

## 4

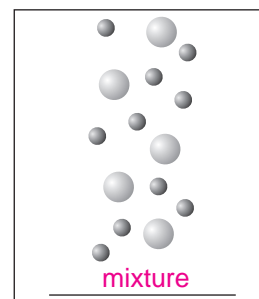
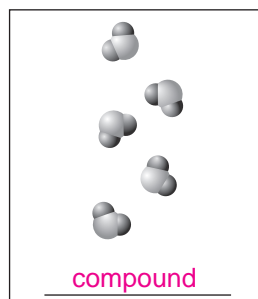
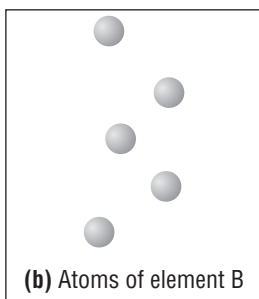
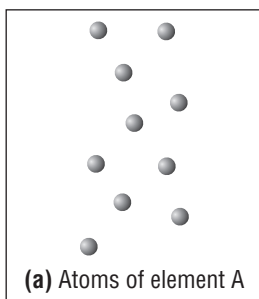
## ATOMIC STRUCTURE

**SECTION 4.1 DEFINING THE ATOM (pages 101–103)**

*This section describes early atomic theories of matter and provides ways to understand the tiny size of individual atoms.*

**► Early Models of the Atom (pages 101–102)**

- Democritus, who lived in Greece during the fourth century B.C., suggested that matter is made up of tiny particles that cannot be divided. He called these particles atoms.
- List two reasons why the ideas of Democritus were not useful in a scientific sense. They did not explain chemical behavior, and they lacked experimental support because Democritus's approach was not based on the scientific method.
- The modern process of discovery about atoms began with the theories of an English schoolteacher named John Dalton.
- Circle the letter of each sentence that is true about Dalton's atomic theory.
  - All elements are composed of tiny, indivisible particles called atoms.
  - An element is composed of several types of atoms.
  - Atoms of different elements can physically mix together, or can chemically combine in simple, whole-number ratios to form compounds.
  - Chemical reactions occur when atoms are separated, joined, or rearranged; however, atoms of one element are never changed into atoms of another element by a chemical reaction.
- In the diagram, use the labels *mixture* and *compound* to identify the mixture of elements A and B and the compound that forms when the atoms of elements A and B combine chemically.



## CHAPTER 4, Atomic Structure (continued)

### ► Sizing up the Atom (page 103)

- Suppose you could grind a sample of the element copper into smaller and smaller particles. The smallest particle that could no longer be divided, yet still has the chemical properties of copper, is an atom of copper.
- About how many atoms of copper when placed side by side would form a line 1 cm long? 100,000,000

## SECTION 4.2 STRUCTURE OF THE NUCLEAR ATOM (pages 104–108)

*This section describes the experiments that led to the discovery of subatomic particles and their properties.*

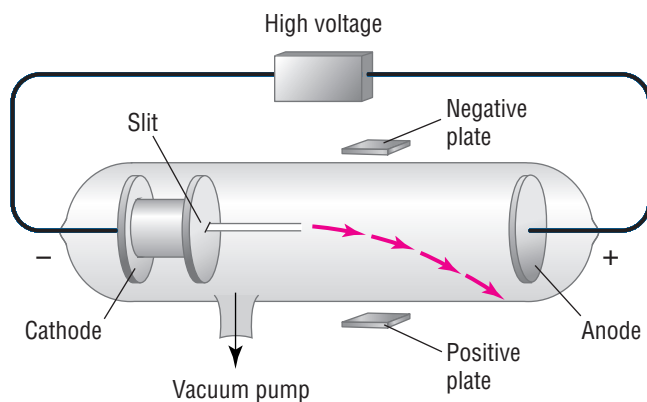
### ► Subatomic Particles (pages 104–106)

- How is the atomic theory that is accepted today different from Dalton's atomic theory? Atoms are now known to be divisible. They can be broken down into even smaller, more fundamental particles.
- Which subatomic particles carry a negative charge? electrons

Match each term from the experiments of J. J. Thomson with the correct description.

- |          |                |  |
|----------|----------------|--|
| <u>c</u> | 3. anode       | a. an electrode with a negative charge                 |
| <u>a</u> | 4. cathode     | b. a glowing beam traveling between charged electrodes |
| <u>b</u> | 5. cathode ray | c. an electrode with a positive charge                 |
| <u>d</u> | 6. electron    | d. a negatively charged particle                       |

- The diagram shows electrons moving from left to right in a cathode-ray tube. Draw an arrow showing how the path of the electrons will be affected by the placement of the negatively and positively charged plates.



8. Thomson observed that the production of cathode rays did not depend on the kind of gas in the tube or the type of metal used for the electrodes. What conclusion did he draw from these observations?

He concluded that electrons must be parts of the atoms of all elements.

9. What two properties of an electron did Robert Millikan determine from his experiments?

He determined the quantity of charge carried by an electron and its mass.

10. Circle the letter of each sentence that is true about atoms, matter, and electric charge.

a. All atoms have an electric charge.

b. Electric charges are carried by particles of matter.

c. Electric charges always exist in whole-number multiples of a single basic unit.

d. When a given number of positively charged particles combines with an equal number of negatively charged particles, an electrically neutral particle is formed.

11. Circle the letter next to the number of units of positive charge that remain if a hydrogen atom loses an electron.

a. 0

b. 1

c. 2

d. 3

12. The positively charged subatomic particle that remains when a hydrogen atom loses an electron is called a proton.

13. What charge does a neutron carry? none.

14. Complete the table about the properties of subatomic particles.

Properties of Subatomic Particles				
Particle	Symbol	Relative Electrical Charge	Relative Mass (mass of proton = 1)	Actual Mass (g)
Electron	e <sup>-</sup>	1-	$\frac{1}{1840}$	$9.11 \times 10^{-28}$
Proton	p <sup>+</sup>	1+	1	$1.67 \times 10^{-24}$
Neutron	n <sup>0</sup>	0	1	$1.67 \times 10^{-24}$

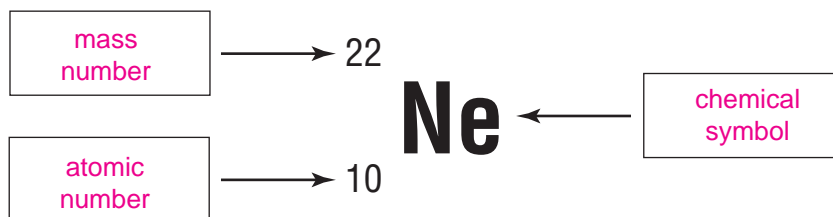


2. Complete the table showing the number of protons and electrons in atoms of six elements.

Atoms of Six Elements				
Name	Symbol	Atomic Number	Number of Protons	Number of Electrons
Hydrogen	H	1	1	1
Helium	He	2	2	2
Lithium	Li	3	3	3
Boron	B	5	5	5
Carbon	C	6	6	6
Oxygen	O	8	8	8

► **Mass Number** (pages 111–112)

3. The total number of protons and neutrons in an atom is its mass number.
4. What is the mass number of a helium atom that has two protons and two neutrons? 4
5. How many neutrons does a beryllium atom with four protons and a mass number of nine have? 5
6. Place the labels *chemical symbol*, *atomic number*, and *mass number* in the shorthand notation below.



7. Designate the atom shown in Question 6 in the form “name of element” - “mass number.” neon-22
8. How many protons, neutrons, and electrons are in the atom discussed in Questions 6 and 7? Protons: 10 Neutrons: 12 Electrons: 10

**CHAPTER 4, Atomic Structure** (continued)**► Isotopes (pages 112–113)**

9. How do atoms of neon-20 and neon-22 differ?

An atom of neon-22 has two more neutrons in its nucleus than an atom of neon-20.

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10. Neon-20 and neon-22 are called isotopes.

11. Is the following sentence true or false? Isotopes are chemically alike because they have identical numbers of protons and electrons. true

Match the designation of each hydrogen isotope with its commonly used name.

b 12. hydrogen-1      a. tritium

c 13. hydrogen-2      b. hydrogen

a 14. hydrogen-3      c. deuterium

**► Atomic Mass (pages 114–117)**

15. Why is the atomic mass unit (amu), rather than the gram, usually used to express atomic mass?

The values of atomic masses measured in grams are inconveniently small and impractical to work with.

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16. What isotope of carbon has been chosen as the reference isotope for atomic mass units? What is the defined atomic mass in amu of this isotope?

carbon-12; 12 amu

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17. Is the following sentence true or false? The atomic mass of an element is always a whole number of atomic mass units. false

18. Circle the letter of each statement that is true about the average atomic mass of an element and the relative abundance of its isotopes.

a. In nature, most elements occur as a mixture of two or more isotopes.

b. Isotopes of an element do not have a specific natural percent abundance.

c. The average atomic mass of an element is usually closest to that of the isotope with the highest natural abundance.

d. Because hydrogen has three isotopes with atomic masses of about 1 amu, 2 amu, and 3 amu, respectively, the average atomic mass of natural hydrogen is 2 amu.

19. Circle the letter of the correct answer. When chlorine occurs in nature, there are three atoms of chlorine-35 for every one atom of chlorine-37. Which atomic mass number is closer to the average atomic mass of chlorine?

a. 35 amu

b. 37 amu

20. In the periodic table, the elements are organized into groups based on a set of repeating properties.

► **The Periodic Table—A Preview (page 118)**

21. What are the horizontal rows in the periodic table called?  
periods



## Reading Skill Practice

Outlining can help you understand and remember what you have read. Prepare an outline of Section 4.3, *Distinguishing Among Atoms*. Begin with the headings in the textbook. Under each heading, write the main idea. Then list the details that support the main idea. Do your work on a separate sheet of paper.

Main heads should be Atomic Number, Mass Number, Isotopes, and Atomic Mass. Subheads should indicate an understanding of each concept.

## CHAPTER 4, Atomic Structure (continued)

# GUIDED PRACTICE PROBLEMS

Fill in the write-on lines and boxes provided as you work through the guided practice problems.

### GUIDED PRACTICE PROBLEM 18 (page 112)

18. Use Table 4.2 to express the compositions of carbon-12, fluorine-19, and beryllium-9 in shorthand notation.

#### Analyze

##### Carbon-12

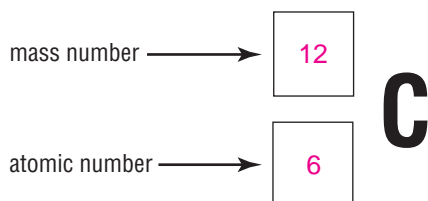
**Step 1.** The number of protons in an atom is called its atomic number.

The number of protons in an atom of carbon-12 is 6.

#### Calculate

**Step 2.** The number of protons plus the number of neutrons in an atom is called its mass number. For carbon-12, this number is 12.

**Step 3.** The shorthand notation for carbon-12 is:



#### Evaluate

**Step 4.** Except for hydrogen-1, the mass number of an isotope is always greater than its atomic number. Is the mass number reasonable? yes

##### Fluorine-19

**Step 1.** The atomic number of fluorine-19 is 9.

**Step 2.** Its mass number is 19.

**Step 3.** The shorthand notation for fluorine-19 is:



**Step 4.** Is your answer reasonable? Why?

Yes it is, because the mass number is greater than the atomic number.



**Beryllium-9**

**Step 1.** The atomic number of beryllium-9 is 4 .

**Step 2.** Its mass number is 9 .

**Step 3.** The shorthand notation for beryllium-9 is:



**Step 4.** Is your answer reasonable? Why?

Yes it is, because the mass number is greater than the atomic number.

**EXTRA PRACTICE (similar to Practice Problem 19, page 113)**

19. Three isotopes of sulfur are sulfur-32, sulfur-33, and sulfur-34. Write the complete symbol for each isotope, including the atomic number and the mass number.

**GUIDED PRACTICE PROBLEM 23 (page 117)**

23. The element copper has naturally occurring isotopes with mass numbers of 63 and 65. The relative abundance and atomic masses are 69.2% for mass = 62.93 amu and 30.8% for mass = 64.93 amu. Calculate the average atomic mass of copper.

**Analyze**

**Step 1.** Will the average atomic mass be closer to 63 or to 65? Explain.

Because about 70% of the element consists of the lighter isotope, the average atomic mass should be closer to the atomic mass number of 63 than 65.

**Solve**

**Step 2.** For Cu-63:  $69.2\% \times 62.93 \text{ amu} = 0.692 \times 62.93 \text{ amu} =$  43.6 amu

**Step 3.** For Cu-65:  $30.8\% \times 64.93 \text{ amu} =$  0.308  $\times$  64.93 amu  $=$  20.0 amu

**Step 4.** Average mass:  $43.6 \text{ amu} +$  20.0 amu  $=$  63.6 amu

**Evaluate**

**Step 5.** Explain why your answer is reasonable.

It is closer to 63 than 65.

**CHAPTER 4, Atomic Structure** (*continued*)

**EXTRA PRACTICE (similar to Practice Problem 24, page 117)**

24. Calculate the atomic mass of rubidium. The two isotopes of rubidium have atomic masses and relative abundances of 84.91 amu (72.16%) and 86.91 amu (27.84%). 85.47 amu