmeter.

Method: Set up a circuit with a battery, a lightbulb and a switch connected in series.

First we will measure the emf of the battery when no current is flowing. Make sure that the switch is open and the lightbulb is not glowing. This is how we check that there is no current flowing through the circuit. Then connect the voltmeter across the terminals of the battery. Make sure that the voltmeter is set to the largest scale when you first start measuring. Write down the voltage you measure. This is the emf. Disconnect the voltmeter.

Second we will measure the potential difference across the battery. This time, make sure the switch is closed and that the lightbulb is glowing. This means that there is current flowing through the circuit. Now connect the voltmeter again across the terminals of the battery. Again make sure that the voltmeter is set on its largest scale. Write down the voltage you measure. This is called the potential difference of the battery. How does it compare to the other value you measured when there was no current flowing?

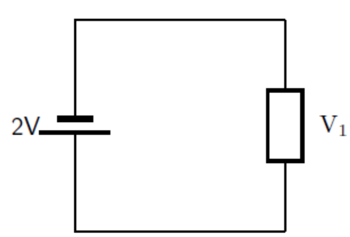
Summary: You will have noticed that the voltages you measured when there was no current flowing (emf) and when there was current flowing (potential difference) were different. The emf was a higher value than the potential difference. Discuss with your classmates and teacher why you think this might happen.

Batteries all have some internal resistance to charges moving through them and therefore some work is required to move the charges through the battery itself. This internal resistance causes the potential difference across the battery terminals to be slightly less than the emf.

In the following exercises, we will assume that we have 'perfect' batteries with no internal resistance. In this special case, the potential difference of the battery and its emf will be the same. We can use this information to solve problems in which the voltages across elements in a circuit add up to the emf.

$$EMF = V_{total} \tag{23.1}$$

Exercise 23.3: Voltages I What is the voltage across the resistor in the circuit shown?



Solution to Exercise

Step 1. We have a circuit with a battery and one resistor. We know the voltage across the battery. We want to find the voltage across the resistor.

$$V_{battery} = 2\text{V} \quad (23.2)$$

Step 2. We know that the voltage across the battery must be equal to the total voltage across all other circuit components.

$$V_{battery} = V_{total} \quad (23.3)$$

There is only one other circuit component, the resistor.

$$V_{total} = V_1 \quad (23.4)$$

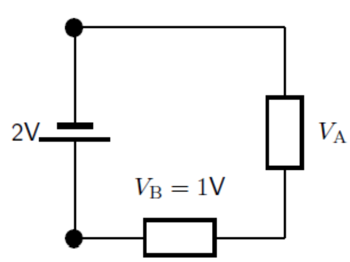
This means that the voltage across the battery is the same as the voltage across the resistor.

$$V_{battery} = V_{total} = V_1 \quad (23.5)$$

$$V_{battery} = V_{total} = V_1 \quad (23.6)$$

$$V_1 = 2\text{V} \quad (23.7)$$

Exercise 23.4: Voltages II What is the voltage across the unknown resistor in the circuit shown?



Solution to Exercise

Step 1. We have a circuit with a battery and two resistors. We know the voltage across the battery and one of the resistors. We want to find that voltage across the resistor.

$$V_{battery} = 2\text{V} \tag{23.8}$$

$$V_{\text{A}} = 1\text{V} \tag{23.9}$$

Step 2. We know that the voltage across the battery must be equal to the total voltage across all other circuit components that are in series.

$$V_{battery} = V_{total} \tag{23.10}$$

The total voltage in the circuit is the sum of the voltages across the individual resistors

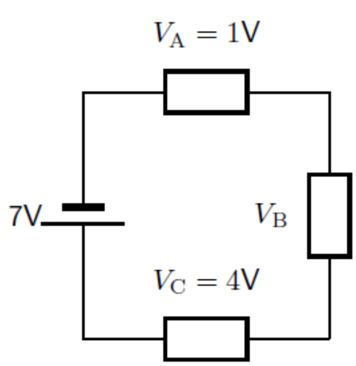
$$V_{total} = V_{\text{A}} + V_{\text{B}} \tag{23.11}$$

Using the relationship between the voltage across the battery and total voltage across the resistors

$$V_{battery} = V_{total} \tag{23.12}$$

$$\begin{aligned} V_{battery} &= V_1 + V_{resistor} \\ 2\text{V} &= V_1 + 1\text{V} \\ V_1 &= 1\text{V} \end{aligned} \tag{23.13}$$

Exercise 23.5: Voltages III What is the voltage across the unknown resistor in the circuit shown?



Solution to Exercise

Step 1. We have a circuit with a battery and three resistors. We know the voltage across the battery and two of the resistors. We want to find that voltage across the unknown resistor.

$$V_{battery} = 7\text{V} \tag{23.14}$$

$$\begin{aligned} V_{known} &= V_{\text{A}} + V_{\text{C}} \\ &= 1\text{V} + 4\text{V} \end{aligned} \tag{23.15}$$

Step 2. We know that the voltage across the battery must be equal to the total voltage across all other circuit components that are in series.

$$V_{battery} = V_{total} \tag{23.16}$$

The total voltage in the circuit is the sum of the voltages across the individual resistors

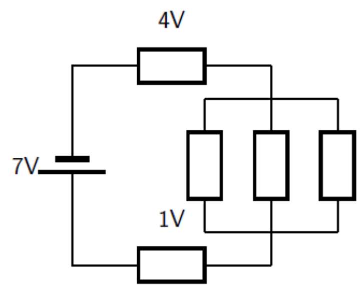
$$V_{total} = V_{\text{B}} + V_{known} \tag{23.17}$$

Using the relationship between the voltage across the battery and total voltage across the resistors

$$V_{battery} = V_{total} \tag{23.18}$$

$$\begin{aligned} V_{battery} &= V_{\text{B}} + V_{known} \\ 7\text{V} &= V_{\text{B}} + 5\text{V} \\ V_{\text{B}} &= 2\text{V} \end{aligned} \tag{23.19}$$

Exercise 23.6: Voltages IV What is the voltage across the parallel resistor combination in the circuit shown? Hint: the rest of the circuit is the same as the previous problem.



Solution to Exercise

Step 1. The circuit is the same as the previous example and we know that the voltage difference between two points in a circuit does not depend on what is between them so the answer is the same as above $V_{parallel} = 2\text{V}$.

Step 2. We have a circuit with a battery and five resistors (two in series and three in parallel). We know the voltage across the battery and two of the resistors. We want to find that voltage across the parallel resistors, $V_{parallel}$.

$$V_{battery} = 7\text{V} \tag{23.20}$$

$$V_{known} = 1\text{V} + 4\text{V} \tag{23.21}$$

Step 3. We know that the voltage across the battery must be equal to the total voltage across all other circuit components.

$$V_{battery} = V_{total} \tag{23.22}$$

Voltages only add for components in series. The resistors in parallel can be thought of as a single component which is in series with the other components and then the voltages can be added.

$$V_{total} = V_{parallel} + V_{known} \tag{23.23}$$

Using the relationship between the voltage across the battery and total voltage across the resistors

$$V_{battery} = V_{total} \tag{23.24}$$

$$\begin{array}{rcl} V_{battery} & = & V_{parallel} + V_{known} \\ 7\text{V} & = & V_1 + 5\text{V} \\ V_{parallel} & = & 2\text{V} \end{array} \tag{23.25}$$

1.176 Current

(section shortcode: P10091)

1.176.1 Flow of Charge

We have been talking about moving charge. But how much charge is moving, and how fast is it moving? The concept that represents this information is called *current*. Current allows us to quantify the movement of charge.

When we talk about current we talk about how much charge moves past a fixed point in circuit in one second. Think of charges being pushed around the circuit by the battery; there are charges in the wires but unless there is a battery they won't move. When one charge moves, the charges next to it also move. They keep their spacing as if you had a tube of marbles like in this picture.

If you push one marble into the tube then one must come out the other side. If you look at any point in the tube and push one marble into the tube, one marble will move past the point you are looking at. This is similar to charges in the wires of a circuit.

If one charge moves then they all move and the same number move at every point in the circuit. This is due to the conservation of charge.

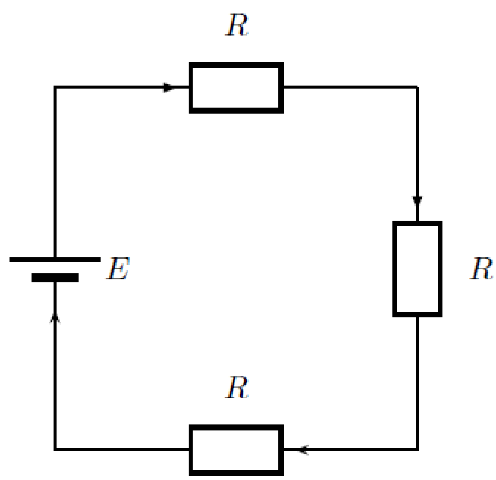
1.176.2 Current

Now that we've thought about moving charges and visualised what is happening we need to get back to quantifying moving charge. We've already said that we call moving charge current but we define it precisely as follows:

Definition: Current
Current is the rate of flow of charge, i.e. the rate at which charges move past a fixed point in a circuit. We use the symbol I to show current and it is measured in amperes (A). One ampere is one coulomb of charge moving in one second. The relationship between current, charge and time is given by:

$$I = \frac{Q}{\Delta t} \tag{23.26}$$

When current flows in a circuit we show this on a diagram by adding arrows. The arrows show the direction of flow in the circuit. By convention we say that charge flows from the positive end (*or* terminal) of a battery, through the circuit, and back to the negative end (*or* terminal) of the battery. This is shown in the diagram below. We measure the current with an instrument called an *ammeter.*



The arrows in this picture show you the direction that charge will flow in the circuit. Note that the arrows point from the positive end of the battery, through the circuit, towards the negative end of the battery.

NOTE: Benjamin Franklin made a guess about the direction of charge flow when rubbing smooth wax with rough wool. He thought that the charges flowed from the wax to the wool (i.e. from positive to negative) which was opposite to the real direction. Due to this, electrons are said to have a *negative* charge and so objects which Ben Franklin called "negative" (meaning a shortage of charge) really have an excess of electrons. By the time the true direction of electron flow was discovered, the convention of ´positive¡ and ´negative¡ had already been so well accepted in the scientific world that no effort was made to change it.

TIP: A battery **does not** produce the same amount of current no matter what is connected to it. While the voltage produced by a battery is constant, the amount of current supplied depends on what is in the circuit.

Exercises: Current

Exercise 23.7 Using the relationship between current, charge, and time, calculate the current in a circuit which has $0,8\,\mathrm{C}$ of charge passing a point every second.

Solution to Exercise

Step 1. $I = \frac{Q}{\Delta t}$

Step 2. Given: $Q = 0,8\,\mathrm{C}$

$$\Delta t = 1\,\mathrm{s} \qquad (23.27)$$

Asked for: I

Step 3.

$$I = \frac{0,8\ \mathrm{C}}{1\ \mathrm{s}} \\ = 0,8\ \mathrm{A} \quad (23.28)$$

Exercise 23.8 How much charge flows per second in a circuit with a current of $1,5\ \mathrm{A}$?

Solution to Exercise

Step 1. Asked for: Charge, Q
Given: Current $I = 1,5\ \mathrm{A}$

$$\Delta t = 1\ \mathrm{s} \quad (23.29)$$

Step 2. $I = \frac{Q}{\Delta t}$

$$Q = I \times \Delta t$$

Step 3.

$$= 1,5\ \mathrm{A} \times 1\ \mathrm{s}$$
$$= 1,5\ \mathrm{C}$$

Exercise 23.9 If $500 \times 10^3\ \mu\mathrm{C}$ of charge flow past a point in a circuit in 1 second, what is the current in the circuit?

Solution to Exercise

Step 1. Asked for: current, I
Given: charge $Q = 500 \times 10^3\ \mu\mathrm{C}$
$\Delta t = 1\ s$

$$Q = 500 \times 10^3 \mu\mathrm{C}$$

Step 2.

$$= 500 \times 10^3 \times 10^{-6}\mathrm{C}$$
$$= 0,5\mathrm{C}$$

Step 3. $I = \frac{Q}{\Delta t}$

$$I = \frac{Q}{\Delta t}$$

Step 4.

$$= \frac{0,5\mathrm{C}}{1\mathrm{s}}$$
$$= 0,5A$$

Exercise 23.10 I measure the current in a circuit to be $500\mathrm{mA}$. How much charge is flowing per second in the circuit?

Solution to Exercise

Step 1. Asked for: Charge, Q
Given: Current $I = 500\mathrm{mA}$,
$\Delta t = 1\mathrm{s}$

Step 2. $500\mathrm{mA} = 500 \times 10^{-3}\mathrm{A}$

Step 3. $I = \frac{Q}{\Delta t}$

Step 4.
$$\begin{aligned} Q &= I \times \Delta t \\ &= 500 \times 10^{-3}\mathrm{A} \times 1\mathrm{s} \\ &= 0,5\,\mathrm{C} \end{aligned}$$

1.176.3 Measuring voltage and current in circuits

As we have seen in previous sections, an electric circuit is made up of a number of different components such as batteries, resistors and light bulbs. There are devices to measure the properties of these components. These devices are called meters.

For example, one may be interested in measuring the amount of current flowing through a circuit using an *ammeter* or measuring the voltage provided by a battery using a *voltmeter*. In this section we will discuss the practical usage of voltmeters and ammeters.

Voltmeter

A voltmeter is an instrument for measuring the voltage between two points in an electric circuit. In analogy with a water circuit, a voltmeter is like a meter designed to measure pressure difference. Since one is interested in measuring the voltage between two points in a circuit, a voltmeter must be connected in *parallel* with the portion of the circuit on which the measurement is made.

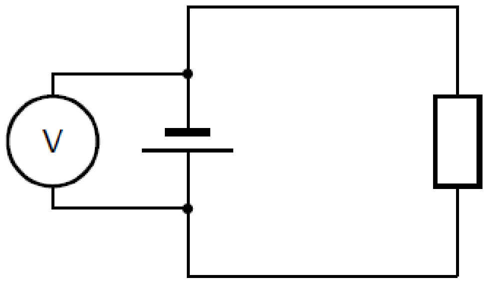


Figure 23.23: A voltmeter should be connected in parallel in a circuit.

Figure 23.23 shows a voltmeter connected in parallel with a battery. The positive lead of the voltmeter must be connected closest to the positive end of the battery and the negative lead closest to the negative end of the battery. The voltmeter may also be used to measure the voltage across a resistor or any other component of a circuit that has a voltage drop or potential difference.

Ammeter

An ammeter is an instrument used to measure the flow of electric current in a circuit. Since one is interested in measuring the current flowing *through* the circuit, the ammeter must be connected in *series* with the measured circuit component (Figure 23.24). The positive lead from the ammeter must be connected closest to the positive end of the battery and the negative lead must be connected closest to the negative end of the battery.

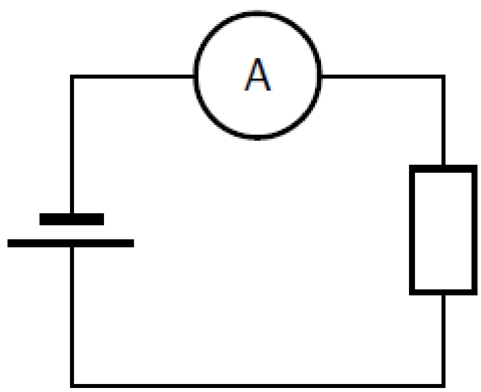


Figure 23.24: An ammeter should be connected in series in a circuit.

Impact of meters on circuits

A good quality meter used correctly will not significantly change the values it is used to measure. This means that an ammeter has very low resistance so as to not slow down the flow of charge. A voltmeter has a very high resistance so that it does not add another parallel pathway to the circuit for the charge to flow along.

Investigation : Using meters

If possible, connect meters in circuits to get used to how to use meters to measure electrical quantities. If the meters have more than one scale, always connect to the **largest scale** first so that the meter will not be damaged by having to measure values that exceed its limits.

The table below summarises the use of each measuring instrument that we discussed and the way it should be connected to a circuit component.

Instrument	Measured Quantity	Proper Connection
Voltmeter	Voltage	In Parallel
Ammeter	Current	In Series

Table 23.2

1.177 Resistance

(section shortcode: P10092)

The resistance of a circuit element can be thought of as how much it opposes the flow of electric current in the circuit.

Definition: Resistance
The resistance of a conductor is defined as the potential difference across it divided by the current flowing though it. We use the symbol **R** to show resistance and it is measured in units called **Ohms** with the symbol Ω.

$$1\ \mathrm{Ohm} = 1\frac{\mathrm{Volt}}{\mathrm{Ampere}}. \quad (23.30)$$

1.177.1 What causes resistance?

We have spoken about resistors that reduce the flow of charge in a conductor. On a microscopic level, electrons moving through the conductor collide with the particles of which the conductor (metal) is made. When they collide, they transfer kinetic energy. The electrons lose kinetic energy and slow down. This leads to resistance. The transferred energy causes the resistor to heat up. You can feel this directly if you touch a cellphone charger when you are charging a cell phone - the charger gets warm because its circuits have some resistors in them!

All conductors have some resistance. For example, a piece of wire has less resistance than a light bulb, but both have resistance. A lightbulb is a very thin wire surrounded by a glass housing The high resistance of the filament (small wire) in a lightbulb causes the electrons to transfer a lot of their kinetic energy in the form of heat[3]. The heat energy is enough to cause the filament to glow white-hot which produces light. The wires connecting the lamp to the cell or battery hardly even get warm while conducting the same amount of current. This is because of their much lower resistance due to their larger cross-section (they are thicker).

An important effect of a resistor is that it *converts* electrical energy into **heat** energy. **Light** is a by-product of the heat that is produced.

Interesting Fact

NOTE: There is a special type of conductor, called a **superconductor** that has no resistance, but the materials that make up all known superconductors only start superconducting at very low temperatures (approximately -170°C).

Why do batteries go flat?

A battery stores chemical potential energy. When it is connected in a circuit, a chemical reaction takes place inside the battery which converts chemical potential energy to electrical energy which powers the charges (electrons) to move through the circuit. All the circuit elements (such as the conducting leads, resistors and lightbulbs) have some resistance to the flow of charge and convert the electrical energy to heat and, in the case of the lightbulb, heat and light. Since energy is always conserved, the battery goes flat when all its chemical potential energy has been converted into other forms of energy.

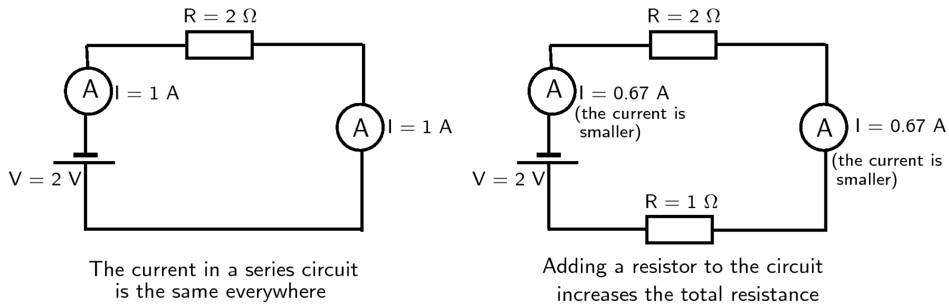
1.177.2 Resistors in electric circuits

It is important to understand what effect adding resistors to a circuit has on the *total* resistance of a circuit and on the current that can flow in the circuit.

Resistors in series

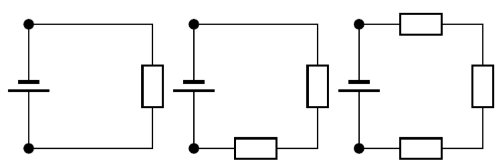
When we add resistors in series to a circuit, we *increase* the resistance to the flow of current. There is only **one path** along which the current can flow and the current is the same at all places in the series circuit. Take a look at the diagram below: On the left there is a circuit with a single resistor and a battery. No matter where we measure the current, it is the same in a series circuit. On the right, we have added a second resistor in series to the circuit. The *total* resistance of the circuit has *increased* and you can see from the reading on the ammeter that the current in the circuit has *decreased* and is still the same everywhere in the circuit.

[3] Flourescent lightbulbs do not use thin wires; they use the fact that certain gases glow when a current flows through them. They are much more efficient (much less resistance) than lightbulbs.

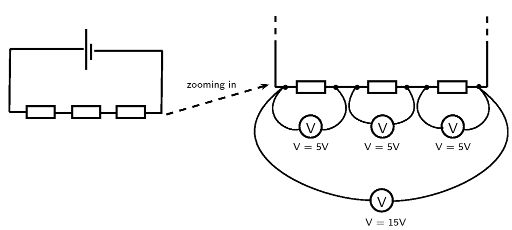


Potential difference and resistors in series When resistors are in series, one after the other, there is a potential difference across each resistor. The total potential difference across a set of resistors in series is the sum of the potential differences across each of the resistors in the set. This is the same as falling a large distance under gravity or falling that same distance (difference) in many smaller steps. The total distance (difference) is the same.

Look at the circuits below. If we measured the potential difference between the black dots in all of these circuits it would be the same; it is just the potential difference across the battery which is the same as the potential difference across the rest of the circuit. So we now know the total potential difference is the same across one, two or three resistors. We also know that some work is required to make charge flow through each one. Each is a step down in potential energy. These steps add up to the total voltage drop which we know is the difference between the two dots. The sum of the potential differences across each individual resistor is equal to the potential difference measured across all of them together. For this reason, series circuits are sometimes called **voltage dividers**.

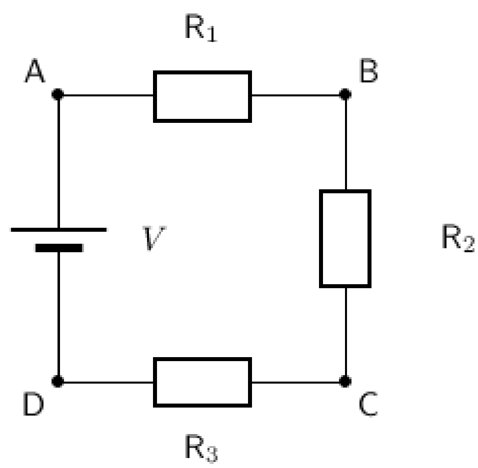


Let us look at this in a bit more detail. In the picture below you can see what the different measurements for 3 identical resistors in series could look like. The total voltage across all three resistors is the sum of the voltages across the individual resistors.

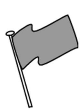


Equivalent Series Resistance

When there is more than one resistor in a circuit, we are usually able to calculate the total combined resitance of all the resistors. The resistance of the single resistor is known as *equivalent resistance* or total resistance. Consider a circuit consisting of three resistors and a single cell connected in series.



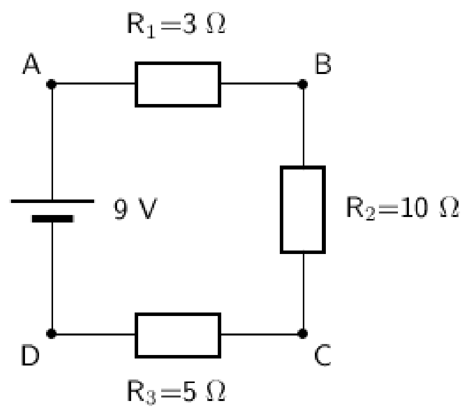
We can define the total resistance in a series circuit as:



Definition: Equivalent resistance in a series circuit, R_s
For n resistors in series the equivalent resistance is:

$$R_s = R_1 + R_2 + R_3 + \cdots + R_n \tag{23.31}$$

The more resistors we add in series, the higher the equivalent resistance in the circuit. Since the resistors act as obstacles to the flow of charge through the circuit, the current in the circuit is reduced. Therefore, the *higher* the resistance in the circuit, the *lower* the current through the battery and the circuit. We say that the current in the battery is inversely proportional to the resistance in the circuit. Let us apply the rule of equivalent resistance in a series circuit to the following circuit.



The resistors are in series, therefore:

$$\begin{aligned} R_s &= R_1 + R_2 + R_3 \\ &= 3\,\Omega + 10\,\Omega + 5\,\Omega \\ &= 18\,\Omega \end{aligned} \tag{23.32}$$

Experiment : Current in Series Circuits

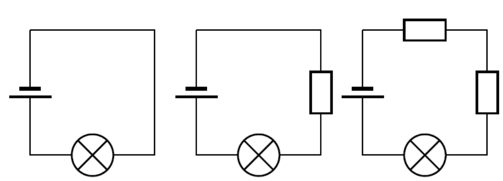
Aim: To determine the effect of multiple resistors on current in a circuit

Apparatus:

- Battery
- Resistors
- Light bulb
- Wires

Method:

1. Construct the following circuits



2. Rank the three circuits in terms of the brightness of the bulb.

Conclusions: The brightness of the bulb is an indicator of how much current is flowing. If the bulb gets brighter because of a change then more current is flowing. If the bulb gets dimmer less current is flowing. You will find that the more resistors you have the dimmer the bulb.

Exercise 23.11: Equivalent series resistance I Two 10 kΩ resistors are connected in series. Calculate the equivalent resistance.

Solution to Exercise

Step 1. Since the resistors are in series we can use:

$$R_s = R_1 + R_2 \tag{23.33}$$

Step 2.

$$\begin{aligned} R_s &= R_1 + R_2 \\ &= 10\,\mathrm{k}\,\Omega + 10\,\mathrm{k}\,\Omega \\ &= 20\,\mathrm{k}\,\Omega \end{aligned} \tag{23.34}$$

Step 3. The equivalent resistance of two 10 kΩ resistors connected in series is 20 kΩ.

Exercise 23.12: Equivalent series resistance II Two resistors are connected in series. The equivalent resistance is 100 Ω. If one resistor is 10 Ω, calculate the value of the second resistor.

Solution to Exercise

Step 1. Since the resistors are in series we can use:

$$R_s = R_1 + R_2 \tag{23.35}$$

We are given the value of R_s and R_1.

Step 2.

$$\begin{aligned} R_s &= R_1 + R_2 \\ \therefore R_2 &= R_s - R_1 \\ &= 100\,\Omega - 10\,\Omega \\ &= 90\,\Omega \end{aligned} \tag{23.36}$$

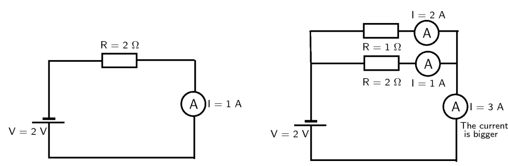
Step 3. The second resistor has a resistance of 90 Ω.

Khan academy video on circuits - 2 (Video: P10093)

Resistors in parallel

In contrast to the series case, when we add resistors in parallel, we create **more paths** along which current can flow. By doing this we *decrease* the total resistance of the circuit!

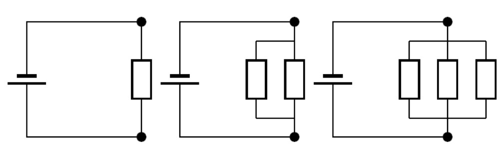
Take a look at the diagram below. On the left we have the same circuit as shown on the left in Figure 23.25 with a battery and a resistor. The ammeter shows a current of 1 ampere. On the right we have added a second resistor in parallel to the first resistor. This has increased the number of paths (branches) the charge can take through the circuit - the total resistance has decreased. You can see that the current in the circuit has increased. Also notice that the current in the different branches can be different (in this case 1 A and 2 A) but must add up to the current through the battery (3 A). Since the total current in the circuit is equal to the sum of the currents in the parallel branches, a parallel circuit is sometimes called a **current divider**.



Adding a resistor to the circuit in parallel decreases the total resistance

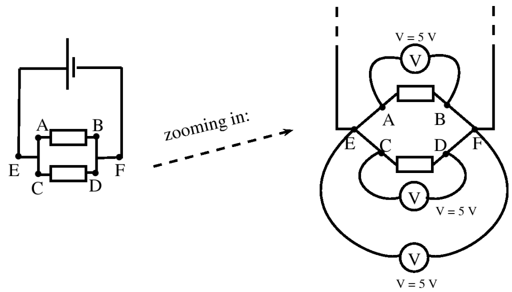
Potential difference and parallel resistors When resistors are connected in parallel the start and end points for all the resistors are the same. These points have the same potential energy and so the potential difference between them is the same no matter what is put in between them. You can have one, two or many resistors between the two points, the potential difference will not change. You can ignore whatever components are between two points in a circuit when calculating the difference between the two points.

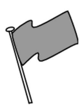
Look at the following circuit diagrams. The battery is the same in all cases. All that changes is that more resistors are added between the points marked by the black dots. If we were to measure the potential difference between the two dots in these circuits we would get the same answer for all three cases.



Let's look at two resistors in parallel more closely. When you construct a circuit you use wires and you might think that measuring the voltage in different places on the wires will make a difference. This is not true. The potential difference or voltage measurement will only be different if you measure a different set of components. All points on the wires that have no circuit components between them will give you the same measurements.

All three of the measurements shown in the picture below (i.e. A–B, C–D and E–F) will give you the same voltage. The different measurement points on the left (i.e. A, E, C) have no components between them so there is no change in potential energy. Exactly the same applies to the different points on the right (i.e. B, F, D). When you measure the potential difference between the points on the left and right you will get the same answer.



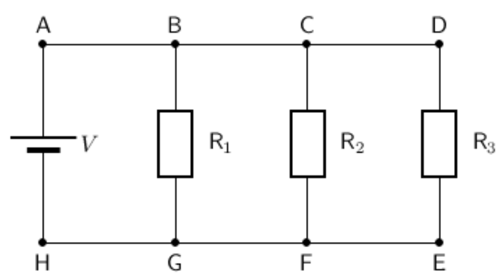


Definition: Equivalent resistance of two parallel resistor, R_p
For 2 resistors in parallel with resistances R_1 and R_2, the equivalent resistance is:

$$R_p = \frac{R_1 R_2}{R_1 + R_2} \tag{23.37}$$

Equivalent parallel resistance

Consider a circuit consisting of a single cell and three resistors that are connected in parallel.

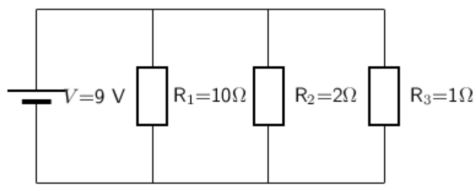


Using what we know about voltage and current in parallel circuits we can define the equivalent resistance of several resistors in parallel as:

Definition: Equivalent resistance in a parallel circuit, R_p
For n resistors in parallel, the equivalent resistance is:

$$\frac{1}{R_p} = \left(\frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} + \cdots + \frac{1}{R_n}\right) \tag{23.38}$$

Let us apply this formula to the following circuit.



What is the total resistance in the circuit?

$$\begin{aligned}
\frac{1}{R_p} &= \left(\frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3}\right) \\
&= \left(\frac{1}{10\,\Omega} + \frac{1}{2\,\Omega} + \frac{1}{1\,\Omega}\right) \\
&= \left(\frac{1+5+10}{10}\right) \\
&= \left(\frac{16}{10}\right) \\
\therefore R_p &= 0{,}625\,\Omega
\end{aligned} \tag{23.39}$$

Experiment : Current in Parallel Circuits

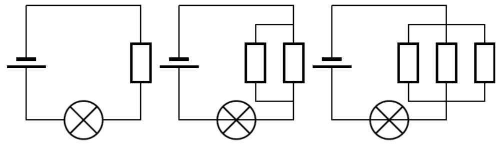
Aim: To determine the effect of multiple resistors on current in a circuit

Apparatus:

- Battery
- Resistors
- Light bulb
- Wires

Method:

1. Construct the following circuits



2. Rank the three circuits in terms of the brightness of the bulb.

Conclusions: The brightness of the bulb is an indicator of how much current is flowing. If the bulb gets brighter because of a change then more current is flowing. If the bulb gets dimmer less current is flowing. You will find that the more resistors you have the brighter the bulb.

Why is this the case? Why do more resistors make it easier for charge to flow in the circuit? It is because they are in parallel so there are more paths for charge to take to move. You can think of it like a highway with more lanes, or the tube of marbles splitting into multiple parallel tubes. The more branches there are, the easier it is for charge to flow. You will learn more about the total resistance of parallel resistors later but always remember that more resistors in parallel mean more pathways. In series the pathways come one after the other so it does not make it easier for charge to flow.

Exercise 23.13 Two $8\,\mathrm{k}\Omega$ resistors are connected in parallel. Calculate the equivalent resistance.

Solution to Exercise

Step 1. Since the resistors are in parallel we can use:

$$\frac{1}{R_p} = \frac{1}{R_1} + \frac{1}{R_2} \tag{23.40}$$

Step 2.

$$\begin{aligned} \frac{1}{R_p} &= \frac{1}{R_1} + \frac{1}{R_2} \\ &= \frac{1}{8\,\mathrm{k}\,\Omega} + \frac{1}{10\,\mathrm{k}\,\Omega} \\ R_p &= \frac{2}{8} \\ &= 4\,\mathrm{k}\,\Omega \end{aligned} \tag{23.41}$$

Step 3. The equivalent resistance of two $8\,\mathrm{k}\Omega$ resistors connected in parallel is $4\,\mathrm{k}\Omega$.

Exercise 23.14 Two resistors are connected in parallel. The equivalent resistance is $100\,\Omega$. If one resistor is $150\,\Omega$, calculate the value of the second resistor.

Solution to Exercise

Step 1. Since the resistors are in parallel we can use:

$$\frac{1}{R_p} = \frac{1}{R_1} + \frac{1}{R_2} \tag{23.42}$$

We are given the value of R_p and R_1.

Step 2.

$$\begin{aligned} \frac{1}{R_p} &= \frac{1}{R_1} + \frac{1}{R_2} \\ \therefore \frac{1}{R_2} &= \frac{1}{R_p} - \frac{1}{R_1} \\ &= \frac{1}{100\,\Omega} - \frac{1}{150\,\Omega} \\ &= \frac{3-2}{300} \\ &= \frac{1}{300} \\ R_2 &= 300\,\Omega \end{aligned} \tag{23.43}$$

Step 3. The second resistor has a resistance of $300\,\Omega$.

Khan academy video on circuits - 3 (Video: P10094)

Resistance

1. What is the unit of resistance called and what is its symbol?
2. Explain what happens to the total resistance of a circuit when resistors are added in series?
3. Explain what happens to the total resistance of a circuit when resistors are added in parallel?
4. Why do batteries go flat?

Find the answers with the shortcodes:

(1.) lqk (2.) lq0 (3.) lq8 (4.) lq9

1.178

(section shortcode: P10095)

Khan academy video on circuits - 4 (Video: P10096)

The following presentation summarizes the concepts covered in this chapter.

(Presentation: P10097)

1.179 Exercises - Electric circuits

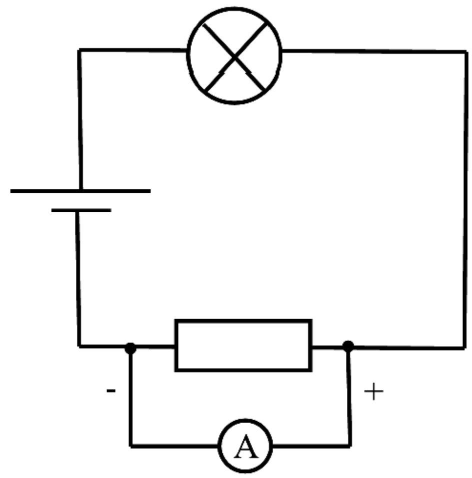
(section shortcode: P10098)

1. Write definitions for each of the following:
   a. resistor
   b. coulomb
   c. voltmeter
2. Draw a circuit diagram which consists of the following components:
   a. 2 batteries in parallel
   b. an open switch
   c. 2 resistors in parallel
   d. an ammeter measuring total current
   e. a voltmeter measuring potential difference across one of the parallel resistors
3. Complete the table below:

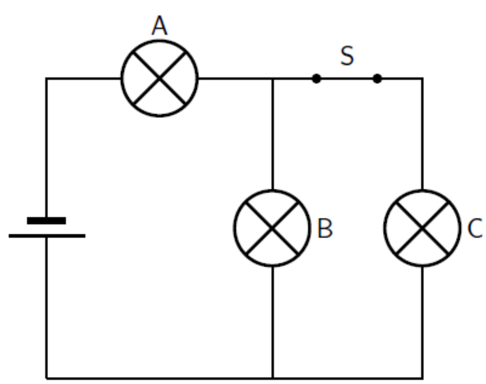
Quantity	**Symbol**	**Unit of meaurement**	**Symbol of unit**
e.g. Distance	e.g. d	e.g. kilometer	e.g. km
Resistance			
Current			
Potential difference			

Table 23.3

4. Draw a diagram of a circuit which contains a battery connected to a lightbulb and a resistor all in series.
   a. Also include in the diagram where you would place an ammeter if you wanted to measure the current through the lightbulb.
   b. Draw where and how you would place a voltmeter in the circuit to measure the potential difference across the resistor.
5. Thandi wants to measure the current through the resistor in the circuit shown below and sets up the circuit as shown below. What is wrong with her circuit setup?



6. (SC 2003/11) The emf of a battery can best be explained as the $\cdots$
   a. rate of energy delivered per unit current
   b. rate at which charge is delivered
   c. rate at which energy is delivered
   d. charge per unit of energy delivered by the battery
7. (IEB 2002/11 HG1) Which of the following is the correct definition of the emf of a battery?
   a. It is the product of current and the external resistance of the circuit.
   b. It is a measure of the cell's ability to conduct an electric current.
   c. It is equal to the "lost volts" in the internal resistance of the circuit.
   d. It is the power supplied by the battery per unit current passing through the battery.
8. (IEB 2005/11 HG) Three identical light bulbs A, B and C are connected in an electric circuit as shown in the diagram below.



   a. How bright is bulb A compared to B and C?
   b. How bright are the bulbs after switch S has been opened?
   c. How do the currents in bulbs A and B change when switch S is opened?

	Current in A	**Current in B**
(a)	decreases	increases
(b)	decreases	decreases
(c)	increases	increases
(d)	increases	decreases

Table 23.4

9. (IEB 2004/11 HG1) When a current I is maintained in a conductor for a time of t, how many electrons with charge e pass any cross-section of the conductor per second?

a. It
b. It/e
c. Ite
d. e/It

Find the answers with the shortcodes:

(1.) IqX (2.) IqI (3.) Iq5 (4.) ITw (5.) ITv (6.) Iqn
(7.) IqR (8.) IqN (9.) IqQ

Safety

.1 Introduction

(section shortcode: S10000)

A laboratory (be it for physics, chemistry or other sciences) can be a very dangerous and daunting place. However, if you follow a few simple guidelines you can safely carry out experiments in the lab without endangering yourself or others around you.

.2 General safety rules

(section shortcode: S10001)

The following are some of the general guidelines and rules that you should always observe when working in a laboratory.

1. Do not eat or drink in the lab. Do not use lab glassware to eat or drink from.
2. Always behave responsibly in the lab. Do not run around or play practical jokes.
3. In case of accidents or chemical spills call your teacher at once.
4. Always check with your teacher how to dispose of waste. Chemicals should not be disposed of down the sink.
5. Only perform the experiments that your teacher instructs you to. Never mix chemicals for fun.
6. Never perform experiments alone.
7. Always check the safety data of any chemicals you are going to use.
8. Follow the given instructions exactly. Do not mix up steps or try things in a different order.
9. Be alert and careful when handling chemicals, hot glassware, etc.
10. Ensure all bunsen burners are turned off at the end of the practical and all chemical containers are sealed.

.3 Hazard signs

(section shortcode: S10002)

The image below lists some of the common hazards signs that you may encounter. You should know what all of these mean.

Symbol	Meaning	Symbol	Meaning
	E Explosive		O Oxidising
	F Highly flammable		T Toxic
	Xn Harmful		Xi Irritating
	C Corrosive		N Harmful for the environment

.4 Notes and information

(section shortcode: S10003)

You can find safety data sheets at Merck[4] . You should always look at these data sheets anytime you work with a new chemical.

You should always try dispose of chemicals correctly and safely. Many chemicals cannot simply be washed down the sink.

[4] http://www.merck-chemicals.co.za/safety-data-sheets/c_O_Sb.s1LQz0AAAEWVOYfVhTo

Glossary

A **Acceleration**

Acceleration is the rate of change of velocity.

Acid rain

Acid rain refers to the deposition of acidic components in rain, snow and dew. Acid rain occurs when sulphur dioxide and nitrogen oxides are emitted into the atmosphere, undergo chemical transformations and are absorbed by water droplets in clouds. The droplets then fall to earth as rain, snow, mist, dry dust, hail, or sleet. This increases the acidity of the soil and affects the chemical balance of lakes and streams.

Amplitude

The **amplitude** is the maximum displacement from equilibrium. For a longitudinal wave which is a pressure wave this would be the maximum increase (or decrease) in pressure from the equilibrium pressure that is cause when a peak (or trough) passes a point.

Amplitude

The *amplitude* is the maximum displacement of a particle from its equilibrium position.

Amplitude

The amplitude of a pulse is a measurement of how far the medium is displaced from rest.

Anti-Node

An anti-node is a point on standing a wave where maximum displacement takes place. A free end of a rope is an anti-node.

Atomic mass number (A)

The number of protons and neutrons in the nucleus of an atom

Atomic number (Z)

The number of protons in an atom

Atomic orbital

An atomic orbital is the region in which an electron may be found around a single atom.

Attraction and Repulsion

Like poles of magnets *repel* each other whilst *unlike* poles *attract* each other.

Average velocity

Average velocity is the total displacement of a body over a time interval.

Avogadro's number

The number of particles in a mole, equal to $6,022 \times 10^{23}$. It is also sometimes referred to as the number of atoms in $12\,\mathrm{g}$ of carbon-12.

B **Boiling point**

The temperature at which a *liquid* changes its phase to become a *gas*. The process is called evaporation and the reverse process is called condensation

C **Chemical bond**

A chemical bond is the physical process that causes atoms and molecules to be attracted to each other, and held together in more stable chemical compounds.

Chemical change

The formation of new substances in a chemical reaction. One type of matter is changed into something different.

Compound

A compound is a group of two or more atoms that are attracted to each other by relatively strong forces or bonds.

Compound

A substance made up of two or more elements that are joined together in a fixed ratio.

Compression

A **compression** is a region in a longitudinal wave where the particles are closest together.

Concentration

Concentration is a measure of the amount of solute that is dissolved in a given volume of liquid. It is measured in $\mathrm{mol \cdot dm^{-3}}$. Another term that is used for concentration is **molarity (M)**

Conductivity

Conductivity is a measure of a solution's ability to conduct an electric current.

Conductors and insulators

A conductor allows the easy movement or flow of something such as heat or electrical charge through it. Insulators are the opposite to conductors because they *inhibit* or reduce the flow of heat, electrical charge, sound etc through them.

Conservation of energy principle

Energy cannot be created or destroyed. It can only be changed from one form to another.

Conservation of Energy

The Law of Conservation of Energy: Energy cannot be created or destroyed, but is merely changed from one form into another.

Conservation of Mechanical Energy

Law of Conservation of Mechanical Energy: The total amount of mechanical energy in a closed system remains constant.

Constructive interference

Constructive interference is when two pulses meet, resulting in a bigger pulse.

Core electrons

All the electrons in an atom, excluding the valence electrons

Covalent bond

Covalent bonding is a form of chemical bonding where pairs of electrons are shared between atoms.

Current

Current is the rate of flow of charge, i.e. the rate at which charges move past a fixed point in a circuit. We use the symbol I to show current and it is measured in amperes (A). One ampere is one coulomb of charge moving in one second. The relationship between current, charge and time is given by:

$$I = \frac{Q}{\Delta t} \tag{23.26}$$

D

Density

Density is a measure of the mass of a substance per unit volume.

Destructive interference

Destructive interference is when two pulses meet, resulting in a smaller pulse.

Displacement

Displacement is the change in an object's position.

Dissociation

Dissociation in chemistry and biochemistry is a general process in which ionic compounds separate or split into smaller molecules or ions, usually in a reversible manner.

E

Electric circuit

An electric circuit is a closed path (with no breaks or gaps) along which electrical charges (electrons) flow powered by an energy source.

Electrolyte

An electrolyte is a substance that contains free ions and behaves as an electrically conductive medium. Because they generally consist of ions in solution, electrolytes are also known as ionic solutions.

Electron configuration

Electron configuration is the arrangement of electrons in an atom, molecule or other physical structure.

Element

An element is a substance that cannot be broken down into other substances through chemical means.

emf

The emf (electromotive force) is the voltage measured across the terminals of a battery when *no current is flowing* through the battery.

Empirical formula

The empirical formula of a chemical compound gives the relative number of each type of atom in that compound.

Empirical formula

This is a way of expressing the *relative* number of each type of atom in a chemical compound. In most cases, the empirical formula does not show the exact number of atoms, but rather the simplest *ratio* of the atoms in the compound.

Equivalent resistance in a parallel circuit, Rp

For n resistors in parallel, the equivalent resistance is:

$$\frac{1}{R_p} = \left(\frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} + \cdots + \frac{1}{R_n}\right) \tag{23.38}$$

Equivalent resistance in a series circuit, Rs

For n resistors in series the equivalent resistance is:

$$R_s = R_1 + R_2 + R_3 + \cdots + R_n \tag{23.31}$$

Equivalent resistance of two parallel resistor, Rp

For 2 resistors in parallel with resistances R_1 and R_2, the equivalent resistance is:

$$R_p = \frac{R_1 R_2}{R_1 + R_2} \tag{23.37}$$

F Frame of Reference

A frame of reference is a reference point combined with a set of directions.

Frequency

The frequency is the number of successive peaks (or troughs) passing a given point in 1 second.

Frequency

The **frequency** of a wave is the number of wavelengths per second.

G Gradient

The gradient of a line can be calculated by dividing the change in the y-value by the change in the x-value.
m = $\frac{\Delta y}{\Delta x}$

H Heterogeneous mixture

A heterogeneous mixture is one that is non-uniform and the different components of the mixture can be seen.

Homogeneous mixture

A homogeneous mixture is one that is uniform, and where the different components of the mixture cannot be seen.

I Instantaneous velocity

Instantaneous velocity is the velocity of a body at a specific instant in time.

Intermolecular force

A force between molecules, which holds them together.

Intramolecular force

The force between the atoms of a molecule, which holds them together.

Ion

An ion is a charged atom. A positively charged ion is called a **cation** e.g. Na^{+}, and a negatively charged ion is called an **anion** e.g. F^{-}. The charge on an ion depends on the number of electrons that have been lost or gained.

Ion exchange reaction

A type of reaction where the positive ions exchange their respective negative ions due to a driving force.

Ionic bond

An ionic bond is a type of chemical bond based on the electrostatic forces between two oppositely-charged ions. When ionic bonds form, a metal donates one or more electrons, due to having a low electronegativity, to form a positive ion or cation. The non-metal atom has a high electronegativity, and therefore readily gains electrons to form a negative ion or anion. The two ions are then attracted to each other by electrostatic forces.

Isotope

The **isotope** of a particular element is made up of atoms which have the same number of protons as the atoms in the original element, but a different number of neutrons.

K Kinetic Energy

Kinetic energy is the energy an object has due to its motion.

L Longitudinal waves

A longitudinal wave is a wave where the particles in the medium move parallel to the direction of propagation of the wave.

M **Magnetism**

Magnetism is one of the phenomena by which materials exert attractive or repulsive forces on other materials.

Medium

A medium is the substance or material in which a wave will move.

Melting point

The temperature at which a *solid* changes its phase or state to become a *liquid*. The process is called melting and the reverse process (change in phase from liquid to solid) is called **freezing**.

Metallic bond

Metallic bonding is the electrostatic attraction between the positively charged atomic nuclei of metal atoms and the delocalised electrons in the metal.

Mixture

A **mixture** is a combination of two or more substances, where these substances are not bonded (or joined) to each other.

Model

A model is a representation of a system in the real world. Models help us to understand systems and their properties. For example, an *atomic model* represents what the structure of an atom *could* look like, based on what we know about how atoms behave. It is not necessarily a true picture of the exact structure of an atom.

Molar mass

Molar mass (M) is the mass of 1 mole of a chemical substance. The unit for molar mass is **grams per mole** or $\mathrm{g \cdot mol^{-1}}$.

Mole

The mole (abbreviation 'n') is the SI (Standard International) unit for 'amount of substance'. It is defined as an amount of substance that contains the same number of particles (atoms, molecules or other particle units) as there are atoms in $12\ \mathrm{g}$ carbon.

Molecular formula

The molecular formula of a chemical compound gives the exact number of atoms of each element in one molecule of that compound.

Molecular formula

This is a concise way of expressing information about the atoms that make up a particular chemical compound. The molecular formula gives the exact number of each type of atom in the molecule.

N **Node**

A node is a point on a standing wave where no displacement takes place at any time. A fixed end of a rope is a node.

P **Parallel circuit**

In a parallel circuit, the charge flowing from the battery can flow along **multiple** paths to return to the battery.

Peaks and troughs

A *peak* is a point on the wave where the displacement of the medium is at a maximum. A point on the wave is a *trough* if the displacement of the medium at that point is at a minimum.

Period (T)

The period (T) is the time taken for two successive peaks (or troughs) to pass a fixed point.

Period

The **period** of a wave is the time taken by the wave to move one wavelength.

Photon

A photon is a quantum (energy packet) of light.

Physical change

A change that can be seen or felt, but that doesn't involve the break up of the particles in the reaction. During a physical change, the *form* of matter may change, but not its *identity*. A change in temperature is an example of a physical change.

Physical Quantity

A physical quantity is anything that you can measure. For example, length, temperature, distance and time are physical quantities.

Planck's constant

Planck's constant is a physical constant named after Max Planck.
$h = 6,626 \times 10^{-34}$ J · s

Position

Position is a measurement of a location, with reference to an origin.

Potential Difference

Electrical potential difference is the difference in electrical potential energy per unit charge between two points. The unit of potential difference is the volt[5] (V). The potential difference of a battery is the voltage measured across it when current is flowing through it.

Potential energy

Potential energy is the energy an object has due to its position or state.

Precipitate

A precipitate is the solid that forms in a solution during a chemical reaction.

Prefix

A prefix is a group of letters that are placed in front of a word. The effect of the prefix is to change meaning of the word. For example, the prefix *un* is often added to a word to mean *not*, as in *un*necessary which means *not necessary*.

Pulse

A pulse is a single disturbance that moves through a medium.

Pulse Speed

Pulse speed is the distance a pulse travels per unit time.

R

Rarefaction

A **rarefaction** is a region in a longitudinal wave where the particles are furthest apart.

Relative atomic mass

Relative atomic mass is the average mass of one atom of all the naturally occurring isotopes of a particular chemical element, expressed in atomic mass units.

Representing circuits

A **physical circuit** is the electric circuit you create with real components.
A **circuit diagram** is a drawing which uses symbols to represent the different components in the physical circuit.

[5] named after the Italian physicist Alessandro Volta (1745–1827)

Resistance

The resistance of a conductor is defined as the potential difference across it divided by the current flowing though it. We use the symbol **R** to show resistance and it is measured in units called **Ohms** with the symbol Ω.

$$1 \text{ Ohm} = 1\frac{\text{Volt}}{\text{Ampere}}. \tag{23.30}$$

Resultant of Vectors

The resultant of a number of vectors is the single vector whose effect is the same as the individual vectors acting together.

S

Scalar

A scalar is a quantity that has only magnitude (size).

Series circuit

In a series circuit, the charge flowing from the battery can only flow along a **single** path to return to the battery.

SI Units

The name *SI units* comes from the French *Système International d'Unités*, which means *international system of units*.

Speed of sound

The speed of sound in air, at sea level, at a temperature of $21°\mathrm{C}$ and under normal atmospheric conditions, is $344\,\mathrm{m\cdot s^{-1}}$.

T

The law of conservation of mass

The mass of a closed system of substances will remain constant, regardless of the processes acting inside the system. Matter can change form, but cannot be created or destroyed. For any chemical process in a closed system, the mass of the reactants must equal the mass of the products.

Transverse Pulse

A pulse where all of the particles disturbed by the pulse move perpendicular (at a right angle) to the direction in which the pulse is moving.

Transverse wave

A *transverse wave* is a wave where the movement of the particles of the medium is perpendicular (at a right angle) to the direction of propagation of the wave.

V

Valence electrons

The electrons in the outer energy level of an atom

Valency

The number of electrons in the outer shell of an atom which are able to be used to form bonds with other atoms.

Vectors

A vector is a quantity that has both magnitude and direction.

Vectors and Scalars

A vector is a physical quantity with magnitude (size) and direction. A scalar is a physical quantity with magnitude (size) only.

Velocity

Velocity is the rate of change of displacement.

Viscosity

Viscosity is a measure of how resistant a liquid is to flowing (in other words, how easy it is to pour the liquid from one container to another).

W

Water hardness

Water hardness is a measure of the mineral content of water. Minerals are substances such as calcite, quartz and mica that occur naturally as a result of geological processes.

Wave

A *wave* is a periodic, continuous disturbance that consists of a *train* of pulses.

Wavelength of wave

The wavelength of a wave is the distance between any two adjacent points that are in phase.

Wavelength

The **wavelength** in a longitudinal wave is the distance between two consecutive points that are in phase.

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