# WHAT'S AHEAD

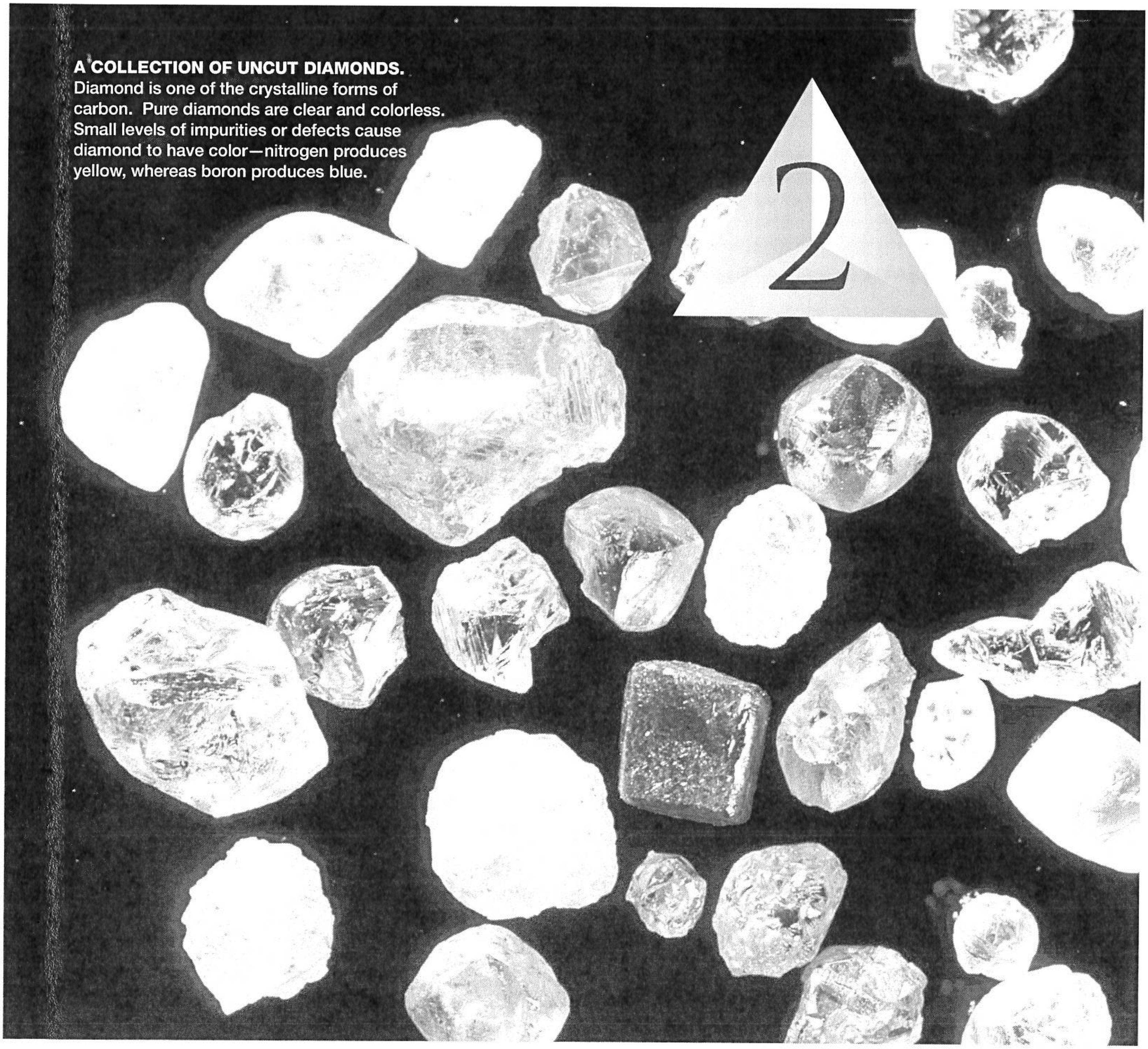
**2.1** THE ATOMIC THEORY OF MATTER
We begin with a brief history of the notion of *atoms*—the smallest pieces of matter.

**2.2** THE DISCOVERY OF ATOMIC STRUCTURE
We then look at some key experiments that led to the discovery of *electrons* and to the *nuclear model* of the atom.

**2.3** THE MODERN VIEW OF ATOMIC STRUCTURE
We explore the modern theory of atomic structure, including the ideas of *atomic numbers, mass numbers,* and *isotopes.*

**2.4** ATOMIC WEIGHTS
We introduce the concept of *atomic weights* and how they relate to the masses of individual atoms.

**2.5** THE PERIODIC TABLE
We examine the organization of the *periodic table,* in which elements are put in order of increasing atomic number and grouped by chemical similarity.




**A COLLECTION OF UNCUT DIAMONDS.** Diamond is one of the crystalline forms of carbon. Pure diamonds are clear and colorless. Small levels of impurities or defects cause diamond to have color—nitrogen produces yellow, whereas boron produces blue.

2.6 MOLECULES AND MOLECULAR COMPOUNDS
We discuss the assemblies of atoms called *molecules* and how their compositions are represented by *empirical* and *molecular formulas*.

2.7 IONS AND IONIC COMPOUNDS
We learn that atoms can gain or lose electrons to form *ions.* We also look at how to use the periodic table to predict the charges on ions and the empirical formulas of *ionic compounds.*

2.8 NAMING INORGANIC COMPOUNDS
We consider the systematic way in which substances are named, called *nomenclature,* and how this nomenclature is applied to inorganic compounds.

2.9 SOME SIMPLE ORGANIC COMPOUNDS
We introduce *organic chemistry,* the chemistry of the element carbon.

# ATOMS, MOLECULES, AND IONS

LOOK AROUND AT THE GREAT variety of colors, textures, and other properties in the materials that surround you—the colors in a garden, the texture of the fabric in your clothes, the solubility of sugar in a cup of coffee, or the transparency and beauty of a diamond. The materials in our world exhibit a striking and seemingly infinite variety of properties, but how do we understand and explain them? What makes diamonds transparent and hard, whereas table salt is brittle and dissolves in water? Why does paper burn, and why does water quench fires? The structure and behavior of atoms are key to understanding both the physical and chemical properties of matter.

Although the materials in our world vary greatly in their properties, everything is formed from only about 100 elements and, therefore, from only about 100 chemically different kinds of atoms. In a sense, the atoms are like the 26 letters of the English alphabet that join in different combinations to form the immense number of words in our language. But what rules govern the ways in which atoms combine? How do the properties of a substance relate to the kinds of atoms it contains? Indeed, what is an atom like, and what makes the atoms of one element different from those of another?

In this chapter we examine the basic structure of atoms and discuss the formation of molecules and ions, thereby providing a foundation for exploring chemistry more deeply in later chapters.

## 2.1 THE ATOMIC THEORY OF MATTER

Philosophers from the earliest times speculated about the nature of the fundamental "stuff" from which the world is made. Democritus (460–370 BC) and other early Greek philosophers described the material world as made up of tiny indivisible particles they called *atomos,* meaning "indivisible or uncuttable." Later, however, Plato and Aristotle formulated the notion that there can be no ultimately indivisible particles, and the "atomic" view of matter faded for many centuries during which Aristotelean philosophy dominated Western culture.

The notion of **atoms** reemerged in Europe during the seventeenth century. As chemists learned to measure the amounts of elements that reacted with one another to form new substances, the ground was laid for an atomic theory that linked the idea of elements with the idea of atoms. That theory came from the work of John Dalton during the period from 1803 to 1807. Dalton's atomic theory was based on the four postulates given in ▼ FIGURE 2.1.

Dalton's theory explains several laws of chemical combination that were known during his time, including the *law of constant composition* (Section 1.2),* based on postulate 4:

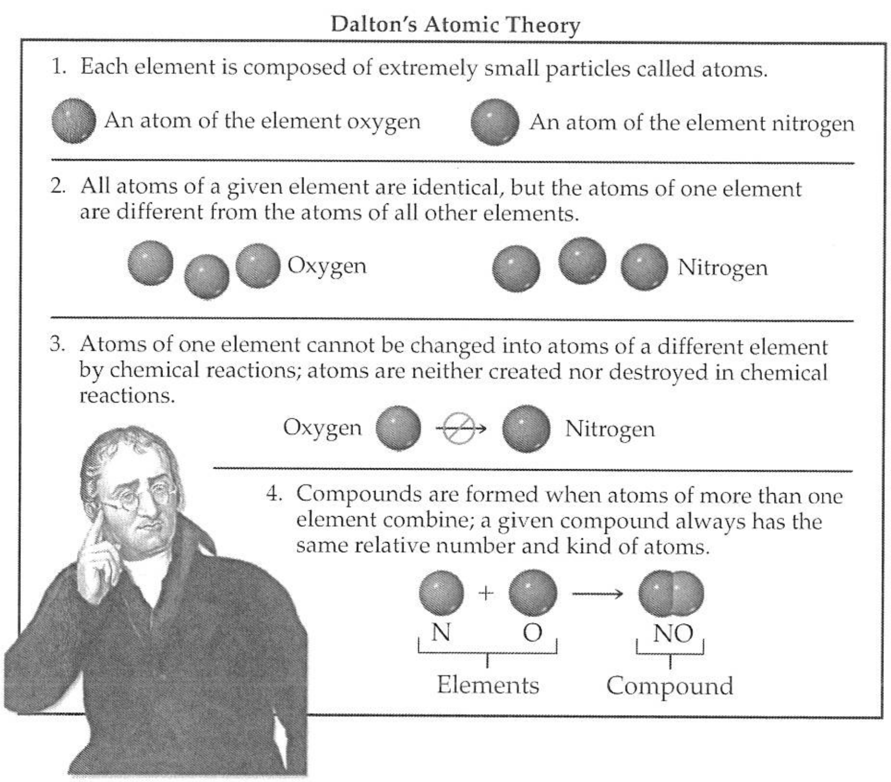
> In a given compound, the relative numbers and kinds of atoms are constant.

It also explains the *law of conservation of mass,* based on postulate 3:

> The total mass of materials present after a chemical reaction is the same as the total mass present before the reaction.

A good theory explains known facts and predicts new ones. Dalton used his theory to deduce the *law of multiple proportions*:

> If two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in the ratio of small whole numbers.




▶ FIGURE 2.1 **Dalton's atomic theory.** John Dalton (1766–1844), the son of a poor English weaver, began teaching at age 12. He spent most of his years in Manchester, where he taught both grammar school and college. His lifelong interest in meteorology led him to study gases, then chemistry, and eventually atomic theory. Despite his humble beginnings, Dalton gained a strong scientific reputation during his lifetime.

*The short chainlike symbol that precedes the section reference indicates a link to ideas presented earlier in the text.

We can illustrate this law by considering water and hydrogen peroxide, both of which consist of the elements hydrogen and oxygen. In forming water, 8.0 g of oxygen combine with 1.0 g of hydrogen. In forming hydrogen peroxide, 16.0 g of oxygen combine with 1.0 g of hydrogen. Thus, the ratio of the mass of oxygen per gram of hydrogen in the two compounds is 2:1. Using Dalton's atomic theory, we conclude that hydrogen peroxide contains twice as many atoms of oxygen per hydrogen atom as does water.

GIVE IT SOME THOUGHT

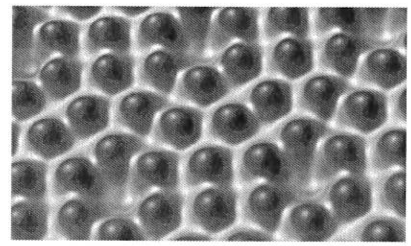
Compound A contains 1.333 g of oxygen per gram of carbon, whereas compound B contains 2.666 g of oxygen per gram of carbon.

a. What chemical law do these data illustrate?
b. If compound A has an equal number of oxygen and carbon atoms, what can we conclude about the composition of compound B?

# 2.2 THE DISCOVERY OF ATOMIC STRUCTURE

Dalton based his conclusions about atoms on chemical observations made in the laboratory. Neither he nor those who followed him during the century after his work was published had any direct evidence for the existence of atoms. Today, however, we can measure the properties of individual atoms and even provide images of them (▶ FIGURE 2.2).

As scientists developed methods for probing the nature of matter, the supposedly indivisible atom began to show signs of a more complex structure, and today we know that the atom is composed of **subatomic particles.** Before we summarize the current model, we briefly consider a few of the landmark discoveries that led to that model. We will see that the atom is composed in part of electrically charged particles, some with a positive charge and some with a negative charge. As we discuss the development of our current model of the atom, keep in mind this fact: *Particles with the same charge repel one another, whereas particles with unlike charges attract one another.*



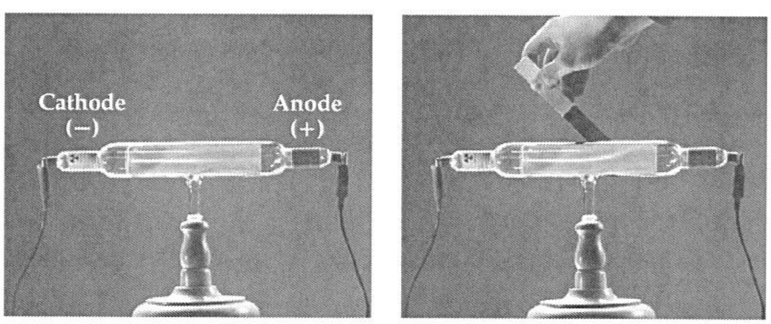
▲ FIGURE 2.2 **An image of the surface of silicon.** The image was obtained by a technique called scanning tunneling microscopy. The color was added to the image by computer to help distinguish its features. Each purple sphere is a silicon atom.

## Cathode Rays and Electrons

During the mid-1800s, scientists began to study electrical discharge through a glass tube pumped almost empty of air (▼ FIGURE 2.3). When a high voltage was applied to the electrodes in the tube, radiation was produced between the electrodes. This radiation, called **cathode rays**, originated at the negative electrode and traveled to the positive electrode. Although the rays could not be seen, their presence was detected because they cause certain materials to *fluoresce,* or to give off light.

GO FIGURE

How do we know that the cathode rays travel from cathode to anode?



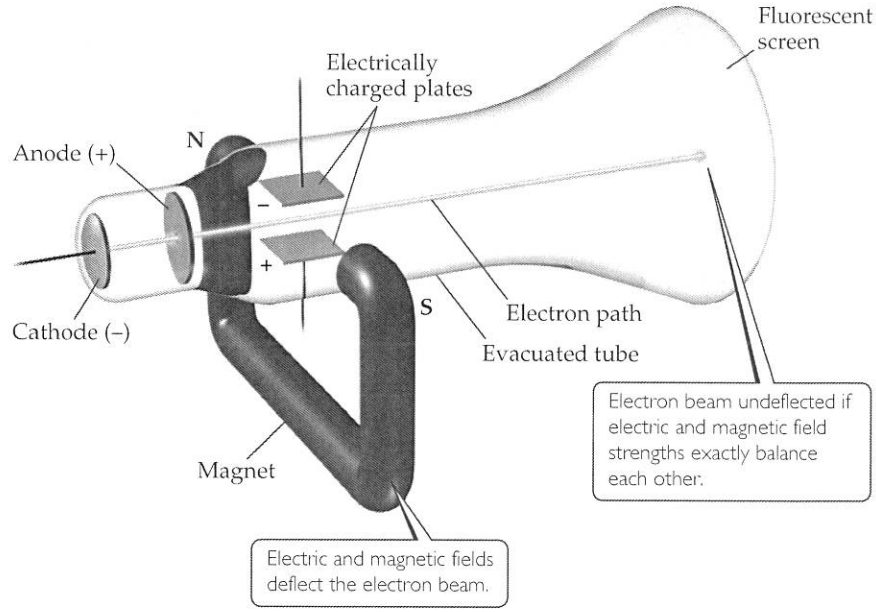

(a) Electrons move from the cathode (negative electrode) to the anode (positive electrode). The tube contains a glass screen (set diagonally to the electron beam) that fluoresces, showing the path of the cathode rays.

(b) The rays are deflected by a magnet.

◀ FIGURE 2.3 **Cathode-ray tube.**

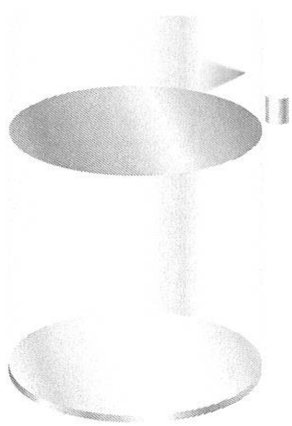
GO FIGURE

**If no magnetic field were applied, would you expect the electron beam to be deflected upward or downward by the electric field?**




▶ FIGURE 2.4 **Cathode-ray tube with perpendicular magnetic and electric fields.** The cathode rays (electrons) originate at the cathode and are accelerated toward the anode, which has a hole in its center. A narrow beam of electrons passes through the hole and travels to the fluorescent screen. The strengths of the electric and magnetic fields are adjusted so their effects cancel each other allowing the beam to travel a straight path.

Experiments showed that cathode rays are deflected by electric or magnetic fields in a way consistent with their being a stream of negative electrical charge. The British scientist J. J. Thomson (1856–1940) observed that cathode rays are the same regardless of the identity of the cathode material. In a paper published in 1897, Thomson described cathode rays as streams of negatively charged particles. His paper is generally accepted as the "discovery" of what became known as the *electron*.



Thomson constructed a cathode-ray tube having a hole in the anode through which a beam of electrons passed. Electrically charged plates and a magnet were positioned perpendicular to the electron beam, and a fluorescent screen was located at one end (▲ FIGURE 2.4). The electric field deflected the rays in one direction, and the magnetic field deflected them in the opposite direction. Thomson adjusted the strengths of the fields so that the effects balanced each other, allowing the electrons to travel in a straight path to the screen. Knowing the strengths that resulted in the straight path made it possible to calculate a value of $1.76 \times 10^8$ coulombs per gram for the ratio of the electron's electrical charge to its mass.*

Once the charge-to-mass ratio of the electron was known, measuring either quantity allowed scientists to calculate the other. In 1909, Robert Millikan (1868–1953) of the University of Chicago succeeded in measuring the charge of an electron by performing the experiment described in ◀ FIGURE 2.5. He then calculated the mass of the electron by using his experimental value for the charge, $1.602 \times 10^{-19}$ C, and Thomson's charge-to-mass ratio, $1.76 \times 10^8$ C/g:

$$\text{Electron mass} = \frac{1.602 \times 10^{-19}\ \text{C}}{1.76 \times 10^{8}\ \text{C/g}} = 9.10 \times 10^{-28}\ \text{g}$$

*The coulomb (C) is the SI unit for electrical charge.

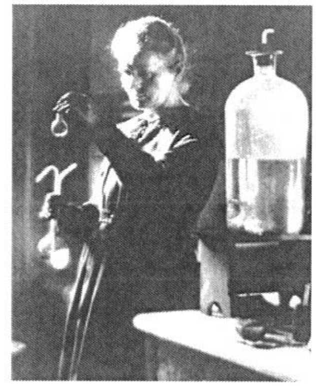
This result agrees well with the currently accepted value for the electron mass, $9.10938 \times 10^{-28}$ g. This mass is about 2000 times smaller than that of hydrogen, the lightest atom.

## Radioactivity

In 1896 the French scientist Henri Becquerel (1852–1908) discovered that a compound of uranium spontaneously emits high-energy radiation. This spontaneous emission of radiation is called **radioactivity.** At Becquerel's suggestion, Marie Curie (▶ FIGURE 2.6) and her husband, Pierre, began experiments to isolate the radioactive components of the compound.

Further study of radioactivity, principally by the British scientist Ernest Rutherford (▶ FIGURE 2.7), revealed three types of radiation: alpha ($\alpha$), beta ($\beta$), and gamma ($\gamma$). The paths of $\alpha$ and $\beta$ radiation are bent by an electric field, although in opposite directions; $\gamma$ radiation is unaffected by the field (▼ FIGURE 2.8).

Rutherford showed that $\alpha$ and $\beta$ rays consist of fast-moving particles. In fact, $\beta$ particles are high-speed electrons and can be considered the radioactive equivalent of cathode rays. They are attracted to a positively charged plate. The $\alpha$ particles have a positive charge and are attracted to a negative plate. In units of the charge of the electron, $\beta$ particles have a charge of $1-$ and $\alpha$ particles a charge of $2+$. Each $\alpha$ particle has a mass about 7400 times that of an electron. Gamma radiation is high-energy radiation similar to X-rays; it does not consist of particles and carries no charge.



▲ FIGURE 2.6 **Marie Sklodowska Curie (1867–1934).** When Marie Curie presented her doctoral thesis, it was described as the greatest single contribution of any doctoral thesis in the history of science. In 1903 Henri Becquerel, Maire Curie, and her husband, Pierre, were jointly awarded the Nobel Prize in Physics for their pioneering work on radioactivity (a term she introduced). In 1911 Marie Curie won a second Nobel Prize, this time in chemistry for her discovery of the elements polonium and radium.

## The Nuclear Model of the Atom

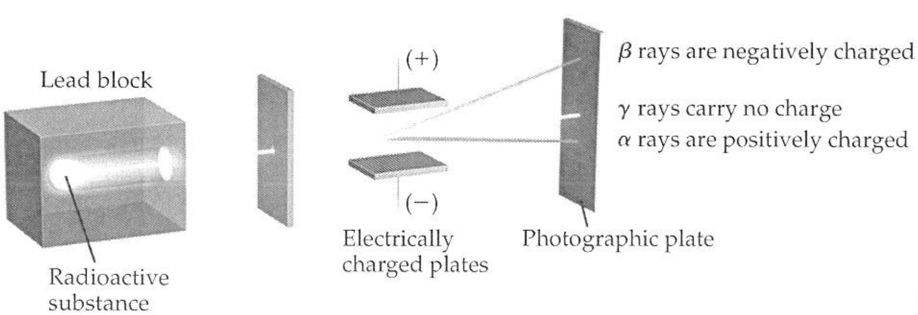
With growing evidence that the atom is composed of smaller particles, attention was given to how the particles fit together. During the early 1900s, Thomson reasoned that because electrons contribute only a very small fraction of an atom's mass they probably were responsible for an equally small fraction of the atom's size. He proposed that the atom consisted of a uniform positive sphere of matter in which the electrons were embedded like raisins in a pudding or seeds in a watermelon (▶ FIGURE 2.9). This *plum-pudding model,* named after a traditional English dessert, was very short-lived.

In 1910, Rutherford was studying the angles at which $\alpha$ particles were deflected, or *scattered,* as they passed through a thin sheet of gold foil (▶ FIGURE 2.10). He discovered that almost all the particles passed directly through the foil without deflection, with a few particles deflected about 1 degree, consistent with Thomson's plum-pudding model. For the sake of completeness, Rutherford suggested that Ernest Marsden, an undergraduate student working in the laboratory, look for scattering at large angles. To everyone's surprise, a small amount of scattering was observed at large angles, with some particles scattered back in the direction from which they had come. The explanation for these results was not immediately obvious, but they were clearly inconsistent with Thomson's plum-pudding model.

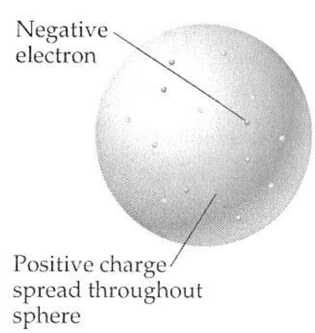

▲ FIGURE 2.7 **Ernest Rutherford (1871–1937).** In 1895, Rutherford was awarded a position at Cambridge University in England, where he worked with J. J. Thomson. In 1898 he moved to McGill University in Montreal, where he did the research on radioactivity that led to his 1908 Nobel Prize in Chemistry. In 1907 Rutherford returned to England as a faculty member at Manchester University, where in 1910 he performed his famous $\alpha$-particle scattering experiments. In 1992 his native New Zealand honored him by putting his likeness on their $100 currency note.

GO FIGURE

**Which of the three kinds of radiation shown consists of electrons? Why are these rays deflected to a greater extent than the others?**



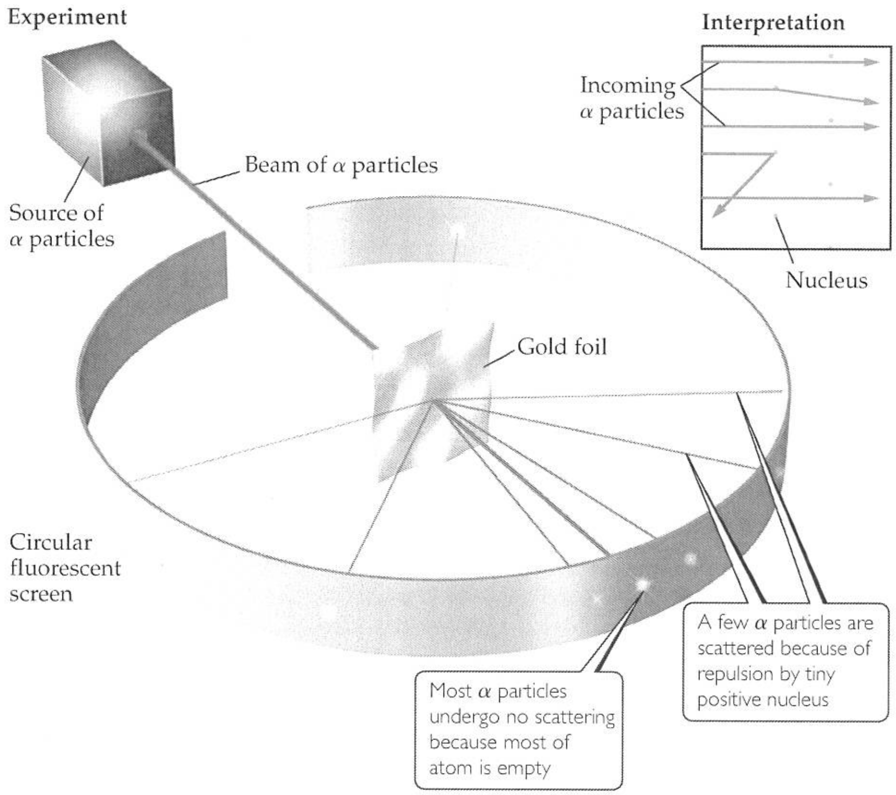

▲ FIGURE 2.8 **Behavior of alpha ($\alpha$), beta ($\beta$), and gamma ($\gamma$) rays in an electric field.**




▲ FIGURE 2.9 **J. J. Thomson's plum-pudding model of the atom.** Ernest Rutherford proved this model wrong.

GO FIGURE

What is the charge on the particles that form the beam?




▶ FIGURE 2.10 **Rutherford's $\alpha$-scattering experiment.** When $\alpha$ particles pass through a gold foil, most pass through undeflected but some are scattered, a few at very large angles. According to the plum-pudding model of the atom, the particles should experience only very minor deflections. The nuclear model of the atom explains why a few $\alpha$ particles are deflected at large angles. For clarity, the nuclear atom is shown here as a colored sphere, but most of the space around the nucleus is empty except for the tiny electrons moving around.

Rutherford explained the results by postulating the **nuclear model** of the atom, a model in which most of the mass of each gold atom and all of its positive charge reside in a very small, extremely dense region that he called the **nucleus**. He postulated further that most of the volume of an atom is empty space in which electrons move around the nucleus. In the $\alpha$-scattering experiment, most of the particles passed through the foil unscattered because they did not encounter the minute nucleus of any gold atom. Occasionally, however, an $\alpha$ particle came close to a gold nucleus. The repulsion between the highly positive charge of the gold nucleus and the positive charge of the $\alpha$ particle was then strong enough to deflect the particle, as shown in Figure 2.10.

Subsequent experiments led to the discovery of positive particles (*protons*) and neutral particles (*neutrons*) in the nucleus. Protons were discovered in 1919 by Rutherford and neutrons in 1932 by British scientist James Chadwick (1891–1972). Thus, the atom is composed of electrons, protons, and neutrons.

GIVE IT SOME THOUGHT

What happens to most of the $\alpha$ particles that strike the gold foil in Rutherford's experiment? Why do they behave that way?

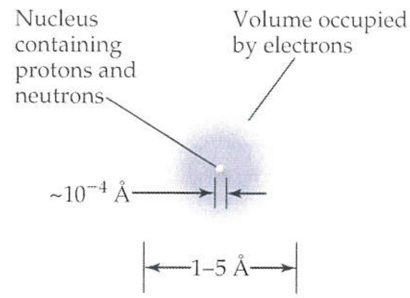
## 2.3 THE MODERN VIEW OF ATOMIC STRUCTURE

Since Rutherford's time, as physicists have learned more and more about atomic nuclei, the list of particles that make up nuclei has grown and continues to increase. As chemists, however, we can take a simple view of the atom because only three subatomic particles—the **proton**, **neutron**, and **electron**—have a bearing on chemical behavior.

As noted earlier, the charge of an electron is $-1.602 \times 10^{-19}$ C. That of a proton is equal in magnitude, $+1.602 \times 10^{-19}$ C. The quantity $1.602 \times 10^{-19}$ C is called the

**electronic charge**. For convenience, the charges of atomic and subatomic particles are usually expressed as multiples of this charge rather than in coulombs. Thus, the charge of the electron is 1− and that of the proton is 1+. Neutrons are electrically neutral (which is how they received their name). *Every atom has an equal number of electrons and protons, so atoms have no net electrical charge.*

Protons and neutrons reside in the tiny nucleus of the atom. The vast majority of an atom's volume is the space in which the electrons reside (▶ FIGURE 2.11). The electrons are attracted to the protons in the nucleus by the electrostatic force that exists between particles of opposite electrical charge. In later chapters we will see that the strength of the attractive forces between electrons and nuclei can be used to explain many of the differences among different elements.




▲ FIGURE 2.11 **The structure of the atom.** A cloud of rapidly moving electrons occupies most of the volume of the atom. The nucleus occupies a tiny region at the center of the atom and is composed of the protons and neutrons. The nucleus contains virtually all the mass of the atom.

GIVE IT SOME THOUGHT

a. If an atom has 15 protons, how many electrons does it have?
b. Where do the protons reside in an atom?

Atoms have extremely small masses. The mass of the heaviest known atom, for example, is approximately $4 \times 10^{-22}$ g. Because it would be cumbersome to express such small masses in grams, we use the **atomic mass unit** (amu),* where 1 amu = $1.66054 \times 10^{-24}$ g. A proton has a mass of 1.0073 amu, a neutron 1.0087 amu, and an electron $5.486 \times 10^{-4}$ amu (▼ TABLE 2.1). Because it takes 1836 electrons to equal the mass of one proton or one neutron, the nucleus contains most of the mass of an atom.

Most atoms have diameters between $1 \times 10^{-10}$ m and $5 \times 10^{-10}$ m. A convenient non–SI unit of length used for atomic dimensions is the **angstrom** (Å), where 1 Å = $1 \times 10^{-10}$ m. Thus, atoms have diameters of approximately 1–5 Å. The diameter of a chlorine atom, for example, is 200 pm, or 2.0 Å.

**SAMPLE EXERCISE 2.1** Atomic Size

The diameter of a US dime is 17.9 mm, and the diameter of a silver atom is 2.88 Å. How many silver atoms could be arranged side by side across the diameter of a dime?

SOLUTION

The unknown is the number of silver (Ag) atoms. Using the relationship 1 Ag atom = 2.88 Å as a conversion factor relating number of atoms and distance, we start with the diameter of the dime, first converting this distance into angstroms and then using the diameter of the Ag atom to convert distance to number of Ag atoms:

$$\text{Ag atoms} = (17.9\ \cancel{\text{mm}})\left(\frac{10^{-3}\ \cancel{\text{m}}}{1\ \cancel{\text{mm}}}\right)\left(\frac{1\ \cancel{\text{Å}}}{10^{-10}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{Ag atom}}{2.88\ \cancel{\text{Å}}}\right) = 6.22 \times 10^{7}\ \text{Ag atoms}$$

That is, 62.2 million silver atoms could sit side by side across a dime!

PRACTICE EXERCISE

The diameter of a carbon atom is 1.54 Å. **(a)** Express this diameter in picometers. **(b)** How many carbon atoms could be aligned side by side across the width of a pencil line that is 0.20 mm wide?

***Answers:*** (a) 154 pm, (b) $1.3 \times 10^{6}$ C atoms

The diameter of an atomic nucleus is approximately $10^{-4}$ Å, only a small fraction of the diameter of the atom as a whole. You can appreciate the relative sizes of the atom and its nucleus by imagining that if the hydrogen atom were as large as a football stadium,

TABLE 2.1 • Comparison of the Proton, Neutron, and Electron

| Particle | Charge | Mass (amu) |
|---|---|---|
| Proton | Positive (1+) | 1.0073 |
| Neutron | None (neutral) | 1.0087 |
| Electron | Negative (1−) | $5.486 \times 10^{-4}$ |

*The SI abbreviation for the atomic mass unit is u. We will use the more common abbreviation amu.

## A CLOSER LOOK

### BASIC FORCES

Four basic forces are known in nature: (1) gravitational, (2) electromagnetic, (3) strong nuclear, and (4) weak nuclear. *Gravitational forces* are attractive forces that act between all objects in proportion to their masses. Gravitational forces between atoms or between subatomic particles are so small that they are of no chemical significance.

*Electromagnetic forces* are attractive or repulsive forces that act between either electrically charged or magnetic objects. Electric forces are important in understanding the chemical behavior of atoms. The magnitude of the electric force between two charged particles is given by *Coulomb's law*: $F = kQ_1Q_2/d^2$, where $Q_1$ and $Q_2$ are the magnitudes of the charges on the two particles, $d$ is the distance between their centers, and $k$ is a constant determined by the units for $Q$ and $d$. A negative value for the force indicates attraction, whereas a positive value indicates repulsion.

All nuclei except those of hydrogen atoms contain two or more protons. Because like charges repel, electrical repulsion would cause the protons to fly apart if the *strong nuclear force* did not keep them together. This force acts between subatomic particles, as in the nucleus. At this distance, the attractive strong nuclear force is stronger than the positive–positive repulsive electric force and holds the nucleus together.

The *weak nuclear force* is weaker than the electric force but stronger than the gravitational force. We are aware of its existence only because it shows itself in certain types of radioactivity.

*RELATED EXERCISE:* 2.88

the nucleus would be the size of a small marble. Because the tiny nucleus carries most of the mass of the atom in such a small volume, it has an incredibly high density—on the order of $10^{13}-10^{14}\ g/cm^3$. A matchbox full of material of such density would weigh over 2.5 billion tons!

An illustration of the atom that incorporates the features we have just discussed is shown in Figure 2.11. The electrons play the major role in chemical reactions. The significance of representing the region containing the electrons as an indistinct cloud will become clear in later chapters when we consider the energies and spatial arrangements of the electrons.

## Atomic Numbers, Mass Numbers, and Isotopes

What makes an atom of one element different from an atom of another element is that the atoms of each element have a *characteristic number of protons.* Indeed, the number of protons in an atom of any particular element is called that element's **atomic number.** Because an atom has no net electrical charge, the number of electrons it contains must equal the number of protons. All atoms of carbon, for example, have six protons and six electrons, whereas all atoms of oxygen have eight protons and eight electrons. Thus, carbon has atomic number 6, and oxygen has atomic number 8. The atomic number of each element is listed with the name and symbol of the element on the inside front cover of the text.

Atoms of a given element can differ in the number of neutrons they contain and, consequently, in mass. For example, most atoms of carbon have six neutrons, although some have more and some have less. The symbol $^{12}_{6}C$ (read "carbon twelve," carbon-12) represents the carbon atom containing six protons and six neutrons. The atomic number is shown by the subscript; the superscript, called the **mass number,** is the number of protons plus neutrons in the atom:

Mass number (number of protons plus neutrons)

$^{12}_{6}C \longleftarrow$ Symbol of element

Atomic number (number of protons or electrons)

Because all atoms of a given element have the same atomic number, the subscript is redundant and is often omitted. Thus, the symbol for carbon-12 can be represented simply as $^{12}C$. As one more example of this notation, carbon atoms that contain six protons and eight neutrons have mass number 14, are represented as $^{14}_{6}C$ or $^{14}C$, and are referred to as carbon-14.

TABLE 2.2 • Some Isotopes of Carbon*

| Symbol | Number of Protons | Number of Electrons | Number of Neutrons |
|---|---|---|---|
| $^{11}C$ | 6 | 6 | 5 |
| $^{12}C$ | 6 | 6 | 6 |
| $^{13}C$ | 6 | 6 | 7 |
| $^{14}C$ | 6 | 6 | 8 |

*Almost 99% of the carbon found in nature is $^{12}C$.

Atoms with identical atomic numbers but different mass numbers (that is, same number of protons but different numbers of neutrons) are called **isotopes** of one another. Several isotopes of carbon are listed in ▲ TABLE 2.2. We will generally use the notation with superscripts only when referring to a particular isotope of an element.

**SAMPLE EXERCISE 2.2** **Determining the Number of Subatomic Particles in Atoms**

How many protons, neutrons, and electrons are in (**a**) an atom of $^{197}Au$, (**b**) an atom of strontium-90?

**SOLUTION**

(**a**) The superscript 197 is the mass number (protons + neutrons). According to the list of elements given on the inside front cover, gold has atomic number 79. Consequently, an atom of $^{197}Au$ has 79 protons, 79 electrons, and $197 - 79 = 118$ neutrons. (**b**) The atomic number of strontium (listed on inside front cover) is 38. Thus, all atoms of this element have 38 protons and 38 electrons. The strontium-90 isotope has $90 - 38 = 52$ neutrons.

**PRACTICE EXERCISE**

How many protons, neutrons, and electrons are in (**a**) a $^{138}Ba$ atom, (**b**) an atom of phosphorus-31?

***Answer:*** (**a**) 56 protons, 56 electrons, and 82 neutrons, (**b**) 15 protons, 15 electrons, and 16 neutrons

**SAMPLE EXERCISE 2.3** **Writing Symbols for Atoms**

Magnesium has three isotopes with mass numbers 24, 25, and 26. (**a**) Write the complete chemical symbol (superscript and subscript) for each. (**b**) How many neutrons are in an atom of each isotope?

**SOLUTION**

(**a**) Magnesium has atomic number 12, so all atoms of magnesium contain 12 protons and 12 electrons. The three isotopes are therefore represented by $^{24}_{12}Mg$, $^{25}_{12}Mg$, and $^{26}_{12}Mg$. (**b**) The number of neutrons in each isotope is the mass number minus the number of protons. The numbers of neutrons in an atom of each isotope are therefore 12, 13, and 14, respectively.

**PRACTICE EXERCISE**

Give the complete chemical symbol for the atom that contains 82 protons, 82 electrons, and 126 neutrons.

***Answer:*** $^{208}_{82}Pb$

# 2.4 ATOMIC WEIGHTS

Atoms are small pieces of matter, so they have mass. In this section we discuss the mass scale used for atoms and introduce the concept of *atomic weights*.

## The Atomic Mass Scale

Scientists of the nineteenth century were aware that atoms of different elements have different masses. They found, for example, that each 100.0 g of water contains 11.1 g of hydrogen and 88.9 g of oxygen. Thus, water contains $88.9/11.1 = 8$ times as much oxygen, by mass, as hydrogen. Once scientists understood that water contains two hydrogen atoms for each oxygen atom, they concluded that an oxygen atom must have $2 \times 8 = 16$ times as much mass as a hydrogen atom. Hydrogen, the lightest atom, was arbitrarily assigned a relative mass of 1 (no units). Atomic masses of other elements were at first determined relative to this value. Thus, oxygen was assigned an atomic mass of 16.

Today we can determine the masses of individual atoms with a high degree of accuracy. For example, we know that the $^{1}H$ atom has a mass of $1.6735 \times 10^{-24}$ g and the $^{16}O$ atom has a mass of $2.6560 \times 10^{-23}$ g. As we noted in Section 2.3, it is convenient to use the *atomic mass unit* (amu) when dealing with these extremely small masses:

$$1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g and } 1 \text{ g} = 6.02214 \times 10^{23} \text{ amu}$$

The atomic mass unit is presently defined by assigning a mass of exactly 12 amu to an atom of the $^{12}C$ isotope of carbon. In these units, an $^{1}H$ atom has a mass of 1.0078 amu and an $^{16}O$ atom has a mass of 15.9949 amu.

## Atomic Weight

Most elements occur in nature as mixtures of isotopes. We can determine the *average atomic mass* of an element, usually called the element's **atomic weight**, by using the masses of its isotopes and their relative abundances:

$$\text{Atomic weight} = \sum[(\text{isotope mass}) \times (\text{fractional isotope abundance})] \quad \text{over all isotopes of the element} \qquad [2.1]$$

Naturally occurring carbon, for example, is composed of 98.93% $^{12}C$ and 1.07% $^{13}C$. The masses of these isotopes are 12 amu (exactly) and 13.00335 amu, respectively, making the atomic weight of carbon

$$(0.9893)(12 \text{ amu}) + (0.0107)(13.00335 \text{ amu}) = 12.01 \text{ amu}$$

The atomic weights of the elements are listed in both the periodic table and the table of elements inside the front cover of this text.

### GIVE IT SOME THOUGHT

A particular atom of chromium has a mass of 52.94 amu, whereas the atomic weight of chromium is 51.99 amu. Explain the difference in the two masses.

### SAMPLE EXERCISE 2.4 Calculating the Atomic Weight of an Element from Isotopic Abundances

Naturally occurring chlorine is 75.78% $^{35}Cl$ (atomic mass 34.969 amu) and 24.22% $^{37}Cl$ (atomic mass 36.966 amu). Calculate the atomic weight of chlorine.

#### SOLUTION

We can calculate the atomic weight by multiplying the abundance of each isotope by its atomic mass and summing these products. Because 75.78% = 0.7578 and 24.22% = 0.2422, we have

$$\begin{aligned}\text{Atomic weight} &= (0.7578)(34.969 \text{ amu}) + (0.2422)(36.966 \text{ amu}) \\ &= 26.50 \text{ amu} + 8.953 \text{ amu} \\ &= 35.45 \text{ amu}\end{aligned}$$

This answer makes sense: The atomic weight, which is actually the average atomic mass, is between the masses of the two isotopes and is closer to the value of $^{35}Cl$, the more abundant isotope.

#### PRACTICE EXERCISE

Three isotopes of silicon occur in nature: $^{28}Si$ (92.23%), atomic mass 27.97693 amu; $^{29}Si$ (4.68%), atomic mass 28.97649 amu; and $^{30}Si$ (3.09%), atomic mass 29.97377 amu. Calculate the atomic weight of silicon.

***Answer:*** 28.09 amu

## A CLOSER LOOK

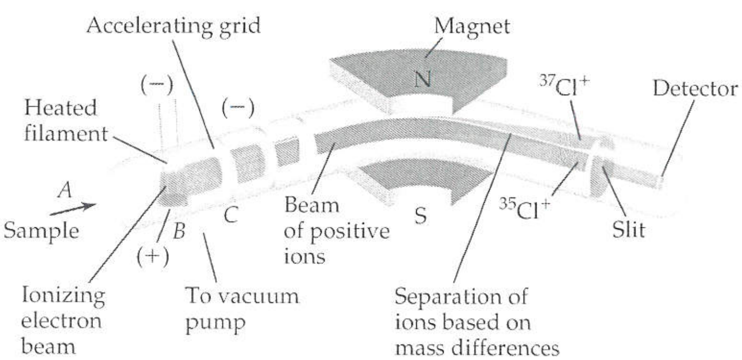
### THE MASS SPECTROMETER

The most accurate means for determining atomic weights is provided by the **mass spectrometer** (▼ FIGURE 2.12). A gaseous sample is introduced at *A* and bombarded by a stream of high-energy electrons at *B*. Collisions between the electrons and the atoms or molecules of the gas produce positively charged particles that are then accelerated toward a negatively charged grid (*C*). After the particles pass through the grid, they encounter two slits that allow only a narrow beam of particles to pass. This beam then passes between the poles of a magnet, which deflects the particles into a curved path. For particles with the same charge, the extent of deflection depends on mass—the more massive the particle, the less the deflection. The particles are thereby separated according to their masses. By changing the strength of the magnetic field or the accelerating voltage on the grid, charged particles of various masses can be selected to enter the detector.

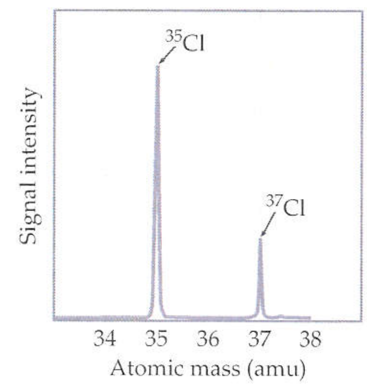
A graph of the intensity of the detector signal versus particle atomic mass is called a *mass spectrum* (▼ FIGURE 2.13). Analysis of a mass spectrum gives both the masses of the charged particles reaching the detector and their relative abundances, which are obtained from the signal intensities. Knowing the atomic mass and the abundance of each isotope allows us to calculate the atomic weight of an element, as shown in Sample Exercise 2.4.

Mass spectrometers are used extensively today to identify chemical compounds and analyze mixtures of substances. Any molecule that loses electrons can fall apart, forming an array of positively charged fragments. The mass spectrometer measures the masses of these fragments, producing a chemical "fingerprint" of the molecule and providing clues about how the atoms were connected in the original molecule. Thus, a chemist might use this technique to determine the molecular structure of a newly synthesized compound or to identify a pollutant in the environment.

*RELATED EXERCISES:* 2.33, 2.34, 2.35(b), 2.36, 2.92, and 2.93




▲ FIGURE 2.12 **A mass spectrometer.** Cl atoms are introduced at *A* and are ionized to form $Cl^+$ ions, which are then directed through a magnetic field. The paths of the ions of the two Cl isotopes diverge as they pass through the field.




▲ FIGURE 2.13 **Mass spectrum of atomic chlorine.** The fractional abundances of the isotopes $^{35}Cl$ and $^{37}Cl$ are indicated by the relative signal intensities of the beams reaching the detector of the mass spectrometer.

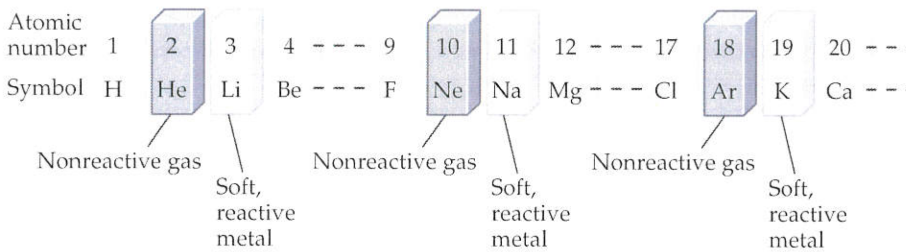
## 2.5 THE PERIODIC TABLE

As the list of known elements expanded during the early 1800s, attempts were made to find patterns in chemical behavior. These efforts culminated in the development of the periodic table in 1869. We will have much to say about the periodic table in later chapters, but it is so important and useful that you should become acquainted with it now. You will quickly learn that *the periodic table is the most significant tool that chemists use for organizing and remembering chemical facts.*

Many elements show strong similarities to one another. The elements lithium (Li), sodium (Na), and potassium (K) are all soft, very reactive metals, for example. The elements helium (He), neon (Ne), and argon (Ar) are all very nonreactive gases. If the elements are arranged in order of increasing atomic number, their chemical and physical properties show a repeating, or *periodic,* pattern. For example, each of the soft, reactive metals—lithium, sodium, and potassium—comes immediately after one of the nonreactive gases—helium, neon, and argon—as shown in ▶ FIGURE 2.14.

GO FIGURE

**If F is a reactive nonmetal, which other element or elements shown here do you expect to also be a reactive nonmetal?**




▶ FIGURE 2.14 **Arranging elements by atomic number reveals a periodic pattern of properties.** This pattern is the basis of the periodic table.

The arrangement of elements in order of increasing atomic number, with elements having similar properties placed in vertical columns, is known as the **periodic table** (▼ FIGURE 2.15). The table shows the atomic number and atomic symbol for each element, and the atomic weight is often given as well, as in this typical entry for potassium:

| | |
|---|---|
| 19 | ← atomic number |
| K | ← atomic symbol |
| 39.0983 | ← atomic weight |

You may notice slight variations in periodic tables from one book to another or between those in the lecture hall and in the text. These are simply matters of style, or they might concern the particular information included. There are no fundamental differences.

Periods - horizontal rows

Groups - vertical columns containing elements with similar properties

Elements arranged in order of increasing atomic number

Steplike line divides metals from nonmetals

| | 1A 1 | 2A 2 | 3B 3 | 4B 4 | 5B 5 | 6B 6 | 7B 7 | 8B 8 | 8B 9 | 8B 10 | 1B 11 | 2B 12 | 3A 13 | 4A 14 | 5A 15 | 6A 16 | 7A 17 | 8A 18 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 | 1 H | | | | | | | | | | | | | | | | | 2 He |
| 2 | 3 Li | 4 Be | | | | | | | | | | | 5 B | 6 C | 7 N | 8 O | 9 F | 10 Ne |
| 3 | 11 Na | 12 Mg | | | | | | | | | | | 13 Al | 14 Si | 15 P | 16 S | 17 Cl | 18 Ar |
| 4 | 19 K | 20 Ca | 21 Sc | 22 Ti | 23 V | 24 Cr | 25 Mn | 26 Fe | 27 Co | 28 Ni | 29 Cu | 30 Zn | 31 Ga | 32 Ge | 33 As | 34 Se | 35 Br | 36 Kr |
| 5 | 37 Rb | 38 Sr | 39 Y | 40 Zr | 41 Nb | 42 Mo | 43 Tc | 44 Ru | 45 Rh | 46 Pd | 47 Ag | 48 Cd | 49 In | 50 Sn | 51 Sb | 52 Te | 53 I | 54 Xe |
| 6 | 55 Cs | 56 Ba | 71 Lu | 72 Hf | 73 Ta | 74 W | 75 Re | 76 Os | 77 Ir | 78 Pt | 79 Au | 80 Hg | 81 Tl | 82 Pb | 83 Bi | 84 Po | 85 At | 86 Rn |
| 7 | 87 Fr | 88 Ra | 103 Lr | 104 Rf | 105 Db | 106 Sg | 107 Bh | 108 Hs | 109 Mt | 110 Ds | 111 Rg | 112 Cn | 113 | 114 | 115 | 116 | 117 | 118 |

| | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 57 La | 58 Ce | 59 Pr | 60 Nd | 61 Pm | 62 Sm | 63 Eu | 64 Gd | 65 Tb | 66 Dy | 67 Ho | 68 Er | 69 Tm | 70 Yb |
| 89 Ac | 90 Th | 91 Pa | 92 U | 93 Np | 94 Pu | 95 Am | 96 Cm | 97 Bk | 98 Cf | 99 Es | 100 Fm | 101 Md | 102 No |

Metals
Metalloids
Nonmetals

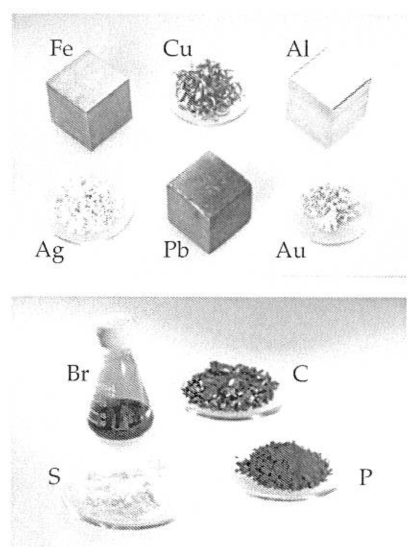
▲ FIGURE 2.15 **Periodic table of the elements.**

The horizontal rows of the periodic table are called **periods**. The first period consists of only two elements, hydrogen (H) and helium (He). The second and third periods consist of eight elements each. The fourth and fifth periods contain 18 elements. The sixth period has 32 elements, but for it to fit on a page, 14 of these elements (atomic numbers 57–70) appear at the bottom of the table. The seventh period is incomplete, but it also has 14 of its members placed in a row at the bottom of the table.

The vertical columns are **groups**. The way in which the groups are labeled is somewhat arbitrary. Three labeling schemes are in common use, two of which are shown in Figure 2.15. The top set of labels, which have A and B designations, is widely used in North America. Roman numerals, rather than Arabic ones, are often employed in this scheme. Group 7A, for example, is often labeled VIIA. Europeans use a similar convention that numbers the columns from 1A through 8A and then from 1B through 8B, thereby giving the label 7B (or VIIB) instead of 7A to the group headed by fluorine (F). In an effort to eliminate this confusion, the International Union of Pure and Applied Chemistry (IUPAC) has proposed a convention that numbers the groups from 1 through 18 with no A or B designations, as shown in Figure 2.15. We will use the traditional North American convention with Arabic numerals and the letters A and B.

Elements in a group often exhibit similarities in physical and chemical properties. For example, the "coinage metals"—copper (Cu), silver (Ag), and gold (Au)—belong to group 1B. These elements are less reactive than most metals, which is why they are used throughout the world to make coins. Many other groups in the periodic table also have names, listed in ▼ TABLE 2.3.

We will learn in Chapters 6 and 7 that elements in a group have similar properties because they have the same arrangement of electrons at the periphery of their atoms. However, we need not wait until then to make good use of the periodic table; after all, chemists who knew nothing about electrons developed the table! We can use the table, as they intended, to correlate behaviors of elements and to help us remember many facts. The color code of Figure 2.15 shows that, except for hydrogen, all the elements on the left and in the middle of the table are **metallic elements**, or **metals**. All the metallic elements share characteristic properties, such as luster and high electrical and heat conductivity, and all of them except mercury (Hg) are solid at room temperature. The metals are separated from the **nonmetallic elements**, or **nonmetals**, by a stepped line that runs from boron (B) to astatine (At). (Note that hydrogen, although on the left side of the table, is a nonmetal.) At room temperature some of the nonmetals are gaseous, some are solid, and one is liquid. Nonmetals generally differ from the metals in appearance (► FIGURE 2.16) and in other physical properties. Many of the elements that lie along the line that separates metals from nonmetals have properties that fall between those of metals and those of nonmetals. These elements are often referred to as **metalloids**.




▲ FIGURE 2.16 **Examples of metals (top) and nonmetals (bottom).**

### GIVE IT SOME THOUGHT

Chlorine is a halogen (Table 2.3). Locate this element in the periodic table.

a. What is its symbol?
b. In which period and in which group is the element located?
c. What is its atomic number?
d. Is it a metal or nonmetal?

**TABLE 2.3 • Names of Some Groups in the Periodic Table**

| Group | Name | Elements |
|---|---|---|
| 1A | Alkali metals | Li, Na, K, Rb, Cs, Fr |
| 2A | Alkaline earth metals | Be, Mg, Ca, Sr, Ba, Ra |
| 6A | Chalcogens | O, S, Se, Te, Po |
| 7A | Halogens | F, Cl, Br, I, At |
| 8A | Noble gases (or rare gases) | He, Ne, Ar, Kr, Xe, Rn |

## A CLOSER LOOK

### GLENN SEABORG AND SEABORGIUM

Prior to 1940 the periodic table ended at uranium, element number 92. Since that time, no scientist has had a greater effect on the periodic table than Glenn Seaborg (▶ FIGURE 2.17). In 1940 Seaborg, Edwin McMillan, and coworkers at the University of California, Berkeley, succeeded in isolating plutonium (Pu) as a product of the reaction between uranium and neutrons. We will talk about reactions of this type, called *nuclear reactions,* in Chapter 21.

Between 1944 and 1958, Seaborg and his coworkers also identified various products of nuclear reactions as being the elements having atomic numbers 95 through 102. All these elements are radioactive and are not found in nature; they can be synthesized only via nuclear reactions. For their efforts in identifying the elements beyond uranium (the *transuranium* elements), McMillan and Seaborg shared the 1951 Nobel Prize in Chemistry.

From 1961 to 1971, Seaborg served as the chairman of the US Atomic Energy Commission (now the Department of Energy). In this position he had an important role in establishing international treaties to limit the testing of nuclear weapons. Upon his return to Berkeley, he was part of the team that in 1974 first identified element number 106. In 1994, to honor Seaborg's many contributions to the discovery of new elements, the American Chemical Society proposed that element number 106 be named seaborgium (Sg). After several years of controversy about whether an element should be named after a living person, the IUPAC officially adopted the name in 1997. Seaborg became the first person to have an element named after him while he was alive.

◀ FIGURE 2.17 **Glenn Seaborg (1912–1999).** Seaborg at Berkeley in 1941 measuring radiation produced by plutonium.

*RELATED EXERCISE:* 2.95

**SAMPLE EXERCISE 2.5** **Using the Periodic Table**

Which two of these elements would you expect to show the greatest similarity in chemical and physical properties: B, Ca, F, He, Mg, P?

**SOLUTION**

Elements in the same group of the periodic table are most likely to exhibit similar properties. We therefore expect Ca and Mg to be most alike because they are in the same group (2A, the alkaline earth metals).

**PRACTICE EXERCISE**

Locate Na (sodium) and Br (bromine) in the periodic table. Give the atomic number of each and classify each as metal, metalloid, or nonmetal.

***Answer:*** Na, atomic number 11, is a metal; Br, atomic number 35, is a nonmetal.

## 2.6 MOLECULES AND MOLECULAR COMPOUNDS

Even though the atom is the smallest representative sample of an element, only the noble-gas elements are normally found in nature as isolated atoms. Most matter is composed of molecules or ions. We examine molecules here and ions in Section 2.7.

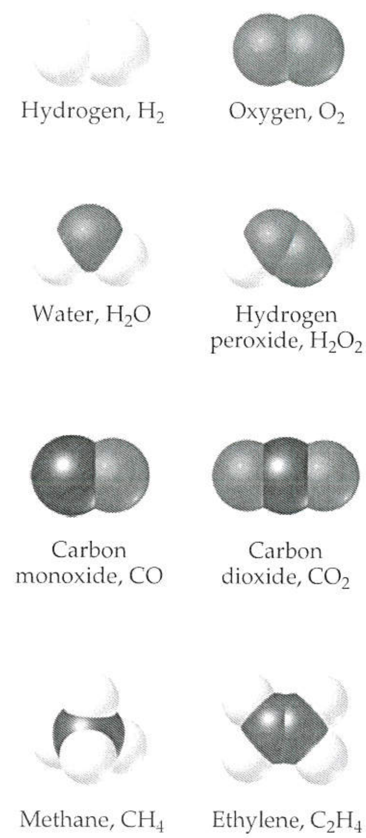
### Molecules and Chemical Formulas

Several elements are found in nature in molecular form—two or more of the same type of atom bound together. For example, most of the oxygen in air consists of molecules that contain two oxygen atoms. As we saw in Section 1.2, we represent this molecular oxygen by the **chemical formula** $O_2$ (read "oh two"). The subscript tells us that two oxygen atoms are present in each molecule. A molecule made up of two atoms is called a **diatomic molecule.**

Oxygen also exists in another molecular form known as *ozone*. Molecules of ozone consist of three oxygen atoms, making the chemical formula $O_3$. Even though "normal" oxygen ($O_2$) and ozone ($O_3$) are both composed only of oxygen atoms, they exhibit very different chemical and physical properties. For example, $O_2$ is essential for life, but $O_3$ is toxic; $O_2$ is odorless, whereas $O_3$ has a sharp, pungent smell.

The elements that normally occur as diatomic molecules are hydrogen, oxygen, nitrogen, and the halogens ($H_2$, $O_2$, $N_2$, $F_2$, $Cl_2$, $Br_2$, and $I_2$). Except for hydrogen, these diatomic elements are clustered on the right side of the periodic table.

Compounds composed of molecules contain more than one type of atom and are called **molecular compounds.** A molecule of the compound methane, for example, consists of one carbon atom and four hydrogen atoms and is therefore represented by the chemical formula $CH_4$. Lack of a subscript on the C indicates one atom of C per methane molecule. Several common molecules of both elements and compounds are shown in ► FIGURE 2.18. Notice how the composition of each substance is given by its chemical formula. Notice also that these substances are composed only of nonmetallic elements. *Most molecular substances we will encounter contain only nonmetals.*




▲ FIGURE 2.18 **Molecular models.** Notice how the chemical formulas of these simple molecules correspond to their compositions.

## Molecular and Empirical Formulas

Chemical formulas that indicate the actual numbers of atoms in a molecule are called **molecular formulas.** (The formulas in Figure 2.18 are molecular formulas.) Chemical formulas that give only the relative number of atoms of each type in a molecule are called **empirical formulas.** The subscripts in an empirical formula are always the smallest possible whole-number ratios. The molecular formula for hydrogen peroxide is $H_2O_2$, for example, whereas its empirical formula is HO. The molecular formula for ethylene is $C_2H_4$, and its empirical formula is $CH_2$. For many substances, the molecular formula and the empirical formula are identical, as in the case of water, $H_2O$.

Whenever we know the molecular formula of a compound, we can determine its empirical formula. The converse is not true, however. If we know the empirical formula of a substance, we cannot determine its molecular formula unless we have more information. So why do chemists bother with empirical formulas? As we will see in Chapter 3, certain common methods of analyzing substances lead to the empirical formula only. Once the empirical formula is known, additional experiments can give the information needed to convert the empirical formula to the molecular one. In addition, there are substances that do not exist as isolated molecules. For these substances, we must rely on empirical formulas.

**SAMPLE EXERCISE 2.6** **Relating Empirical and Molecular Formulas**

Write the empirical formulas for **(a)** glucose, a substance also known as either blood sugar or dextrose, molecular formula $C_6H_{12}O_6$; **(b)** nitrous oxide, a substance used as an anesthetic and commonly called laughing gas, molecular formula $N_2O$.

**SOLUTION**

**(a)** The subscripts of an empirical formula are the smallest whole-number ratios. The smallest ratios are obtained by dividing each subscript by the largest common factor, in this case 6. The resultant empirical formula for glucose is $CH_2O$.

**(b)** Because the subscripts in $N_2O$ are already the lowest integral numbers, the empirical formula for nitrous oxide is the same as its molecular formula, $N_2O$.

**PRACTICE EXERCISE**

Give the empirical formula for *diborane*, whose molecular formula is $B_2H_6$.

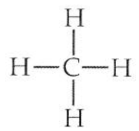
***Answer:*** $BH_3$

GO FIGURE

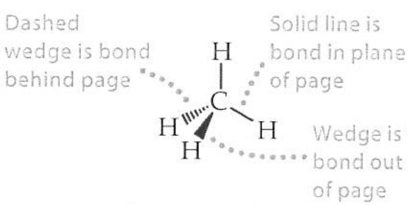
**What advantage does a ball-and-stick model have over a space-filling model?**

$CH_4$

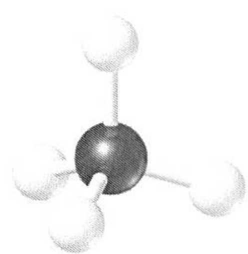
Molecular formula



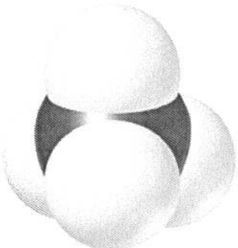
Structural formula




Perspective drawing



Ball-and-stick model

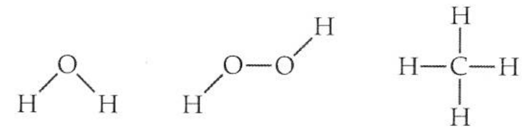


Space-filling model

▲ FIGURE 2.19 **Different representations of the methane ($CH_4$) molecule.** Structural formulas, perspective drawings, ball-and-stick models, and space-filling models correspond to the molecular formula, and each helps us visualize the ways atoms are attached to each other.

## Picturing Molecules

The molecular formula of a substance summarizes the composition of the substance but does not show how the atoms are joined together in the molecule. A **structural formula** shows which atoms are attached to which, as in the following examples:



Water Hydrogen peroxide Methane

The atoms are represented by their chemical symbols, and lines are used to represent the bonds that hold the atoms together.

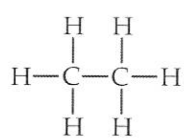
A structural formula usually does not depict the actual geometry of the molecule, that is, the actual angles at which atoms are joined together. A structural formula can be written as a *perspective drawing* (◀ FIGURE 2.19), however, to give some sense of three-dimensional shape.

Scientists also rely on various models to help visualize molecules. *Ball-and-stick models* show atoms as spheres and bonds as sticks. This type of model has the advantage of accurately representing the angles at which the atoms are attached to one another in the molecule (Figure 2.19). Sometimes the chemical symbols of the elements are superimposed on the balls, but often the atoms are identified simply by color.

A *space-filling model* depicts what the molecule would look like if the atoms were scaled up in size (Figure 2.19). These models show the relative sizes of the atoms, but the angles between atoms, which help define their molecular geometry, are often more difficult to see than in ball-and-stick models. As in ball-and-stick models, the identities of the atoms are indicated by color, but they may also be labeled with the element's symbol.

GIVE IT SOME THOUGHT

The structural formula for ethane is

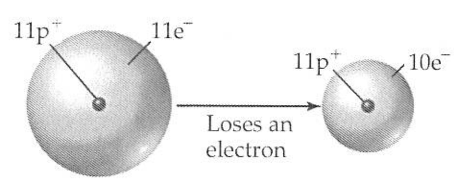


a. What is the molecular formula for ethane?
b. What is its empirical formula?
c. Which kind of molecular model would most clearly show the angles between atoms?

## 2.7 IONS AND IONIC COMPOUNDS

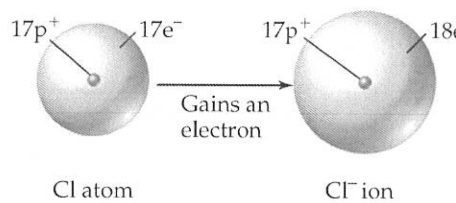
The nucleus of an atom is unchanged by chemical processes, but some atoms can readily gain or lose electrons. If electrons are removed from or added to an atom, a charged particle called an **ion** is formed. An ion with a positive charge is a **cation** (pronounced CAT-ion); a negatively charged ion is an **anion** (AN-ion).

To see how ions form, consider the sodium atom, which has 11 protons and 11 electrons. This atom easily loses one electron. The resulting cation has 11 protons and 10 electrons, which means it has a net charge of 1+.




Na atom Na$^+$ ion

The net charge on an ion is represented by a superscript. The superscripts +, 2+, and 3+, for instance, mean a net charge resulting from the *loss* of one, two, and three electrons, respectively. The superscripts −, 2−, and 3− represent net charges resulting from the *gain* of one, two, and three electrons, respectively. Chlorine, with 17 protons and 17 electrons, for example, can gain an electron in chemical reactions, producing the $Cl^-$ ion:




*In general, metal atoms tend to lose electrons to form cations and nonmetal atoms tend to gain electrons to form anions. Thus, ionic compounds tend to be composed of metals bonded with nonmetals, as in NaCl.*

**SAMPLE EXERCISE 2.7** **Writing Chemical Symbols for Ions**

Give the chemical symbol, including superscript indicating mass number, for **(a)** the ion with 22 protons, 26 neutrons, and 19 electrons; **(b)** the ion of sulfur that has 16 neutrons and 18 electrons.

**SOLUTION**

**(a)** The number of protons is the atomic number of the element. A periodic table or list of elements tells us that the element with atomic number 22 is titanium (Ti). The mass number (protons plus neutrons) of this isotope of titanium is $22 + 26 = 48$. Because the ion has three more protons than electrons, it has a net charge of 3+: $^{48}Ti^{3+}$.

**(b)** The periodic table tells us that sulfur (S) has an atomic number of 16. Thus, each atom or ion of sulfur contains 16 protons. We are told that the ion also has 16 neutrons, meaning the mass number is $16 + 16 = 32$. Because the ion has 16 protons and 18 electrons, its net charge is 2− and the ion symbol is $^{32}S^{2-}$.

In general, we will focus on the net charges of ions and ignore their mass numbers unless the circumstances dictate that we specify a certain isotope.

**PRACTICE EXERCISE**

How many protons, neutrons, and electrons does the $^{79}Se^{2-}$ ion possess?

***Answer:*** 34 protons, 45 neutrons, and 36 electrons

In addition to simple ions such as $Na^+$ and $Cl^-$, there are **polyatomic ions**, such as $NH_4^+$ (ammonium ion) and $SO_4^{2-}$ (sulfate ion). These latter ions consist of atoms joined as in a molecule, but they have a net positive or negative charge. We consider polyatomic ions in Section 2.8.

It is important to realize that the chemical properties of ions are very different from the chemical properties of the atoms from which the ions are derived. Although a given atom and its ion may be essentially the same (plus or minus a few electrons), the behavior of the ion is very different from that of its associated atom.

## Predicting Ionic Charges

Many atoms gain or lose electrons to end up with the same number of electrons as the noble gas closest to them in the periodic table. Noble-gas elements are chemically nonreactive and form very few compounds. We might deduce that this is because their electron arrangements are very stable. Nearby elements can obtain these same stable arrangements by losing or gaining electrons. For example, the loss of one electron from an atom of sodium leaves it with the same number of electrons as in a neon atom (10). Similarly, when chlorine gains an electron, it ends up with 18, the same number of electrons as in argon. We will use this simple observation to explain the formation of ions until Chapter 8, where we discuss chemical bonding.

**SAMPLE EXERCISE 2.8** **Predicting Ionic Charge**

Predict the charge expected for the most stable ion of barium and the most stable ion of oxygen.

**SOLUTION**

We will assume that these elements form ions that have the same number of electrons as the nearest noble-gas atom. From the periodic table, we see that barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can attain a stable arrangement of 54 electrons by losing two electrons, forming the $Ba^{2+}$ cation.

Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can attain this stable electron arrangement by gaining two electrons, forming the $O^{2-}$ anion.

**PRACTICE EXERCISE**

Predict the charge expected for the most stable ion of **(a)** aluminum and **(b)** fluorine.

***Answer:*** (**a**) 3+, (**b**) 1−

The periodic table is very useful for remembering ionic charges, especially those of elements on the left and right sides of the table. As ▼ FIGURE 2.20 shows, the charges of these ions relate in a simple way to their positions in the table: The group 1A elements (alkali metals) form 1+ ions, the group 2A elements (alkaline earths) form 2+ ions, the group 7A elements (halogens) form 1− ions, and the group 6A elements form 2− ions. (Many of the other groups do not lend themselves to such simple rules.)

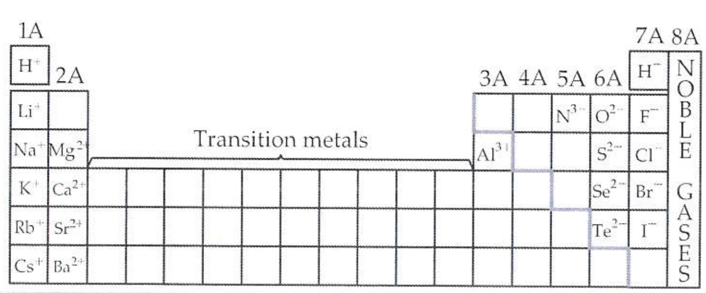
## Ionic Compounds

A great deal of chemical activity involves the transfer of electrons from one substance to another. ▶ FIGURE 2.21 shows that when elemental sodium is allowed to react with elemental chlorine, an electron transfers from a sodium atom to a chlorine atom, forming a $Na^+$ ion and a $Cl^-$ ion. Because objects of opposite charge attract, the $Na^+$ and the $Cl^-$ ions bind together to form the compound sodium chloride (NaCl). Sodium chloride, which we know better as common table salt, is an example of an **ionic compound**, a compound made up of cations and anions.

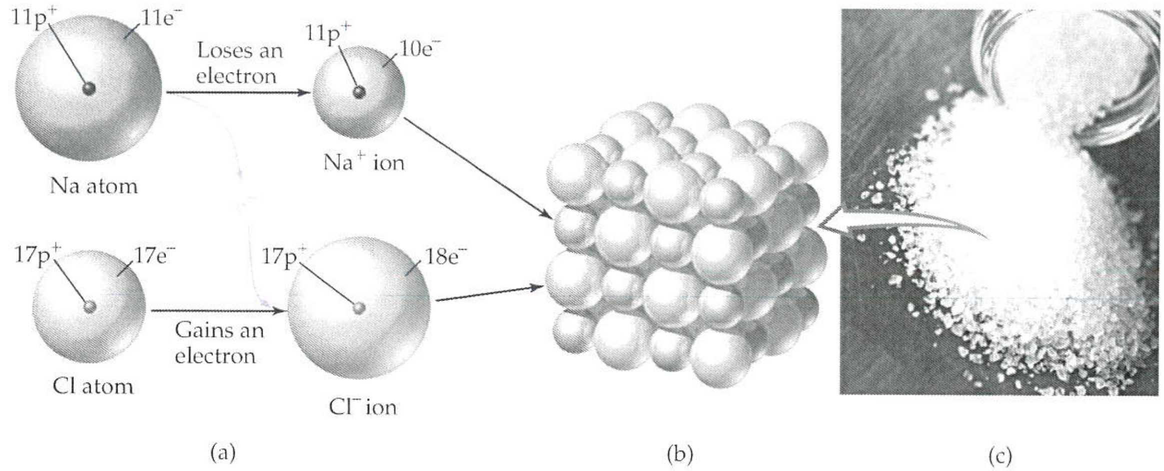
We can often tell whether a compound is ionic (consisting of ions) or molecular (consisting of molecules) from its composition. In general, cations are metal ions and anions are nonmetal ions. Consequently, *ionic compounds are generally combinations of metals and nonmetals*, as in NaCl. In contrast, *molecular compounds are generally composed of nonmetals only*, as in $H_2O$.

GO FIGURE

The most common ions for silver, zinc, and scandium are $Ag^+$, $Zn^{2+}$, and $Sc^{3+}$. Locate the boxes in which you would place these ions in this table. Which of these ions have the same number of electrons as a noble-gas element?




▶ FIGURE 2.20 **Predictable charges of some common ions.** Notice that the red stepped line that divides metals from nonmetals also separates cations from anions. Hydrogen forms both 1+ and 1− ions.




▲ FIGURE 2.21 **Formation of an ionic compound.** (a) The transfer of an electron from a Na atom to a Cl atom leads to the formation of a $Na^+$ ion and a $Cl^-$ ion. (b) Arrangement of these ions in solid sodium chloride, NaCl. (c) A sample of sodium chloride crystals.

## SAMPLE EXERCISE 2.9 Identifying Ionic and Molecular Compounds

Which of these compounds would you expect to be ionic: $N_2O$, $Na_2O$, $CaCl_2$, $SF_4$?

### SOLUTION

We predict that $Na_2O$ and $CaCl_2$ are ionic compounds because they are composed of a metal combined with a nonmetal. We predict (correctly) that $N_2O$ and $SF_4$ are molecular compounds because they are composed entirely of nonmetals.

### PRACTICE EXERCISE

Which of these compounds are molecular: $CBr_4$, FeS, $P_4O_6$, $PbF_2$?

***Answer:*** $CBr_4$ and $P_4O_6$

The ions in ionic compounds are arranged in three-dimensional structures, as Figure 2.21(b) shows for NaCl. Because there is no discrete "molecule" of NaCl, we are able to write only an empirical formula for this substance. This is true for most other ionic compounds.

We can write the empirical formula for an ionic compound if we know the charges of the ions. This is true because chemical compounds are always electrically neutral. Consequently, the ions in an ionic compound always occur in such a ratio that the total positive charge equals the total negative charge. Thus, there is one $Na^+$ to one $Cl^-$ (giving NaCl), one $Ba^{2+}$ to two $Cl^-$ (giving $BaCl_2$), and so forth.

As you consider these and other examples, you will see that if the charges on the cation and anion are equal, the subscript on each ion is 1. If the charges are not equal, the charge on one ion (without its sign) will become the subscript on the other ion. For example, the ionic compound formed from Mg (which forms $Mg^{2+}$ ions) and N (which forms $N^{3-}$ ions) is $Mg_3N_2$:

$$\mathrm{Mg}^{2+} \quad \mathrm{N}^{3-} \longrightarrow \mathrm{Mg}_3\mathrm{N}_2$$

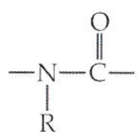
### GIVE IT SOME THOUGHT

Why don't we write the formula for the compound formed by $Ca^{2+}$ and $O^{2-}$ as $Ca_2O_2$?

## CHEMISTRY AND LIFE

### ELEMENTS REQUIRED BY LIVING ORGANISMS

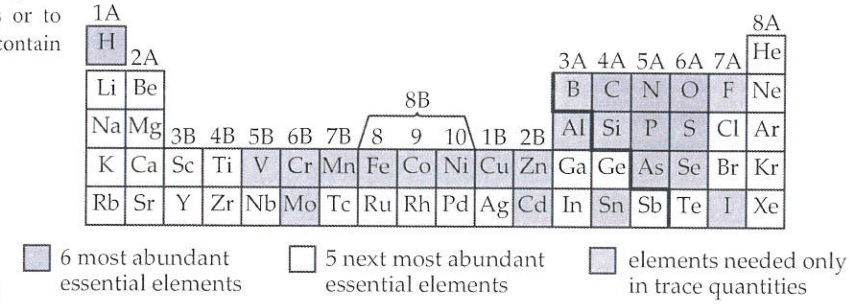
The colored regions of ▶ FIGURE 2.22 shows the elements essential to life. More than 97% of the mass of most organisms is made up of just six of these elements—oxygen, carbon, hydrogen, nitrogen, phosphorus, and sulfur. Water is the most common compound in living organisms, accounting for at least 70% of the mass of most cells. In the solid components of cells, carbon is the most prevalent element by mass. Carbon atoms are found in a vast variety of organic molecules, bonded either to other carbon atoms or to atoms of other elements. All proteins, for example, contain the group



which occurs repeatedly in the molecules. (R is either an H atom or a combination of atoms, such as $CH_3$.)

In addition, 23 more elements have been found in various living organisms. Five are ions required by all organisms: $Ca^{2+}$, $Cl^-$, $Mg^{2+}$, $K^+$, and $Na^+$. Calcium ions, for example, are necessary for the formation of bone and transmission of nervous system signals. Many other elements are needed in only very small quantities and consequently are called *trace* elements. For example, trace quantities of copper are required in the diet of humans to aid in the synthesis of hemoglobin.

*RELATED EXERCISE: 2.96*




▲ FIGURE 2.22 **Elements essential to life.**

## STRATEGIES IN CHEMISTRY

### PATTERN RECOGNITION

Someone once said that drinking at the fountain of knowledge in a chemistry course is like drinking from a fire hydrant. Indeed, the pace can sometimes seem brisk. More to the point, however, we can drown in the facts if we do not see the general patterns. The value of recognizing patterns and learning rules and generalizations is that they free us from having to learn (or trying to memorize) many individual facts. The patterns, rules, and generalizations tie ideas together so that we do not get lost in the details.

Many students struggle with chemistry because they do not see how different topics relate to one another so they treat every idea and problem as being unique instead of as an example or application of a general rule, procedure, or relationship. You can avoid this pitfall by remembering the following.

Notice the structure of the topic you are studying. Pay attention to trends and rules given to summarize a large body of information. Notice, for example, how atomic structure helps us understand the existence of isotopes (as Table 2.2 shows) and how the periodic table helps us remember ionic charges (as Figure 2.20 shows).

You may surprise yourself by observing patterns that are not explicitly spelled out yet. Perhaps you have noticed certain trends in chemical formulas, for instance. Moving across the periodic table from element 11 (Na), we find that the elements form compounds with F having the following compositions: NaF, $MgF_2$, and $AlF_3$. Does this trend continue? Do $SiF_4$, $PF_5$, and $SF_6$ exist? Indeed they do. If you have noticed trends like this from the scraps of information you have seen so far, then you are ahead of the game and have prepared yourself for some topics we will address in later chapters.

**SAMPLE EXERCISE 2.10** **Using Ionic Charge to Write Empirical Formulas for Ionic Compounds**

Write the empirical formula of the compound formed by **(a)** $Al^{3+}$ and $Cl^-$ ions, **(b)** $Al^{3+}$ and $O^{2-}$ ions, **(c)** $Mg^{2+}$ and $NO_3^-$ ions.

**SOLUTION**

**(a)** Three $Cl^-$ ions are required to balance the charge of one $Al^{3+}$ ion, making the formula $AlCl_3$.

**(b)** Two $Al^{3+}$ ions are required to balance the charge of three $O^{2-}$ ions. That is, the total positive charge is 6+, and the total negative charge is 6−. The formula is $Al_2O_3$.

(c) Two $NO_3^-$ ions are needed to balance the charge of one $Mg^{2+}$, yielding $Mg(NO_3)_2$. Note that the formula for the polyatomic ion, $NO_3^-$, must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion.

**PRACTICE EXERCISE**

Write the empirical formula for the compound formed by (a) $Na^+$ and $PO_4^{3-}$, (b) $Zn^{2+}$ and $SO_4^{2-}$, (c) $Fe^{3+}$ and $CO_3^{2-}$.

***Answers:*** (a) $Na_3PO_4$, (b) $ZnSO_4$, (c) $Fe_2(CO_3)_3$

# 2.8 NAMING INORGANIC COMPOUNDS

The names and chemical formulas of compounds are essential vocabulary in chemistry. The system used in naming substances is called **chemical nomenclature**, from the Latin words *nomen* (name) and *calare* (to call).

There are more than 50 million known chemical substances. Naming them all would be a hopelessly complicated task if each had a name independent of all others. Many important substances that have been known for a long time, such as water ($H_2O$) and ammonia ($NH_3$), do have traditional names (called *common names*). For most substances, however, we rely on a set of rules that leads to an informative and unique name for each substance, a name based on the composition of the substance.

The rules for chemical nomenclature are based on the division of substances into categories. The major division is between organic and inorganic compounds. *Organic compounds* contain carbon and hydrogen, often in combination with oxygen, nitrogen, or other elements. All others are *inorganic compounds*. Early chemists associated organic compounds with plants and animals and inorganic compounds with the nonliving portion of our world. Although this distinction is no longer pertinent, the classification between organic and inorganic compounds continues to be useful. In this section we consider the basic rules for naming three categories of inorganic compounds: ionic compounds, molecular compounds, and acids.

## Names and Formulas of Ionic Compounds

Recall from Section 2.7 that ionic compounds usually consist of metal ions combined with nonmetal ions. The metals form the cations, and the nonmetals form the anions.

1. Cations
   a. *Cations formed from metal atoms have the same name as the metal:*

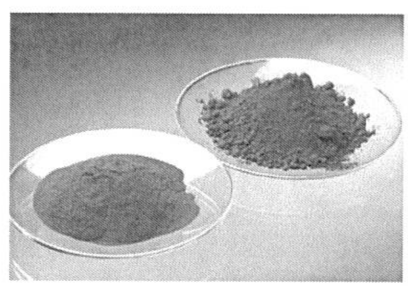
      $Na^+$ sodium ion    $Zn^{2+}$ zinc ion    $Al^{3+}$ aluminum ion

   b. *If a metal can form cations with different charges, the positive charge is indicated by a Roman numeral in parentheses following the name of the metal:*

| | | | |
|---|---|---|---|
| $Fe^{2+}$ | iron(II) ion | $Cu^+$ | copper(I) ion |
| $Fe^{3+}$ | iron(III) ion | $Cu^{2+}$ | copper(II) ion |

Ions of the same element that have different charges have different properties, such as different colors (▶ FIGURE 2.23).

Most metals that form cations with different charges are *transition metals*, elements that occur in the middle of the periodic table, from group 3B to group 2B. The metals that form only one cation (only one possible charge) are those of group 1A and group 2A, as well as $Al^{3+}$ (group 3A) and two transition-metal ions: $Ag^+$ (group 1B) and $Zn^{2+}$ (group 2B). Charges are not expressed when naming these ions. However, if there is any doubt in your mind whether a metal forms more than one cation, use a Roman numeral to indicate the charge. It is never wrong to do so, even though it may be unnecessary.



▲ FIGURE 2.23 **Different ions of the same element have different properties.** Both substances shown are compounds of iron. The substance on the left is $Fe_3O_4$, which contains $Fe^{2+}$ and $Fe^{3+}$ ions. The substance on the right is $Fe_2O_3$, which contains $Fe^{3+}$ ions.

An older method still widely used for distinguishing between differently charged ions of a metal uses the endings *-ous* and *-ic* added to the root of the element's Latin name:

| | | | |
|---|---|---|---|
| $Fe^{2+}$ | ferrous ion | $Cu^{+}$ | cuprous ion |
| $Fe^{3+}$ | ferric ion | $Cu^{2+}$ | cupric ion |

Although we will only rarely use these older names in this text, you might encounter them elsewhere.

c. *Cations formed from nonmetal atoms have names that end in* -ium:

$NH_4^+$ ammonium ion $\quad$ $H_3O^+$ hydronium ion

These two ions are the only ions of this kind that we will encounter frequently in the text.

The names and formulas of some common cations are shown in ▼ TABLE 2.4 and on the back inside cover of the text. The ions on the left side in Table 2.4 are the monatomic ions that do not have more than one possible charge. Those on the right side are either polyatomic cations or cations with more than one possible charge. The $Hg_2^{2+}$ ion is unusual because, even though it is a metal ion, it is not monatomic. It is called the mercury(I) ion because it can be thought of as two $Hg^+$ ions bound together. The cations that you will encounter most frequently are shown in boldface. You should learn these cations first.

GIVE IT SOME THOUGHT

a. Why is CrO named using a Roman numeral, chromium(II) oxide, whereas CaO is named without a Roman numeral, calcium oxide?
b. What does the *-ium* ending on the name ammonium ion tell you about the composition of the ion?

**TABLE 2.4 • Common Cations***

| Charge | Formula | Name | Formula | Name |
|---|---|---|---|---|
| 1+ | **$H^+$** | **hydrogen ion** | **$NH_4^+$** | **ammonium ion** |
| | $Li^+$ | lithium ion | $Cu^+$ | copper(I) or cuprous ion |
| | **$Na^+$** | **sodium ion** | | |
| | **$K^+$** | **potassium ion** | | |
| | $Cs^+$ | cesium ion | | |
| | **$Ag^+$** | **silver ion** | | |
| 2+ | **$Mg^{2+}$** | **magnesium ion** | $Co^{2+}$ | cobalt(II) or cobaltous ion |
| | **$Ca^{2+}$** | **calcium ion** | **$Cu^{2+}$** | **copper(II)** or cupric ion |
| | $Sr^{2+}$ | strontium ion | **$Fe^{2+}$** | **iron(II)** or ferrous ion |
| | $Ba^{2+}$ | barium ion | $Mn^{2+}$ | manganese(II) or manganous ion |
| | **$Zn^{2+}$** | **zinc ion** | $Hg_2^{2+}$ | mercury(I) or mercurous ion |
| | $Cd^{2+}$ | cadmium ion | **$Hg^{2+}$** | **mercury(II)** or mercuric ion |
| | | | $Ni^{2+}$ | nickel(II) or nickelous ion |
| | | | **$Pb^{2+}$** | **lead(II)** or plumbous ion |
| | | | $Sn^{2+}$ | tin(II) or stannous ion |
| 3+ | **$Al^{3+}$** | **aluminum ion** | $Cr^{3+}$ | chromium(III) or chromic ion |
| | | | **$Fe^{3+}$** | **iron(III)** or ferric ion |

*The ions we use most often in this course are in boldface. Learn them first.

2. Anions

a. *The names of monatomic anions are formed by replacing the ending of the name of the element with* -ide:

| | | | | | |
|---|---|---|---|---|---|
| $H^-$ | hydride ion | $O^{2-}$ | oxide ion | $N^{3-}$ | nitride ion |

A few polyatomic anions also have names ending in *-ide*:

| | | | | | |
|---|---|---|---|---|---|
| $OH^-$ | hydroxide ion | $CN^-$ | cyanide ion | $O_2^{2-}$ | peroxide ion |

b. *Polyatomic anions containing oxygen have names ending in either* -ate *or* -ite and are called **oxyanions**. The *-ate* is used for the most common or representative oxyanion of an element, and *-ite* is used for an oxyanion that has the same charge but one O atom fewer:

| | | | |
|---|---|---|---|
| $NO_3^-$ | nitrate ion | $SO_4^{2-}$ | sulfate ion |
| $NO_2^-$ | nitrite ion | $SO_3^{2-}$ | sulfite ion |

Prefixes are used when the series of oxyanions of an element extends to four members, as with the halogens. The prefix *per-* indicates one more O atom than the oxyanion ending in *-ate*; *hypo-* indicates one O atom fewer than the oxyanion ending in *-ite*:

| | |
|---|---|
| $ClO_4^-$ | perchlorate ion (one more O atom than chlorate) |
| $\mathbf{ClO_3^-}$ | **chlorate ion** |
| $ClO_2^-$ | chlorite ion (one O atom fewer than chlorate) |
| $ClO^-$ | hypochlorite ion (one O atom fewer than chlorite) |

These rules are summarized in ▼ FIGURE 2.24.

GIVE IT SOME THOUGHT

What information is conveyed by the endings *-ide, -ate,* and *-ite* in the name of an anion?

▶ FIGURE 2.25 can help you remember the charge and number of oxygen atoms in the various oxyanions. Notice that C and N, both period 2 elements, have only three O atoms each, whereas the period 3 elements P, S, and Cl have four O atoms each. Beginning at the lower right in Figure 2.25, note that ionic charge increases from right to left, from 1− for $ClO_4^-$ to 3− for $PO_4^{3-}$. In the second period the charges also increase from right to left, from 1− for $NO_3^-$ to 2− for $CO_3^{2-}$. Notice also that although each of the anions in Figure 2.25 ends in *-ate*, the $ClO_4^-$ ion also has a *per-* prefix.

GO FIGURE

Name the anion obtained by removing one oxygen atom from the perbromate ion, $BrO_4^-$.

Simple anion: ________ide (chloride, $Cl^-$)

Oxyanions: per______ate (perchlorate, $ClO_4^-$) $\xleftarrow{+\text{O atom}}$ ________ate (chlorate, $ClO_3^-$) $\xrightarrow{-\text{O atom}}$ ________ite (chlorite, $ClO_2^-$) $\xrightarrow{-\text{O atom}}$ hypo____ite (hypochlorite, $ClO^-$)

Common or representative oxyanion

▲ FIGURE 2.24 **Procedure for naming anions.** The first part of the element's name, such as "chlor" for chlorine or "sulf" for sulfur, goes in the blank.

Maximum of 3 O atoms in period 2.

| | Group 4A | Group 5A | Group 6A | Group 7A |
|---|---|---|---|---|
| Period 2 | $CO_3^{2-}$ Carbonate ion | $NO_3^-$ Nitrate ion | | |
| Period 3 | | $PO_4^{3-}$ Phosphate ion | $SO_4^{2-}$ Sulfate ion | $ClO_4^-$ Perchlorate ion |

Maximum of 4 O atoms in period 3.

Charges increase right to left.

▲ FIGURE 2.25 **Common oxyanions.** The composition and charges of common oxyanions are related to their location in the periodic table.

GIVE IT SOME THOUGHT

Predict the formulas for the borate ion and silicate ion, assuming they contain a single B and Si atom, respectively, and follow the trends shown in Figure 2.25.

**SAMPLE EXERCISE 2.11 Determining the Formula of an Oxyanion from Its Name**

Based on the formula for the sulfate ion, predict the formula for **(a)** the selenate ion and **(b)** the selenite ion. (Sulfur and selenium are both in group 6A and form analogous oxyanions.)

**SOLUTION**

**(a)** The sulfate ion is $SO_4^{2-}$. The analogous selenate ion is therefore $SeO_4^{2-}$.
**(b)** The ending *-ite* indicates an oxyanion with the same charge but one O atom fewer than the corresponding oxyanion that ends in *-ate*. Thus, the formula for the selenite ion is $SeO_3^{2-}$.

**PRACTICE EXERCISE**

The formula for the bromate ion is analogous to that for the chlorate ion. Write the formula for the hypobromite and bromite ions.
***Answer:*** $BrO^-$ and $BrO_2^-$

**c.** *Anions derived by adding $H^+$ to an oxyanion are named by adding as a prefix the word* hydrogen *or* dihydrogen, *as appropriate:*

| | | | |
|---|---|---|---|
| $CO_3^{2-}$ | carbonate ion | $PO_4^{3-}$ | phosphate ion |
| $HCO_3^-$ | hydrogen carbonate ion | $H_2PO_4^-$ | dihydrogen phosphate ion |

Notice that each $H^+$ added reduces the negative charge of the parent anion by one. An older method for naming some of these ions uses the prefix *bi-*. Thus, the $HCO_3^-$ ion is commonly called the bicarbonate ion, and $HSO_4^-$ is sometimes called the bisulfate ion.

The names and formulas of the common anions are listed in ▶ TABLE 2.5 and on the back inside cover of the text. Those anions whose names end in *-ide* are listed on the left portion of Table 2.5, and those whose names end in *-ate* are listed on the right. The most common of these ions are shown in boldface. You should learn names and formulas of these anions first. The formulas of the ions whose names end with *-ite* can be derived from those ending in *-ate* by removing an O atom. Notice the location of the monatomic ions in the periodic table. Those of group 7A always have a 1− charge ($F^-$, $Cl^-$, $Br^-$, and $I^-$), and those of group 6A have a 2− charge ($O^{2-}$ and $S^{2-}$).

3. Ionic Compounds

*Names of ionic compounds consist of the cation name followed by the anion name:*

| | |
|---|---|
| $CaCl_2$ | calcium chloride |
| $Al(NO_3)_3$ | aluminum nitrate |
| $Cu(ClO_4)_2$ | copper(II) perchlorate (or cupric perchlorate) |

In the chemical formulas for aluminum nitrate and copper(II) perchlorate, parentheses followed by the appropriate subscript are used because the compounds contain two or more polyatomic ions.

## SAMPLE EXERCISE 2.12 Determining the Names of Ionic Compounds from Their Formulas

Name the ionic compounds (a) $K_2SO_4$, (b) $Ba(OH)_2$, (c) $FeCl_3$.

### SOLUTION

In naming ionic compounds, it is important to recognize polyatomic ions and to determine the charge of cations with variable charge.

(a) The cation is $K^+$, the potassium ion, and the anion is $SO_4^{2-}$, the sulfate ion, making the name potassium sulfate. (If you thought the compound contained $S^{2-}$ and $O^{2-}$ ions, you failed to recognize the polyatomic sulfate ion.)

(b) The cation is $Ba^{2+}$, the barium ion, and the anion is $OH^-$, the hydroxide ion: barium hydroxide.

(c) You must determine the charge of Fe in this compound because an iron atom can form more than one cation. Because the compound contains three chloride ions, $Cl^-$, the cation must be $Fe^{3+}$, the iron(III), or ferric, ion. Thus, the compound is iron(III) chloride or ferric chloride.

### PRACTICE EXERCISE

Name the ionic compounds (a) $NH_4Br$, (b) $Cr_2O_3$, (c) $Co(NO_3)_2$.

***Answers:*** (a) ammonium bromide, (b) chromium(III) oxide, (c) cobalt(II) nitrate

**TABLE 2.5 • Common Anions***

| Charge | Formula | Name | Formula | Name |
|---|---|---|---|---|
| 1− | $H^-$ | hydride ion | $CH_3COO^-$ (or $C_2H_3O_2^-$) | **acetate ion** |
| | $F^-$ | **fluoride ion** | $ClO_3^-$ | chlorate ion |
| | $Cl^-$ | **chloride ion** | $ClO_4^-$ | **perchlorate ion** |
| | $Br^-$ | **bromide ion** | $NO_3^-$ | **nitrate ion** |
| | $I^-$ | **iodide ion** | $MnO_4^-$ | permanganate ion |
| | $CN^-$ | cyanide ion | | |
| | $OH^-$ | **hydroxide ion** | | |
| 2− | $O^{2-}$ | **oxide ion** | $CO_3^{2-}$ | **carbonate ion** |
| | $O_2^{2-}$ | peroxide ion | $CrO_4^{2-}$ | chromate ion |
| | $S^{2-}$ | **sulfide ion** | $Cr_2O_7^{2-}$ | dichromate ion |
| | | | $SO_4^{2-}$ | **sulfate ion** |
| 3− | $N^{3-}$ | nitride ion | $PO_4^{3-}$ | **phosphate ion** |

*The ions we use most often are in boldface. Learn them first.

## SAMPLE EXERCISE 2.13 Determining the Formulas of Ionic Compounds from Their Names

Write the chemical formulas for (a) potassium sulfide, (b) calcium hydrogen carbonate, (c) nickel(II) perchlorate.

### SOLUTION

In going from the name of an ionic compound to its chemical formula, you must know the charges of the ions to determine the subscripts.

(a) The potassium ion is $K^+$, and the sulfide ion is $S^{2-}$. Because ionic compounds are electrically neutral, two $K^+$ ions are required to balance the charge of one $S^{2-}$ ion, giving $K_2S$ for the empirical formula.

(**b**) The calcium ion is $Ca^{2+}$. The carbonate ion is $CO_3^{2-}$, so the hydrogen carbonate ion is $HCO_3^-$. Two $HCO_3^-$ ions are needed to balance the positive charge of $Ca^{2+}$, giving $Ca(HCO_3)_2$.

(**c**) The nickel(II) ion is $Ni^{2+}$. The perchlorate ion is $ClO_4^-$. Two $ClO_4^-$ ions are required to balance the charge on one $Ni^{2+}$ ion, giving $Ni(ClO_4)_2$.

**PRACTICE EXERCISE**

Give the chemical formulas for (**a**) magnesium sulfate, (**b**) silver sulfide, (**c**) lead(II) nitrate.

***Answers:*** (**a**) $MgSO_4$, (**b**) $Ag_2S$, (**c**) $Pb(NO_3)_2$

## Names and Formulas of Acids

Acids are an important class of hydrogen-containing compounds, and they are named in a special way. For our present purposes, an *acid* is a substance whose molecules yield hydrogen ions ($H^+$) when dissolved in water. When we encounter the chemical formula for an acid at this stage of the course, it will be written with H as the first element, as in HCl and $H_2SO_4$.

An acid is composed of an anion connected to enough $H^+$ ions to neutralize, or balance, the anion's charge. Thus, the $SO_4^{2-}$ ion requires two $H^+$ ions, forming $H_2SO_4$. The name of an acid is related to the name of its anion, as summarized in ▼ FIGURE 2.26.

1. *Acids containing anions whose names end in* -ide *are named by changing the* -ide *ending to* -ic, *adding the prefix* hydro- *to this anion name, and then following with the word* acid:

| Anion | Corresponding Acid |
|---|---|
| $Cl^-$ (chloride) | HCl (hydrochloric acid) |
| $S^{2-}$ (sulfide) | $H_2S$ (hydrosulfuric acid) |

2. *Acids containing anions whose names end in* -ate *or* -ite *are named by changing* -ate *to* -ic *and* -ite *to* -ous *and then adding the word* acid. Prefixes in the anion name are retained in the name of the acid:

| Anion | | Corresponding Acid | |
|---|---|---|---|
| $ClO_4^-$ | (perchlorate) | $HClO_4$ | (perchloric acid) |
| $ClO_3^-$ | (chlorate) | $HClO_3$ | (chloric acid) |
| $ClO_2^-$ | (chlorite) | $HClO_2$ | (chlorous acid) |
| $ClO^-$ | (hypochlorite) | HClO | (hypochlorous acid) |

| Anion | | Acid |
|---|---|---|
| ________ide<br>(chloride, $Cl^-$) | $\xrightarrow[\text{ions}]{\text{add } H^+}$ | hydro______ic acid<br>(hydrochloric acid, HCl) |
| ________ate<br>(chlorate, $ClO_3^-$)<br>(perchlorate, $ClO_4^-$) | $\xrightarrow[\text{ions}]{\text{add } H^+}$ | ________ic acid<br>(chloric acid, $HClO_3$)<br>(perchloric acid, $HClO_4$) |
| ________ite<br>(chlorite, $ClO_2^-$)<br>(hypochlorite, $ClO^-$) | $\xrightarrow[\text{ions}]{\text{add } H^+}$ | ________ous acid<br>(chlorous acid, $HClO_2$)<br>(hypochlorous acid, HClO) |

▶ FIGURE 2.26 **How anion names and acid names relate.** The prefixes *per-* and *hypo-* are retained in going from the anion to the acid.

GIVE IT SOME THOUGHT

Name the acid obtained by adding $H^+$ to the iodate ion, $IO_3^-$.

**SAMPLE EXERCISE 2.14** **Relating the Names and Formulas of Acids**

Name the acids (a) HCN, (b) $HNO_3$, (c) $H_2SO_4$, (d) $H_2SO_3$.

SOLUTION

(a) The anion from which this acid is derived is $CN^-$, the cyanide ion. Because this ion has an *-ide* ending, the acid is given a *hydro-* prefix and an *-ic* ending: hydrocyanic acid. Only water solutions of HCN are referred to as hydrocyanic acid. The pure compound, which is a gas under normal conditions, is called hydrogen cyanide. Both hydrocyanic acid and hydrogen cyanide are *extremely* toxic.

(b) Because $NO_3^-$ is the nitrate ion, $HNO_3$ is called nitric acid (the *-ate* ending of the anion is replaced with an *-ic* ending in naming the acid).

(c) Because $SO_4^{2-}$ is the sulfate ion, $H_2SO_4$ is called sulfuric acid.

(d) Because $SO_3^{2-}$ is the sulfite ion, $H_2SO_3$ is sulfurous acid (the *-ite* ending of the anion is replaced with an *-ous* ending).

PRACTICE EXERCISE

Give the chemical formulas for (a) hydrobromic acid, (b) carbonic acid.

***Answers:*** (a) HBr, (b) $H_2CO_3$

## Names and Formulas of Binary Molecular Compounds

The procedures used for naming *binary* (two-element) molecular compounds are similar to those used for naming ionic compounds:

1. *The name of the element farther to the left in the periodic table (closest to the metals) is usually written first.* An exception occurs when the compound contains oxygen and chlorine, bromine, or iodine (any halogen except fluorine), in which case oxygen is written last.
2. *If both elements are in the same group, the lower one is named first.*
3. *The name of the second element is given an* -ide *ending.*
4. *Greek prefixes* (▶ TABLE 2.6) are used to indicate the number of atoms of each element. The prefix *mono-* is never used with the first element. When the prefix ends in *a* or *o* and the name of the second element begins with a vowel, the *a* or *o* of the prefix is often dropped.

The following examples illustrate these rules:

| | | | |
|---|---|---|---|
| $Cl_2O$ | dichlorine monoxide | $NF_3$ | nitrogen trifluoride |
| $N_2O_4$ | dinitrogen tetroxide | $P_4S_{10}$ | tetraphosphorus decasulfide |

Rule 4 is necessary because we cannot predict formulas for most molecular substances the way we can for ionic compounds. Molecular compounds that contain hydrogen and one other element are an important exception, however. These compounds can be treated as if they were neutral substances containing $H^+$ ions and anions. Thus, you can predict that the substance named hydrogen chloride has the formula HCl, containing one $H^+$ to balance the charge of one $Cl^-$. (The name hydrogen chloride is used only for the pure compound; water solutions of HCl are called hydrochloric acid.) Similarly, the formula for hydrogen sulfide is $H_2S$ because two $H^+$ are needed to balance the charge on $S^{2-}$.

**TABLE 2.6 • Prefixes Used in Naming Binary Compounds Formed between Nonmetals**

| Prefix | Meaning |
|---|---|
| *Mono-* | 1 |
| *Di-* | 2 |
| *Tri-* | 3 |
| *Tetra-* | 4 |
| *Penta-* | 5 |
| *Hexa-* | 6 |
| *Hepta-* | 7 |
| *Octa-* | 8 |
| *Nona-* | 9 |
| *Deca-* | 10 |

**SAMPLE EXERCISE 2.15** **Relating the Names and Formulas of Binary Molecular Compounds**

Name the compounds (**a**) $SO_2$, (**b**) $PCl_5$, (**c**) $Cl_2O_3$.

**SOLUTION**

The compounds consist entirely of nonmetals, so they are molecular rather than ionic. Using the prefixes in Table 2.6, we have (**a**) sulfur dioxide, (**b**) phosphorus pentachloride, and (**c**) dichlorine trioxide.

**PRACTICE EXERCISE**

Give the chemical formulas for (**a**) silicon tetrabromide, (**b**) disulfur dichloride.

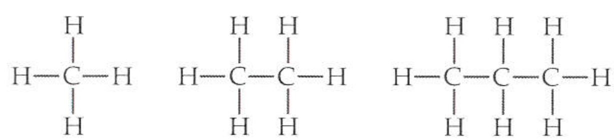
***Answers:*** (**a**) $SiBr_4$, (**b**) $S_2Cl_2$

## 2.9 SOME SIMPLE ORGANIC COMPOUNDS

The study of compounds of carbon is called **organic chemistry**, and as noted earlier, compounds that contain carbon and hydrogen, often in combination with oxygen, nitrogen, or other elements, are called *organic compounds*. We will examine organic compounds in Chapter 24, but here we present a brief introduction to some of the simplest organic compounds.

### Alkanes

Compounds that contain only carbon and hydrogen are called **hydrocarbons**. In the simplest class of hydrocarbons, **alkanes**, each carbon is bonded to four other atoms. The three smallest alkanes are methane ($CH_4$), ethane ($C_2H_6$), and propane ($C_3H_8$). The structural formulas of these three alkanes are as follows:

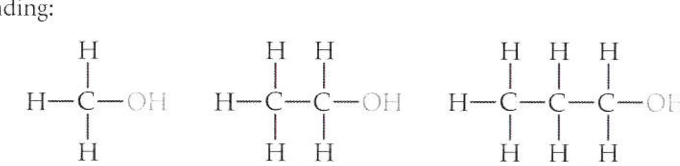


Methane Ethane Propane

Although hydrocarbons are binary molecular compounds, they are not named like the binary inorganic compounds discussed in Section 2.8. Instead, each alkane has a name that ends in *-ane*. The alkane with four carbons is called *butane*. For alkanes with five or more carbons, the names are derived from prefixes like those in Table 2.6. An alkane with eight carbon atoms, for example, is *octane* ($C_8H_{18}$), where the *octa-* prefix for eight is combined with the *-ane* ending for an alkane.

### Some Derivatives of Alkanes

Other classes of organic compounds are obtained when one or more hydrogen atoms in an alkane are replaced with *functional groups*, which are specific groups of atoms. An **alcohol**, for example, is obtained by replacing an H atom of an alkane with an —OH group. The name of the alcohol is derived from that of the alkane by adding an *-ol* ending:

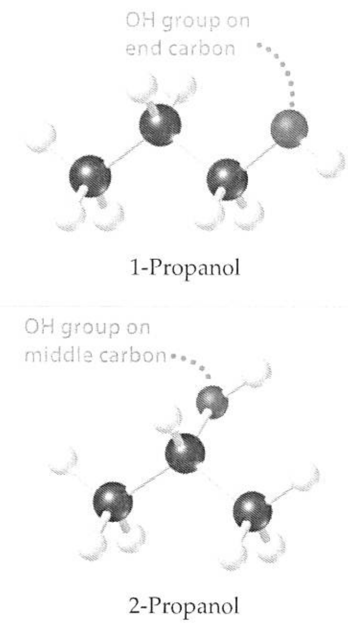


Methanol Ethanol 1-Propanol

Alcohols have properties that are very different from the properties of the alkanes from which the alcohols are obtained. For example, methane, ethane, and propane are all colorless gases under normal conditions, whereas methanol, ethanol, and propanol are colorless liquids. We will discuss the reasons for these differences in Chapter 11.

The prefix "1" in the name 1-propanol indicates that the replacement of H with OH has occurred at one of the "outer" carbon atoms rather than the "middle" carbon atom. A different compound, called either 2-propanol or isopropyl alcohol, is obtained when the OH functional group is attached to the middle carbon atom (▶ FIGURE 2.27).

Compounds with the same molecular formula but different arrangements of atoms are called **isomers**. There are many different kinds of isomers, as we will discover later in this book. What we have here with 1-propanol and 2-propanol are *structural isomers*, compounds having the same molecular formula but different structural formulas.



▲ FIGURE 2.27 **The two forms (isomers) of propanol.**

### GIVE IT SOME THOUGHT

Draw the structural formulas of the two isomers of butane, $C_4H_{10}$.

Much of the richness of organic chemistry is possible because organic compounds can form long chains of carbon–carbon bonds. The series of alkanes that begins with methane, ethane, and propane and the series of alcohols that begins with methanol, ethanol, and propanol can both be extended for as long as we desire, in principle. The properties of alkanes and alcohols change as the chains get longer. Octanes, which are alkanes with eight carbon atoms, are liquids under normal conditions. If the alkane series is extended to tens of thousands of carbon atoms, we obtain *polyethylene*, a solid substance that is used to make thousands of plastic products, such as plastic bags, food containers, and laboratory equipment.

### SAMPLE EXERCISE 2.16 Writing Structural and Molecular Formulas for Hydrocarbons

Assuming the carbon atoms in *pentane* are in a linear chain, write (**a**) the structural formula and (**b**) the molecular formula for this alkane.

#### SOLUTION

(**a**) Alkanes contain only carbon and hydrogen, and each carbon is attached to four other atoms. The name pentane contains the prefix *penta-* for five (Table 2.6), and we are told that the carbons are in a linear chain. If we then add enough hydrogen atoms to make four bonds to each carbon, we obtain the structural formula

$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & H & \\ & & | & & | & & | & & | & & | & \\ H & - & C & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & | & \\ & & H & & H & & H & & H & & H & \end{array}$$

This form of pentane is often called *n*-pentane, where the *n*- stands for "normal" because all five carbon atoms are in one line in the structural formula.

(**b**) Once the structural formula is written, we determine the molecular formula by counting the atoms present. Thus, *n*-pentane has the molecular formula $C_5H_{12}$.

#### PRACTICE EXERCISE

(**a**) What is the molecular formula of butane, the alkane with four carbons? (**b**) What are the name and molecular formula of an alcohol derived from butane?

***Answers:*** (**a**) $C_4H_{10}$, (**b**) butanol, $C_4H_{10}O$ or $C_4H_9OH$

## CHAPTER SUMMARY AND KEY TERMS

**SECTIONS 2.1 AND 2.2** **Atoms** are the basic building blocks of matter. They are the smallest units of an element that can combine with other elements. Atoms are composed of even smaller particles, called **subatomic particles**. Some of these subatomic particles are charged and follow the usual behavior of charged particles: Particles with the same charge repel one another, whereas particles with unlike charges are attracted to one another. We considered some of the important experiments that led to the discovery and characterization of subatomic particles. Thomson's experiments on the behavior of **cathode rays** in magnetic and electric fields led to the discovery of the electron and allowed its charge-to-mass ratio to be measured. Millikan's oil-drop experiment determined the charge of the electron. Becquerel's discovery of **radioactivity**, the spontaneous emission of radiation by atoms, gave further evidence that the atom has a substructure. Rutherford's studies of how thin metal foils scatter $\alpha$ particles led to the **nuclear model** of the atom, showing that the atom has a dense, positively charged **nucleus**.

**SECTION 2.3** Atoms have a nucleus that contains **protons** and **neutrons**; **electrons** move in the space around the nucleus. The magnitude of the charge of the electron, $1.602 \times 10^{-19}$ C, is called the **electronic charge**. The charges of particles are usually represented as multiples of this charge—an electron has a 1− charge, and a proton has a 1+ charge. The masses of atoms are usually expressed in terms of **atomic mass units** (1 amu = $1.66054 \times 10^{-24}$ g). The dimensions of atoms are often expressed in units of **angstroms** (1 Å = $10^{-10}$ m).

Elements can be classified by **atomic number**, the number of protons in the nucleus of an atom. All atoms of a given element have the same atomic number. The **mass number** of an atom is the sum of the numbers of protons and neutrons. Atoms of the same element that differ in mass number are known as **isotopes**.

**SECTION 2.4** The atomic mass scale is defined by assigning a mass of exactly 12 amu to a $^{12}C$ atom. The **atomic weight** (average atomic mass) of an element can be calculated from the relative abundances and masses of that element's isotopes. The **mass spectrometer** provides the most direct and accurate means of experimentally measuring atomic (and molecular) weights.

**SECTION 2.5** The **periodic table** is an arrangement of the elements in order of increasing atomic number. Elements with similar properties are placed in vertical columns. The elements in a column are known as a **group**. The elements in a horizontal row are known as a **period**. The **metallic elements** (**metals**), which comprise the majority of the elements, dominate the left side and the middle of the table; the **nonmetallic elements** (**nonmetals**) are located on the upper right side. Many of the elements that lie along the line that separates metals from nonmetals are **metalloids**.

**SECTION 2.6** Atoms can combine to form **molecules**. Compounds composed of molecules (**molecular compounds**) usually contain only nonmetallic elements. A molecule that contains two atoms is called a **diatomic molecule**. The composition of a substance is given by its **chemical formula**. A molecular substance can be represented by its **empirical formula**, which gives the relative numbers of atoms of each kind. It is usually represented by its **molecular formula**, however, which gives the actual numbers of each type of atom in a molecule. **Structural formulas** show the order in which the atoms in a molecule are connected. Ball-and-stick models and space-filling models are often used to represent molecules.

**SECTION 2.7** Atoms can either gain or lose electrons, forming charged particles called **ions**. Metals tend to lose electrons, becoming positively charged ions (**cations**). Nonmetals tend to gain electrons, forming negatively charged ions (**anions**). Because **ionic compounds** are electrically neutral, containing both cations and anions, they usually contain both metallic and nonmetallic elements. Atoms that are joined together, as in a molecule, but carry a net charge are called **polyatomic ions**. The chemical formulas used for ionic compounds are empirical formulas, which can be written readily if the charges of the ions are known. The total positive charge of the cations in an ionic compound equals the total negative charge of the anions.

**SECTION 2.8** The set of rules for naming chemical compounds is called **chemical nomenclature**. We studied the systematic rules used for naming three classes of inorganic substances: ionic compounds, acids, and binary molecular compounds. In naming an ionic compound, the cation is named first and then the anion. Cations formed from metal atoms have the same name as the metal. If the metal can form cations of differing charges, the charge is given using Roman numerals. Monatomic anions have names ending in *-ide*. Polyatomic anions containing oxygen and another element (**oxyanions**) have names ending in *-ate* or *-ite*.

**SECTION 2.9** **Organic chemistry** is the study of compounds that contain carbon. The simplest class of organic molecules is the **hydrocarbons**, which contain only carbon and hydrogen. Hydrocarbons in which each carbon atom is attached to four other atoms are called **alkanes**. Alkanes have names that end in *-ane*, such as methane and ethane. Other organic compounds are formed when an H atom of a hydrocarbon is replaced with a functional group. An **alcohol**, for example, is a compound in which an H atom of a hydrocarbon is replaced by an OH functional group. Alcohols have names that end in *-ol*, such as methanol and ethanol. Compounds with the same molecular formula but a different bonding arrangement of their constituent atoms are called **isomers**.

## KEY SKILLS

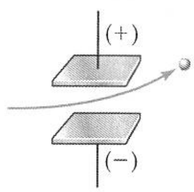
- Describe the basic postulates of Dalton's atomic theory. (Section 2.1)
- Describe the key experiments that led to the discovery of electrons and to the nuclear model of the atom. (Section 2.2)
- Describe the structure of the atom in terms of protons, neutrons, and electrons. (Section 2.3)
- Describe the electrical charge and relative masses of protons, neutrons, and electrons. (Section 2.3)
- Use chemical symbols together with atomic number and mass number to express the subatomic composition of isotopes. (Section 2.3)
- Understand how atomic weights relate to the masses of individual atoms and to their natural abundances. (Section 2.4)
- Describe how elements are organized in the periodic table by atomic number and by similarities in chemical behavior, giving rise to periods and groups. (Section 2.5)
- Describe the locations of metals and nonmetals in the periodic table. (Section 2.5)
- Distinguish between molecular substances and ionic substances in terms of their composition. (Sections 2.6 and 2.7)
- Distinguish between empirical formulas and molecular formulas. (Section 2.6)
- Describe how molecular formulas and structural formulas are used to represent the compositions of molecules. (Section 2.6)
- Explain how ions are formed by the gain or loss of electrons and be able to use the periodic table to predict the charges of common ions. (Section 2.7)
- Write the empirical formulas of ionic compounds, given the charges of their component ions. (Section 2.7)
- Write the name of an ionic compound given its chemical formula, or write the chemical formula given its name. (Section 2.8)
- Name or write chemical formulas for binary inorganic compounds and for acids. (Section 2.8)
- Identify organic compounds and name simple alkanes and alcohols. (Section 2.9)

# EXERCISES

## VISUALIZING CONCEPTS

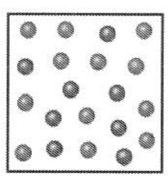
These exercises are intended to probe your understanding of key concepts rather than your ability to utilize formulas and perform calculations. Exercises with red exercise numbers have answers in the back of the book.

2.1 A charged particle is caused to move between two electrically charged plates, as shown here.

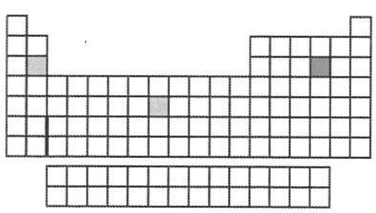



(**a**) Why does the path of the charged particle bend? (**b**) What is the sign of the electrical charge on the particle? (**c**) As the charge on the plates is increased, would you expect the bending to increase, decrease, or stay the same? (**d**) As the mass of the particle is increased while the speed of the particles remains the same, would you expect the bending to increase, decrease, or stay the same? [Section 2.2]

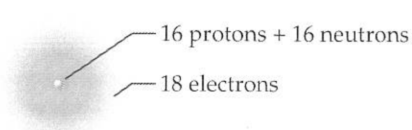
**2.2** The following diagram is a representation of 20 atoms of a fictitious element, which we will call nevadium (Nv). The red spheres are $^{293}Nv$, and the blue spheres are $^{295}Nv$. (**a**) Assuming that this sample is a statistically representative sample of the element, calculate the percent abundance of each element. (**b**) If the mass of $^{293}Nv$ is 293.15 amu and that of $^{295}Nv$ is 295.15 amu, what is the atomic weight of Nv? [Section 2.4]



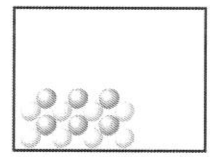
**2.3** Four of the boxes in the following periodic table are colored. Which of these are metals and which are nonmetals? Which one is an alkaline earth metal? Which one is a noble gas? [Section 2.5]



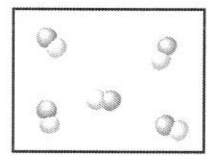
2.4 Does the following drawing represent a neutral atom or an ion? Write its complete chemical symbol including mass number, atomic number, and net charge (if any). [Sections 2.3 and 2.7]




**2.5** Which of the following diagrams most likely represents an ionic compound, and which represents a molecular one? Explain your choice. [Sections 2.6 and 2.7]

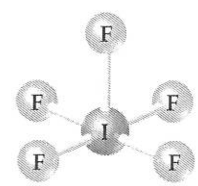


(i)

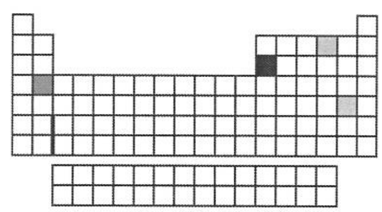


(ii)

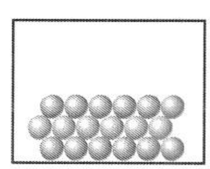
2.6 Write the chemical formula for the following compound. Is the compound ionic or molecular? Name the compound. [Sections 2.6 and 2.8]




**2.7** Five of the boxes in the following periodic table are colored. Predict the charge on the ion associated with each of these elements. [Section 2.7]



**2.8** The following diagram represents an ionic compound in which the red spheres represent cations and blue spheres represent anions. Which of the following formulas is consistent with the drawing: KBr, $K_2SO_4$, $Ca(NO_3)_2$, $Fe_2(SO_4)_3$? Name the compound. [Sections 2.7 and 2.8]



These exercises are divided into sections that deal with specific topics in the chapter. The exercises are grouped in pairs, with the answers given in the back of the book to the odd-numbered exercises, as indicated by the red exercise numbers. Those exercises whose numbers appear in brackets are more challenging than the nonbracketed exercises.

## ATOMIC THEORY AND THE DISCOVERY OF ATOMIC STRUCTURE (sections 2.1–2.2)

**2.9** How does Dalton's atomic theory account for the fact that when 1.000 g of water is decomposed into its elements, 0.111 g of hydrogen and 0.889 g of oxygen are obtained regardless of the source of the water?

**2.10** Hydrogen sulfide is composed of two elements: hydrogen and sulfur. In an experiment, 6.500 g of hydrogen sulfide is fully decomposed into its elements. **(a)** If 0.384 g of hydrogen is obtained in this experiment, how many grams of sulfur must be obtained? **(b)** What fundamental law does this experiment demonstrate? **(c)** How is this law explained by Dalton's atomic theory?

**2.11** A chemist finds that 30.82 g of nitrogen will react with 17.60 g, 35.20 g, 70.40 g, or 88.00 g of oxygen to form four different compounds. **(a)** Calculate the mass of oxygen per gram of nitrogen in each compound. **(b)** How do the numbers in part (a) support Dalton's atomic theory?

**2.12** In a series of experiments, a chemist prepared three different compounds that contain only iodine and fluorine and determined the mass of each element in each compound:

| Compound | Mass of Iodine (g) | Mass of Fluorine (g) |
|---|---|---|
| 1 | 4.75 | 3.56 |
| 2 | 7.64 | 3.43 |
| 3 | 9.41 | 9.86 |

**(a)** Calculate the mass of fluorine per gram of iodine in each compound. **(b)** How do the numbers in part (a) support the atomic theory?

**2.13** Summarize the evidence used by J. J. Thomson to argue that cathode rays consist of negatively charged particles.

**2.14** An unknown particle is caused to move between two electrically charged plates, as illustrated in Figure 2.8. Its path is deflected by a smaller magnitude in the opposite direction from that of a beta particle. What can you conclude about the charge and mass of this unknown particle?

**2.15** How did Rutherford interpret the following observations made during his $\alpha$-particle scattering experiments? **(a)** Most $\alpha$ particles were not appreciably deflected as they passed through the gold foil. **(b)** A few $\alpha$ particles were deflected at very large angles. **(c)** What differences would you expect if beryllium foil were used instead of gold foil in the $\alpha$-particle scattering experiment?

**2.16** Millikan determined the charge on the electron by studying the static charges on oil drops falling in an electric field (Figure 2.5). A student carried out this experiment using several oil drops for her measurements and calculated the charges on the drops. She obtained the following data:

| Droplet | Calculated Charge (C) |
|---|---|
| A | $1.60 \times 10^{-19}$ |
| B | $3.15 \times 10^{-19}$ |
| C | $4.81 \times 10^{-19}$ |
| D | $6.31 \times 10^{-19}$ |

**(a)** What is the significance of the fact that the droplets carried different charges? **(b)** What conclusion can the student draw from these data regarding the charge of the electron? **(c)** What value (and to how many significant figures) should she report for the electronic charge?

## MODERN VIEW OF ATOMIC STRUCTURE; ATOMIC WEIGHTS (sections 2.3–2.4)

**2.17** The radius of an atom of gold (Au) is about 1.35 Å. **(a)** Express this distance in nanometers (nm) and in picometers (pm). **(b)** How many gold atoms would have to be lined up to span 1.0 mm? **(c)** If the atom is assumed to be a sphere, what is the volume in $cm^3$ of a single Au atom?

**2.18** An atom of rhodium (Rh) has a diameter of about $2.7 \times 10^{-8}$ cm. **(a)** What is the radius of a rhodium atom in angstroms (Å) and in meters (m)? **(b)** How many Rh atoms would have to be placed side by side to span a distance of 6.0 $\mu$m? **(c)** If you assume that the Rh atom is a sphere, what is the volume in $m^3$ of a single atom?

**2.19** Answer the following questions without referring to Table 2.1: **(a)** What are the main subatomic particles that make up the atom? **(b)** What is the relative charge (in multiples of the electronic charge) of each of the particles? **(c)** Which of the particles is the most massive? **(d)** Which is the least massive?

**2.20** Determine whether each of the following statements is true or false. If false, correct the statement to make it true: **(a)** The nucleus has most of the mass and comprises most of the volume of an atom. **(b)** Every atom of a given element has the same number of protons. **(c)** The number of electrons in an atom equals the number of neutrons in the atom. **(d)** The protons in the nucleus of the helium atom are held together by a force called the strong nuclear force.

**2.21** **(a)** Define atomic number and mass number. **(b)** Which of these can vary without changing the identity of the element?

**2.22** **(a)** Which two of the following are isotopes of the same element: $^{31}_{16}X$, $^{31}_{15}X$, $^{32}_{16}X$? **(b)** What is the identity of the element whose isotopes you have selected?

**2.23** How many protons, neutrons, and electrons are in the following atoms: **(a)** $^{40}Ar$, **(b)** $^{65}Zn$, **(c)** $^{70}Ga$, **(d)** $^{80}Br$, **(e)** $^{184}W$, **(f)** $^{243}Am$?

**2.24** Each of the following isotopes is used in medicine. Indicate the number of protons and neutrons in each isotope: **(a)** phosphorus-32, **(b)** chromium-51, **(c)** cobalt-60, **(d)** technetium-99, **(e)** iodine-131, **(f)** thallium-201.

2.25 Fill in the gaps in the following table, assuming each column represents a neutral atom.

| Symbol | $^{52}Cr$ | | | | |
|---|---|---|---|---|---|
| Protons | | 25 | | | 82 |
| Neutrons | | 30 | 64 | | |
| Electrons | | | 48 | 86 | |
| Mass no. | | | | 222 | 207 |

**2.26** Fill in the gaps in the following table, assuming each column represents a neutral atom.

| Symbol | $^{65}Zn$ | | | | |
|---|---|---|---|---|---|
| Protons | | 38 | | | 92 |
| Neutrons | | 58 | 49 | | |
| Electrons | | | 38 | 36 | |
| Mass no. | | | | 81 | 235 |

2.27 Write the correct symbol, with both superscript and subscript, for each of the following. Use the list of elements inside the front cover as needed: **(a)** the isotope of platinum that contains 118 neutrons, **(b)** the isotope of krypton with mass number 84, **(c)** the isotope of arsenic with mass number 75, **(d)** the isotope of magnesium that has an equal number of protons and neutrons.

**2.28** One way in which Earth's evolution as a planet can be understood is by measuring the amounts of certain isotopes in rocks. One quantity recently measured is the ratio of $^{129}Xe$ to $^{130}Xe$ in some minerals. In what way do these two isotopes differ from one another? In what respects are they the same?

2.29 **(a)** What isotope is used as the standard in establishing the atomic mass scale? **(b)** The atomic weight of boron is reported as 10.81, yet no atom of boron has the mass of 10.81 amu. Explain.

**2.30** **(a)** What is the mass in amu of a carbon-12 atom? **(b)** Why is the atomic weight of carbon reported as 12.011 in the table of elements and the periodic table in the front inside cover of this text?

2.31 Only two isotopes of copper occur naturally, $^{63}Cu$ (atomic mass = 62.9296 amu; abundance 69.17%) and $^{65}Cu$ (atomic mass = 64.9278 amu; abundance 30.83%). Calculate the atomic weight (average atomic mass) of copper.

**2.32** Rubidium has two naturally occurring isotopes, rubidium-85 (atomic mass = 84.9118 amu; abundance = 72.15%) and rubidium-87 (atomic mass = 86.9092 amu; abundance = 27.85%). Calculate the atomic weight of rubidium.

2.33 **(a)** In what fundamental way is mass spectrometry related to Thomson's cathode-ray experiments (Figure 2.4)? **(b)** What are the labels on the axes of a mass spectrum? **(c)** To measure the mass spectrum of an atom, the atom must first lose one or more electrons. Why is this so?

**2.34** **(a)** The mass spectrometer in Figure 2.12 has a magnet as one of its components. What is the purpose of the magnet? **(b)** The atomic weight of Cl is 35.5 amu. However, the mass spectrum of Cl (Figure 2.13) does not show a peak at this mass. Explain. **(c)** A mass spectrum of phosphorus (P) atoms shows only a single peak at a mass of 31. What can you conclude from this observation?

2.35 Naturally occurring magnesium has the following isotopic abundances:

| Isotope | Abundance | Atomic mass (amu) |
|---|---|---|
| $^{24}Mg$ | 78.99 % | 23.98504 |
| $^{25}Mg$ | 10.00 % | 24.98584 |
| $^{26}Mg$ | 11.01 % | 25.98259 |

**(a)** What is the average atomic mass of Mg? **(b)** Sketch the mass spectrum of Mg.

**2.36** Mass spectrometry is more often applied to molecules than to atoms. We will see in Chapter 3 that the *molecular weight* of a molecule is the sum of the atomic weights of the atoms in the molecule. The mass spectrum of $H_2$ is taken under conditions that prevent decomposition into H atoms. The two naturally occurring isotopes of hydrogen are $^{1}H$ (atomic mass = 1.00783 amu; abundance 99.9885%) and $^{2}H$ (atomic mass = 2.01410 amu; abundance 0.0115%). **(a)** How many peaks will the mass spectrum have? **(b)** Give the relative atomic masses of each of these peaks. **(c)** Which peak will be the largest and which the smallest?

## THE PERIODIC TABLE; MOLECULES AND IONS (sections 2.5–2.7)

2.37 For each of the following elements, write its chemical symbol, locate it in the periodic table, give its atomic number, and indicate whether it is a metal, metalloid, or nonmetal: **(a)** chromium, **(b)** helium, **(c)** phosphorus, **(d)** zinc, **(e)** magnesium, **(f)** bromine, **(g)** arsenic.

**2.38** Locate each of the following elements in the periodic table; give its name and atomic number, and indicate whether it is a metal, metalloid, or nonmetal: **(a)** Li, **(b)** Sc, **(c)** Ge, **(d)** Yb, **(e)** Mn, **(f)** Sb, **(g)** Xe.

2.39 For each of the following elements, write its chemical symbol, determine the name of the group to which it belongs (Table 2.3), and indicate whether it is a metal, metalloid, or nonmetal: **(a)** potassium, **(b)** iodine, **(c)** magnesium, **(d)** argon, **(e)** sulfur.

**2.40** The elements of group 4A show an interesting change in properties moving down the group. Give the name and chemical symbol of each element in the group and label it as a nonmetal, metalloid, or metal.

2.41 What can we tell about a compound when we know the empirical formula? What additional information is conveyed by the molecular formula? By the structural formula? Explain in each case.

**2.42** Two compounds have the same empirical formula. One substance is a gas, whereas the other is a viscous liquid. How is it possible for two substances with the same empirical formula to have markedly different properties?

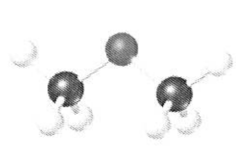
2.43 Write the empirical formula corresponding to each of the following molecular formulas: **(a)** $Al_2Br_6$, **(b)** $C_8H_{10}$, **(c)** $C_4H_8O_2$, **(d)** $P_4O_{10}$, **(e)** $C_6H_4Cl_2$, **(f)** $B_3N_3H_6$.

**2.44** Determine the molecular and empirical formulas of the following: (a) the organic solvent *benzene*, which has six carbon atoms and six hydrogen atoms; (b) the compound *silicon tetrachloride*, which has a silicon atom and four chlorine atoms and is used in the manufacture of computer chips; (c) the reactive substance *diborane*, which has two boron atoms and six hydrogen atoms; (d) the sugar called *glucose*, which has six carbon atoms, twelve hydrogen atoms, and six oxygen atoms.

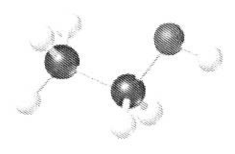
2.45 How many hydrogen atoms are in each of the following: (a) $C_2H_5OH$, (b) $Ca(CH_3COO)_2$, (c) $(NH_4)_3PO_4$?

**2.46** How many of the indicated atoms are represented by each chemical formula: (a) carbon atoms in $C_2H_5COOCH_3$, (b) oxygen atoms in $Ca(ClO_4)_2$, (c) hydrogen atoms in $(NH_4)_2HPO_4$?

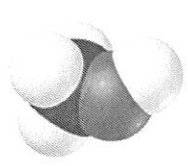
2.47 Write the molecular and structural formulas for the compounds represented by the following molecular models:



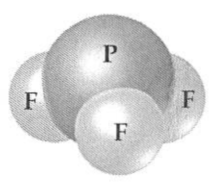
(a)



(b)

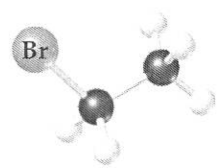


(c)

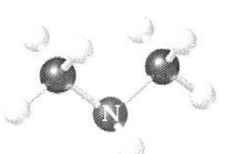



(d)

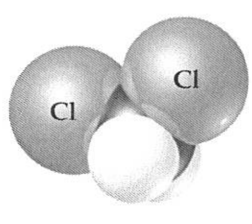
**2.48** Write the molecular and structural formulas for the compounds represented by the following models:



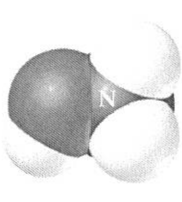

(a)




(b)




(c)




(d)

2.49 Fill in the gaps in the following table:

| Symbol | $^{59}Co^{3+}$ | | | |
|---|---|---|---|---|
| Protons | | 34 | 76 | 80 |
| Neutrons | | 46 | 116 | 120 |
| Electrons | | 36 | | 78 |
| Net charge | | | 2+ | |

**2.50** Fill in the gaps in the following table:

| Symbol | $^{31}P^{3-}$ | | | |
|---|---|---|---|---|
| Protons | | 34 | 50 | |
| Neutrons | | 45 | 69 | 118 |
| Electrons | | | 46 | 76 |
| Net charge | | 2− | | 3+ |

2.51 Each of the following elements is capable of forming an ion in chemical reactions. By referring to the periodic table, predict the charge of the most stable ion of each: (a) Mg, (b) Al, (c) K, (d) S, (e) F.

**2.52** Using the periodic table, predict the charges of the ions of the following elements: (a) Ga, (b) Sr, (c) As, (d) Br, (e) Se.

2.53 Using the periodic table to guide you, predict the chemical formula and name of the compound formed by the following elements: (a) Ga and F, (b) Li and H, (c) Al and I, (d) K and S.

**2.54** The most common charge associated with scandium in its compounds is 3+. Indicate the chemical formulas you would expect for compounds formed between scandium and (a) iodine, (b) sulfur, (c) nitrogen.

2.55 Predict the chemical formula for the ionic compound formed by (a) $Ca^{2+}$ and $Br^-$, (b) $K^+$ and $CO_3^{2-}$, (c) $Al^{3+}$ and $CH_3COO^-$, (d) $NH_4^+$ and $SO_4^{2-}$, (e) $Mg^{2+}$ and $PO_4^{3-}$.

**2.56** Predict the chemical formulas of the compounds formed by the following pairs of ions: (a) $Cr^{3+}$ and $Br^-$, (b) $Fe^{3+}$ and $O^{2-}$, (c) $Hg_2^{2+}$ and $CO_3^{2-}$, (d) $Ca^{2+}$ and $ClO_3^-$, (e) $NH_4^+$ and $PO_4^{3-}$.

2.57 Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

| Ion | $K^+$ | $NH_4^+$ | $Mg^{2+}$ | $Fe^{3+}$ |
|---|---|---|---|---|
| $Cl^-$ | KCl | | | |
| $OH^-$ | | | | |
| $CO_3^{2-}$ | | | | |
| $PO_4^{3-}$ | | | | |

**2.58** Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

| Ion | $Na^+$ | $Ca^{2+}$ | $Fe^{2+}$ | $Al^{3+}$ |
|---|---|---|---|---|
| $O^{2-}$ | $Na_2O$ | | | |
| $NO_3^-$ | | | | |
| $SO_4^{2-}$ | | | | |
| $AsO_4^{3-}$ | | | | |

2.59 Predict whether each of the following compounds is molecular or ionic: (a) $B_2H_6$, (b) $CH_3OH$, (c) $LiNO_3$, (d) $Sc_2O_3$, (e) CsBr, (f) NOCl, (g) $NF_3$, (h) $Ag_2SO_4$.

**2.60** Which of the following are ionic, and which are molecular? (a) $PF_5$, (b) NaI, (c) $SCl_2$, (d) $Ca(NO_3)_2$, (e) $FeCl_3$, (f) LaP, (g) $CoCO_3$, (h) $N_2O_4$.

## NAMING INORGANIC COMPOUNDS; ORGANIC MOLECULES (sections 2.8–2.9)

2.61 Give the chemical formula for (a) chlorite ion, (b) chloride ion, (c) chlorate ion, (d) perchlorate ion, (e) hypochlorite ion.

**2.62** Selenium, an element required nutritionally in trace quantities, forms compounds analogous to sulfur. Name the following ions: (a) $SeO_4^{2-}$, (b) $Se^{2-}$, (c) $HSe^-$, (d) $HSeO_3^-$.

2.63 Give the names and charges of the cation and anion in each of the following compounds: (a) CaO, (b) $Na_2SO_4$, (c) $KClO_4$, (d) $Fe(NO_3)_2$, (e) $Cr(OH)_3$.

**2.64** Give the names and charges of the cation and anion in each of the following compounds: (a) CuS, (b) $Ag_2SO_4$, (c) $Al(ClO_3)_3$, (d) $Co(OH)_2$, (e) $PbCO_3$.

2.65 Name the following ionic compounds: (a) $Li_2O$, (b) $FeCl_3$, (c) NaClO, (d) $CaSO_3$, (e) $Cu(OH)_2$, (f) $Fe(NO_3)_2$, (g) $Ca(CH_3COO)_2$, (h) $Cr_2(CO_3)_3$, (i) $K_2CrO_4$, (j) $(NH_4)_2SO_4$.

**2.66** Name the following ionic compounds: (a) KCN, (b) $NaBrO_2$, (c) $Sr(OH)_2$, (d) CoS, (e) $Fe_2(CO_3)_3$, (f) $Cr(NO_3)_3$, (g) $(NH_4)_2SO_3$, (h) $NaH_2PO_4$, (i) $KMnO_4$, (j) $Ag_2Cr_2O_7$.

2.67 Write the chemical formulas for the following compounds: (a) aluminum hydroxide, (b) potassium sulfate, (c) copper(I) oxide, (d) zinc nitrate, (e) mercury(II) bromide, (f) iron(III) carbonate, (g) sodium hypobromite.

**2.68** Give the chemical formula for each of the following ionic compounds: (a) sodium phosphate, (b) zinc nitrate, (c) barium bromate, (d) iron(II) perchlorate, (e) cobalt(II) hydrogen carbonate, (f) chromium(III) acetate, (g) potassium dichromate.

2.69 Give the name or chemical formula, as appropriate, for each of the following acids: (a) $HBrO_3$, (b) HBr, (c) $H_3PO_4$, (d) hypochlorous acid, (e) iodic acid, (f) sulfurous acid.

**2.70** Provide the name or chemical formula, as appropriate, for each of the following acids: (a) hydroiodic acid, (b) chloric acid, (c) nitrous acid, (d) $H_2CO_3$, (e) $HClO_4$, (f) $CH_3COOH$

2.71 Give the name or chemical formula, as appropriate, for each of the following binary molecular substances: (a) $SF_6$, (b) $IF_5$, (c) $XeO_3$, (d) dinitrogen tetroxide, (e) hydrogen cyanide, (f) tetraphosphorus hexasulfide.

**2.72** The oxides of nitrogen are very important components in urban air pollution. Name each of the following compounds: (a) $N_2O$, (b) NO, (c) $NO_2$, (d) $N_2O_5$, (e) $N_2O_4$.

2.73 Write the chemical formula for each substance mentioned in the following word descriptions (use the front inside cover to find the symbols for the elements you don't know). (a) Zinc carbonate can be heated to form zinc oxide and carbon dioxide. (b) On treatment with hydrofluoric acid, silicon dioxide forms silicon tetrafluoride and water. (c) Sulfur dioxide reacts with water to form sulfurous acid. (d) The substance phosphorus trihydride, commonly called phosphine, is a toxic gas. (e) Perchloric acid reacts with cadmium to form cadmium(II) perchlorate. (f) Vanadium(III) bromide is a colored solid.

**2.74** Assume that you encounter the following sentences in your reading. What is the chemical formula for each substance mentioned? (a) Sodium hydrogen carbonate is used as a deodorant. (b) Calcium hypochlorite is used in some bleaching solutions. (c) Hydrogen cyanide is a very poisonous gas. (d) Magnesium hydroxide is used as a cathartic. (e) Tin(II) fluoride has been used as a fluoride additive in toothpastes. (f) When cadmium sulfide is treated with sulfuric acid, fumes of hydrogen sulfide are given off.

2.75 (a) What is a hydrocarbon? (b) Butane is the alkane with a chain of four carbon atoms. Write a structural formula for this compound and determine its molecular and empirical formulas.

**2.76** (a) What ending is used for the names of alkanes? (b) Hexane is an alkane whose structural formula has all its carbon atoms in a straight chain. Draw the structural formula for this compound and determine its molecular and empirical formulas. (*Hint:* You might need to refer to Table 2.6.)

2.77 (a) What is a functional group? (b) What functional group characterizes an alcohol? (c) With reference to Exercise 2.75, write a structural formula for 1-butanol, the alcohol derived from butane, by making a substitution on one of the carbon atoms.

**2.78** (a) What do ethane and ethanol have in common? (b) How does 1-propanol differ from propane?

2.79 Chloropropane is a compound derived from propane by substituting Cl for H on one of the carbon atoms. (a) Draw the structural formulas for the two isomers of chloropropane. (b) Suggest names for these two compounds.

**2.80** Draw the structural formulas for three isomers of pentane, $C_5H_{12}$.

## ADDITIONAL EXERCISES

These exercises are not divided by category, although they are roughly in the order of the topics in the chapter. They are not paired.

**2.81** Suppose a scientist repeats the Millikan oil-drop experiment but reports the charges on the drops using an unusual (and imaginary) unit called the *warmomb* (wa). The scientist obtains the following data for four of the drops:

| Droplet | Calculated Charge (wa) |
|---|---|
| A | $3.84 \times 10^{-8}$ |
| B | $4.80 \times 10^{-8}$ |
| C | $2.88 \times 10^{-8}$ |
| D | $8.64 \times 10^{-8}$ |

(a) If all the droplets were the same size, which would fall most slowly through the apparatus? (b) From these data, what is the best choice for the charge of the electron in warmombs? (c) Based on your answer to part (b), how many electrons are there on each of the droplets? (d) What is the conversion factor between warmombs and coulombs?