Chapter 10
"Chemical Quantities"

Yes, you <u>will</u> need a calculator for this chapter!

SECTION 10.1
THE MOLE: A MEASUREMENT OF MATTER

- OBJECTIVES:
  - Describe methods of *measuring* the amount of something.
  - Define *Avogadro's number* as it relates to a mole of a substance.
  - Distinguish between the *atomic mass* of an element and its *molar mass.*
  - Describe how the mass of a mole of a compound is *calculated.*

HOW DO WE MEASURE ITEMS?

- You can measure *mass*,
- or *volume*,
- We measure mass in <u>grams</u>.
- We measure volume in <u>liters</u>.
- or you can *count pieces*.
- or you can *count pieces*.
- We count pieces in <u>MOLES</u>.

What is the mole?

We're not talking about this kind of mole!

MOLES (IS ABBREVIATED: MOL)

- It is an amount, defined as the number of carbon atoms in exactly 12 grams of carbon-12.
- 1 mole = 6.022×10^{23} of the representative particles.
- Treat it like a very large dozen
- 6.022×10^{23} is called: Avogadro's number.

SIMILAR WORDS FOR AN AMOUNT

- Pair: 1 pair of shoelaces

  = 2 shoelaces

- Dozen: 1 dozen oranges

  = 12 oranges

- Gross: 1 gross of pencils

  = 144 pencils

- Ream: 1 ream of paper

  = 500 sheets of paper

WHAT ARE REPRESENTATIVE PARTICLES?

- The smallest pieces of a substance:

1) For a molecular compound: it is the <u>molecule</u>.
2) For an ionic compound: it is the <u>formula unit</u> (made of ions).
3) For an element: it is the <u>atom</u>.

TYPES OF QUESTIONS

- How many *oxygen atoms* in the following?

  $CaCO_3$ — 3 atoms of oxygen

  $Al_2(SO_4)_3$ — 12 (3 x 4) atoms of oxygen

- How many *ions* in the following?

  $CaCl_2$ — 3 total ions (1 Ca^{2+} ion and 2 Cl^{1-} ions)

  NaOH — 2 total ions (1 Na^{1+} ion and 1 OH^{1-} ion)

  $Al_2(SO_4)_3$ — 5 total ions (2 Al^{3+} + 3 SO_4^{2-} ions)

PRACTICE PROBLEMS (ROUND APPROPRIATELY.)

- How many molecules of CO_2 are in 4.56 moles of CO_2?
  2.75×10^{24} molecules
- How many moles of water is 5.87×10^{22} molecules?
  0.0975 mol (or 9.75×10^{-2})
- How many atoms of carbon are in 1.23 moles of $C_6H_{12}O_6$?
  4.44×10^{24} atoms C
- How many moles is 7.78×10^{24} formula units of $MgCl_2$?
  12.9 moles

MEASURING MOLES

- Remember relative atomic mass?
  - The amu was one twelfth the mass of a carbon-12 atom.
- Since the mole is the number of atoms in 12 grams of carbon-12,
- the <u>decimal number</u> on the periodic table is also the mass of 1 mole of those atoms in grams.

GRAM ATOMIC MASS (GAM)

- Equals the mass of 1 mole of an element in grams (from periodic table)
- 12.01 grams of C has the same number of pieces as 1.008 grams of H and 55.85 grams of iron.
- We can write this as: 12.01 g C = 1 mole C (this is also the molar mass)
- We can count things by weighing them.

EXAMPLES

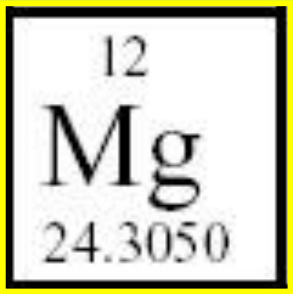
- What is the mass of 2.34 moles of carbon? 28.1 grams C
- How many moles of magnesium is 48.62 g of Mg? 2 mol Mg
- How many atoms of lithium is 1.00 g of Li? 8.68×10^{22} atoms Li
- What is the mass of 3.45×10^{22} atoms of U? 13.6 grams U

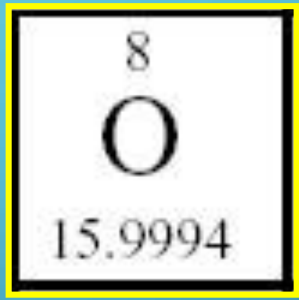
WHAT ABOUT COMPOUNDS?

- in 1 mole of H_2O molecules there are two moles of H atoms and 1 mole of O atoms (think of a compound as a molar ratio)
- To find the mass of one mole of a compound
  - determine the number of moles of the elements present
  - Multiply the number times their mass (from the periodic table)
  - add them up for the total mass

CALCULATING FORMULA MASS

Calculate the formula mass of magnesium carbonate, $MgCO_3$





24.3050 g + 12.0107 g + 3 x (15.9994 g) = 84.3139 g

Thus, 84.3139 grams is the formula mass for **$MgCO_3$**.

TEXT BOOK QUESTIONS

PAGE 291 #3, 4

PAGE 296 #7, 8,13, 14, 15

SECTION 10.2
MOLE-MASS AND MOLE-VOLUME RELATIONSHIPS

- OBJECTIVES:
  - Describe how to *convert* the mass of a substance to the number of moles of a substance, and moles to mass.
  - Identify the *volume* of a quantity of gas at STP.

MOLAR MASS

- Molar mass is the generic term for the mass of one mole of *any* substance (expressed in grams/mol)
- The same as:

1) Gram Molecular Mass (for molecules)
2) Gram Formula Mass (ionic compounds)
3) Gram Atomic Mass (for elements)
    - molar mass is just a much broader term than these other specific masses

EXAMPLES

- Calculate the molar mass of the following and tell what type it is:

Na_2S = 78 g/mol gram formula mass

N_2O_4 = 92 g/mol gram molecular mass

C = 12 g/mol gram atomic mass

$Ca(NO_3)_2$ = 164 g/mol gram formula mass

$C_6H_{12}O_6$ = 180 g/mol gram molecular mass

$(NH_4)_3PO_4$ = 149 g/mol gram formula mass

SINCE MOLAR MASS IS…

- The number of grams in 1 mole of atoms, ions, or molecules,
- We can make *conversion factors* from these.
  - To change between grams of a compound and moles of a compound.

FOR EXAMPLE

- How many moles is 5.69 g of NaOH?

(Solution on next slides)

FOR EXAMPLE

- How many moles is 5.69 g of NaOH?

$$5.69\ \text{g}\left(\frac{}{}\right)$$

FOR EXAMPLE

- How many moles is 5.69 g of NaOH?

$$5.69 \text{ g}\left(\frac{\text{mole}}{g}\right)$$

- **We need to change 5.69 grams NaOH to moles**

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- **1mole Na = 23 g 1 mol O = 16 g 1 mole of H = 1 g**

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- **1mole Na = 23 g   1 mol O = 16 g**
  1 mole of H = 1 g
- **1 mole NaOH = 40 g**

FOR EXAMPLE

- How many moles is 5.69 g of NaOH?

$$5.69\ \text{g}\left(\frac{1\ \ \text{mole}}{40.00\ \ g}\right)$$

- **We need to change 5.69 grams NaOH to moles**
- **1mole Na = 23 g   1 mol O = 16 g**
  1 mole of H = 1 g
- **1 mole NaOH = 40 g**

FOR EXAMPLE

- How many moles is 5.69 g of NaOH?

$$5.69 \text{ g}\left(\frac{1 \text{ mole}}{40.00 \ g}\right) = 0.142 \text{ mol NaOH}$$

- **We need to change 5.69 grams NaOH to moles**
- **1mole Na = 23 g   1 mol O = 16 g**
  1 mole of H = 1 g
- **1 mole NaOH = 40 g**

TEXT BOOK QUESTIONS

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THE MOLE-VOLUME RELATIONSHIP

- Many of the chemicals we deal with are in the physical state as: gases.

  - They are difficult to *weigh* (*or mass*).

- But, we may still need to know how many moles of gas we have.
- Two things effect the volume of a gas:

  a) Temperature and b) Pressure

- We need to compare all gases at the same temperature and pressure.

STANDARD TEMPERATURE AND PRESSURE

- 0°C and 1 atm pressure

  - is abbreviated "STP"

- At STP, 1 mole of *any gas* occupies a volume of 22.4 L

  - Called the molar volume

- This is our fourth equality:
  - 1 mole of *any gas at STP* = 22.4 L

PRACTICE EXAMPLES

- What is the volume of 4.59 mole of CO_2 gas at STP? = 103 L
- How many moles is 5.67 L of O_2 at STP? = 0.253 mol
- What is the volume of 8.8 g of CH_4 gas at STP? = 12.3 L

DENSITY OF A GAS

- $D = m / V$ (density = mass/volume)

  - for a gas the units will be: g / L

- We can determine the density of any gas at STP if we know its formula.
- To find the density we need: 1) mass and 2) volume.
- If you assume you have 1 mole, then the mass is the molar mass (from periodic table)
- And, at STP the volume is 22.4 L.

PRACTICE EXAMPLES (D=M/V)

- Find the density of CO_2 at STP.

  $D = 44g/22.4L = 1.96\ g/L$

- Find the density of CH_4 at STP.

  $D = 16g/22.4L = 0.716\ g/L$

ANOTHER WAY:

- *If given the density*, we can find the molar mass of the gas.
- Again, pretend you have 1 mole at STP, so V = 22.4 L.

modify: D = m/V to show:

$$\boxed{m = D \times V}$$

- "m" will be the mass of 1 mole, since you have 22.4 L of the stuff.
- What is the molar mass of a gas with a density of 1.964 g/L?

= 44.0 g/mol

- How about a density of 2.86 g/L?

= 64.0 g/mol

SUMMARY

- These four items are all equal:

a) 1 mole

b) molar mass (in grams/mol)

c) 6.022×10^{23} representative particles

(atoms, molecules, or formula units)

d) 22.4 L of gas at STP

Thus, we can make conversion factors from these 4 values!

TEXT BOOK QUESTIONS

PAGE 301 #20, 21

PAGE 302 #22, 23

PAGE 303 #25 – 31

SECTION 10.3
PERCENT COMPOSITION AND CHEMICAL FORMULAS

- OBJECTIVES:
  - Describe how to *calculate the percent* by mass of an element in a compound.
  - Interpret an *empirical formula.*
  - Distinguish between *empirical* and *molecular formulas.*

CALCULATING PERCENT COMPOSITION OF A COMPOUND

- Like all percent problems:

$$\left[\frac{\text{part}}{\text{whole}}\ \mathbf{x\ 100\ \%}\right] \mathbf{= percent}$$

1) Find the mass of each of the components (the elements),
2) Next, divide by the total mass of the compound; then x 100

EXAMPLE

- Calculate the percent composition of a compound that is made of 29.0 grams of Ag with 4.30 grams of S.

$$\frac{29.0 \text{ g Ag}}{33.3 \text{ g total}} \times 100 = 87.1\ \%\ \text{Ag}$$

$$\frac{4.30 \text{ g S}}{33.3 \text{ g total}} \times 100 = 12.9\ \%\ \text{S}$$

Total = 100 %

GETTING IT FROM THE FORMULA

- If we know the formula, assume you have 1 mole,
- then you know the mass of the elements and the whole compound (these values come from the periodic table!).

EXAMPLES

- Calculate the percent composition of C_2H_4?

85.7% C, 14.3 % H

- How about Aluminum carbonate?

23.1% Al, 15.4% C, and 61.5 % O

- Sample Problem 10.10, p.307
- We can also use the percent as a *conversion factor*
- Sample Problem page 308

TEXT BOOK QUESTIONS

PAGE 306 – 307 #32 – 35

FORMULAS

***Empirical formula*: the lowest whole number ratio of atoms in a compound.**

***Molecular formula*: the true number of atoms of each element in the formula of a compound.**

- Example: molecular formula for benzene is C_6H_6 (note that everything is divisible by 6)
  - Therefore, the empirical formula = CH (the lowest *whole* number ratio)

FORMULAS (CONTINUED)

Formulas for *ionic compounds* are ALWAYS empirical (the lowest whole number ratio = cannot be reduced).

Examples:

$NaCl$ $MgCl_2$ $Al_2(SO_4)_3$ K_2CO_3

FORMULAS (CONTINUED)

Formulas for *molecular compounds* MIGHT be empirical (lowest whole number ratio).

Molecular: (Correct formula) H_2O $C_6H_{12}O_6$ $C_{12}H_{22}O_{11}$

$\downarrow$ $\downarrow$ $\downarrow$

Empirical: (Lowest whole number ratio) H_2O CH_2O $C_{12}H_{22}O_{11}$

CALCULATING EMPIRICAL

- Just find the lowest whole number ratio

  $C_6H_{12}O_6$ = CH_2O

  CH_4N = this is already the lowest ratio.

- A formula is not just the ratio of *atoms*, it is also the ratio of *moles*.

- In 1 mole of CO_2 there is 1 mole of carbon and 2 moles of oxygen.

- In one molecule of CO_2 there is 1 atom of C and 2 atoms of O.

CALCULATING EMPIRICAL

- We can get a ratio *from the percent composition.*

1) Assume you have a 100 g sample

   - the percentage become grams (75.1% = 75.1 grams)

2) Convert grams to moles.

3) Find lowest whole number ratio by dividing each number of moles by the smallest value.

EXAMPLE

- Calculate the empirical formula of a compound composed of 38.67 % C, 16.22 % H, and 45.11 %N.

- Assume 100 g sample, so
- $\frac{38.67 \text{ g C}}{12.0 \text{ g C}} \times 1 \text{mol C} = 3.22$ mole C
- $\frac{16.22 \text{ g H}}{1.0 \text{ g H}} \times 1 \text{mol H} = 16.22$ mole H
- $\frac{45.11 \text{ g N}}{14.0 \text{ g N}} \times 1 \text{mol N} = 3.22$ mole N

Now divide each value by the smallest value

EXAMPLE

- The ratio is $\frac{3.22 \text{ mol C}}{3.22 \text{ mol N}} = \frac{1 \text{ mol C}}{1 \text{ mol N}}$
- The ratio is $\frac{16.22 \text{ mol H}}{3.22 \text{ mol N}} = \frac{5 \text{ mol H}}{1 \text{ mol N}}$

$= C_1H_5N_1$ which is $= CH_5N$

- A compound is 43.64 % P and 56.36 % O. What is the empirical formula?

  $= P_2O_5$

- Caffeine is 49.48% C, 5.15% H, 28.87% N and 16.49% O. What is its empirical formula?

  $= C_4H_5N_2O$

EMPIRICAL TO MOLECULAR

- Since the empirical formula is the *lowest ratio*, the actual molecule would weigh more.
  - By a whole number multiple.
- Divide the actual molar mass by the empirical formula mass – you get a whole number to *increase each coefficient* in the empirical formula
- Caffeine has a molar mass of 194 g. what is its molecular formula? $= C_8H_{10}N_4O_2$

TEXT BOOK QUESTIONS

PAGE 310 #36, 37

PAGE 312 #38, 39, 42 – 46

End of Chapter 10