2

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 beauty and complexity of a geode like the one shown to the right. How can we explain the striking and seemingly infinite variety of properties of the materials that make up our world? What makes diamonds transparent and hard? A large crystal of sodium chloride, table salt, looks a bit like a diamond, but is brittle and readily dissolves in water. What accounts for the differences? Why does paper burn, and why does water quench fires? The answers to all such questions lie in the structures of atoms, which determine 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, these different 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 introduce 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.

▶ **A SECTION THROUGH A GEODE.** A geode is a mass of mineral matter (often containing quartz) that accumulates slowly within the shell of a roughly spherical, hollow rock. Eventually, perfectly formed crystals may develop at a geode's center. The colors of a geode depend upon its composition. Here, agate crystallized out as the geode formed.

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.

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.

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 BCE) and other early Greek philosophers described the material world as made up of tiny indivisible particles that 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 four postulates (see ▼ 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:

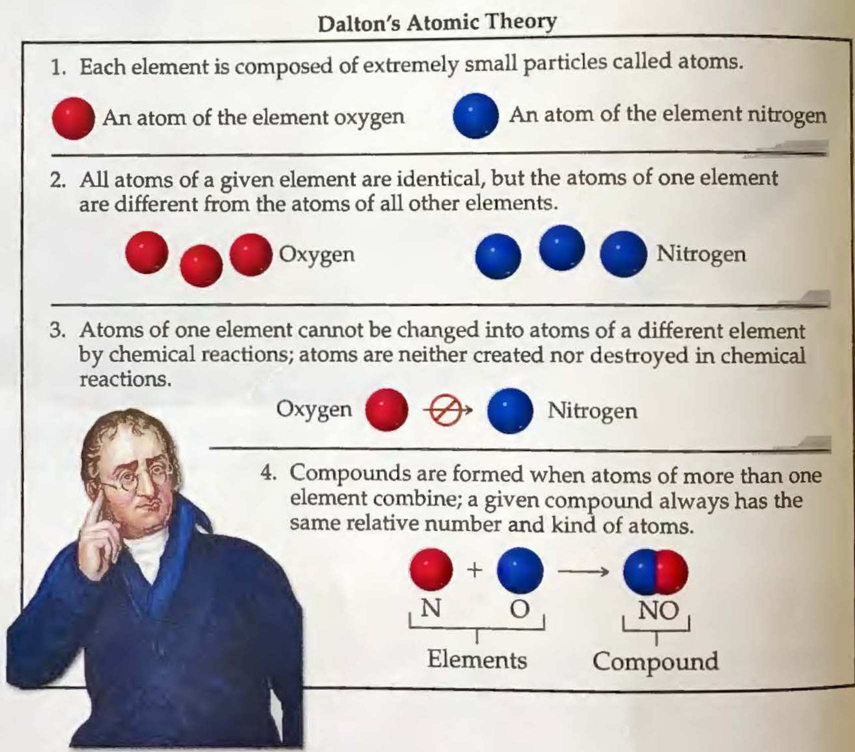
> In a given compound, the relative numbers and kinds of atoms are constant.

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.

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

†Dalton, John. "Atomic Theory." 1844.

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 masses 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 than does water.

Give It Some Thought

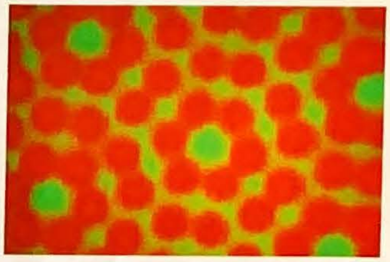
Compound A contains 1.333 g of oxygen per gram of carbon, whereas compound B contains 2.666 g of oxygen per gram of carbon.

(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. By assuming the existence of atoms he was able to account for the laws of constant composition and of multiple proportions. But neither Dalton 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**).



▲ 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 red sphere is a silicon atom.

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.*

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.

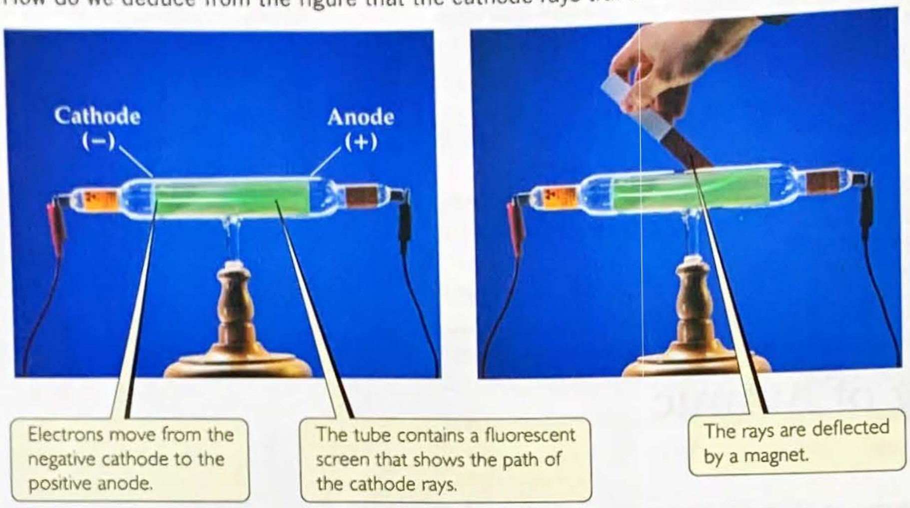
Experiments showed that cathode rays are deflected by electric or magnetic fields in a way consistent with their being a stream of negative electrical charge. The British scientist J. J. Thomson (1856–1940) observed that cathode rays are the same regardless of the identity of the cathode material. In 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 1.76×10^8 coulombs* per gram for the ratio of the electron's electrical charge to its mass.

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

GO FIGURE

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



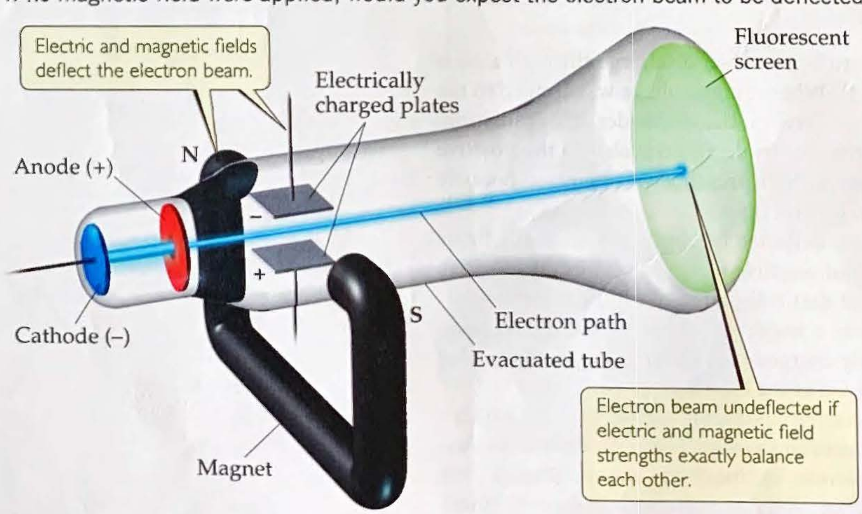
▲ **Figure 2.3 Cathode-ray tube.**

Give It Some Thought

Thomson observed that the cathode rays produced in the cathode-ray tube behaved identically, regardless of the particular metal used as cathode. What is the significance of this observation?

GO FIGURE

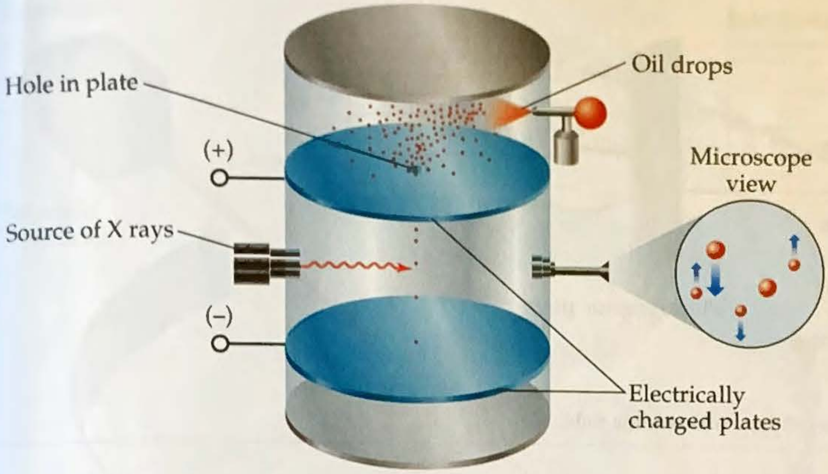
If no magnetic field were applied, would you expect the electron beam to be deflected upward or downward by the electric field?



▲ **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.

GO FIGURE

Would the masses of the oil drops be changed significantly by any electrons that accumulate on them?



▲ **Figure 2.5 Millikan's oil-drop experiment to measure the charge of the electron.** Small drops of oil are allowed to fall between electrically charged plates. The drops pick up extra electrons as a result of irradiation by X-rays and so became negatively charged. Millikan measured how varying the voltage between the plates affected the rate of fall. From these data he calculated the negative charge on the drops. Because the charge on any drop was always some integral multiple of 1.602×10^{-19} C, Millikan deduced this value to be the charge of a single electron.

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, 1.602×10^{-19} C, and Thomson's charge-to-mass ratio, 1.76×10^8 C/g:

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

This result agrees well with the currently accepted value for the electron mass, 9.10938×10^{-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 identify and isolate the source of radioactivity in the compound. They concluded that it was the uranium atoms.

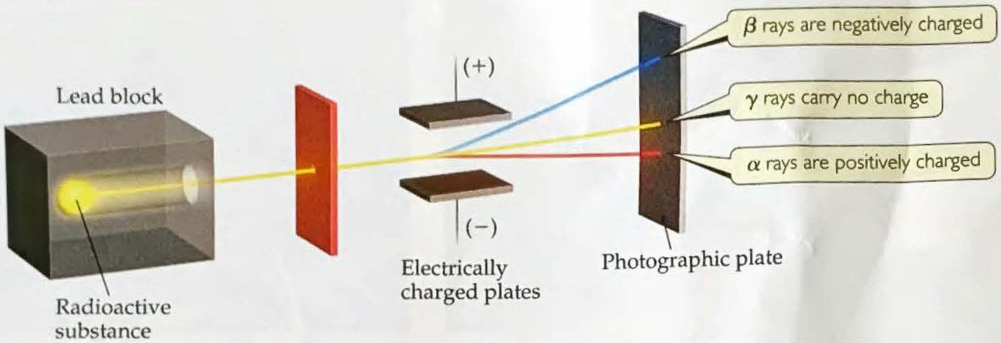
Further study of radioactivity, principally by the British scientist Ernest Rutherford, revealed three types of radiation: alpha (α), beta (β), and gamma (γ). The paths of α and β radiation are bent by an electric field, although in opposite directions; γ radiation is unaffected by the field (**Figure 2.7**). Rutherford (1871–1937) was a very important figure in this period of atomic science. After working at Cambridge University with J. J. Thomson, he moved to McGill University in Montreal, where he did research on radioactivity that led to his 1908 Nobel Prize in Chemistry. In 1907 he returned to England as a faculty member at Manchester University, where he did his famous α-particle scattering experiments, described below.

Rutherford showed that α and β rays consist of fast-moving particles. In fact, β particles are high-speed electrons and can be considered the radioactive equivalent of cathode rays. They are attracted to a positively charged plate. The α particles have a

▲ **Figure 2.6 Marie Sklodowska Curie (1867–1934).** In 1903 Henri Becquerel, Marie 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.

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?

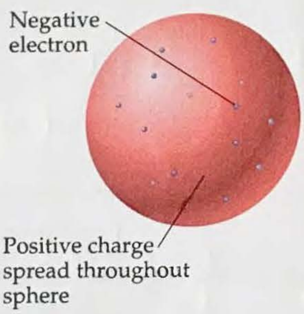


▲ **Figure 2.7 Behavior of alpha (α), beta (β), and gamma (γ) rays in an electric field.**

positive charge and are attracted to a negative plate. In units of the charge of the electron, β particles have a charge of $1-$ and α particles a charge of $2+$. Each α 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.

The Nuclear Model of the Atom

With growing evidence that the atom is composed of smaller particles, scientist gave attention 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 are responsible for an equally small fraction of the atom's size. He proposed that the atom consists of a uniform positive sphere of matter in which the mass is evenly distributed and in which the electrons are embedded like raisins in a pudding or seeds in a watermelon (◀ **Figure 2.8**). This *plum-pudding model*, named after a traditional English dessert, was very short-lived.



▲ **Figure 2.8 J. J. Thomson's plum-pudding model of the atom.** Ernest Rutherford and Ernest Marsden proved this model wrong.

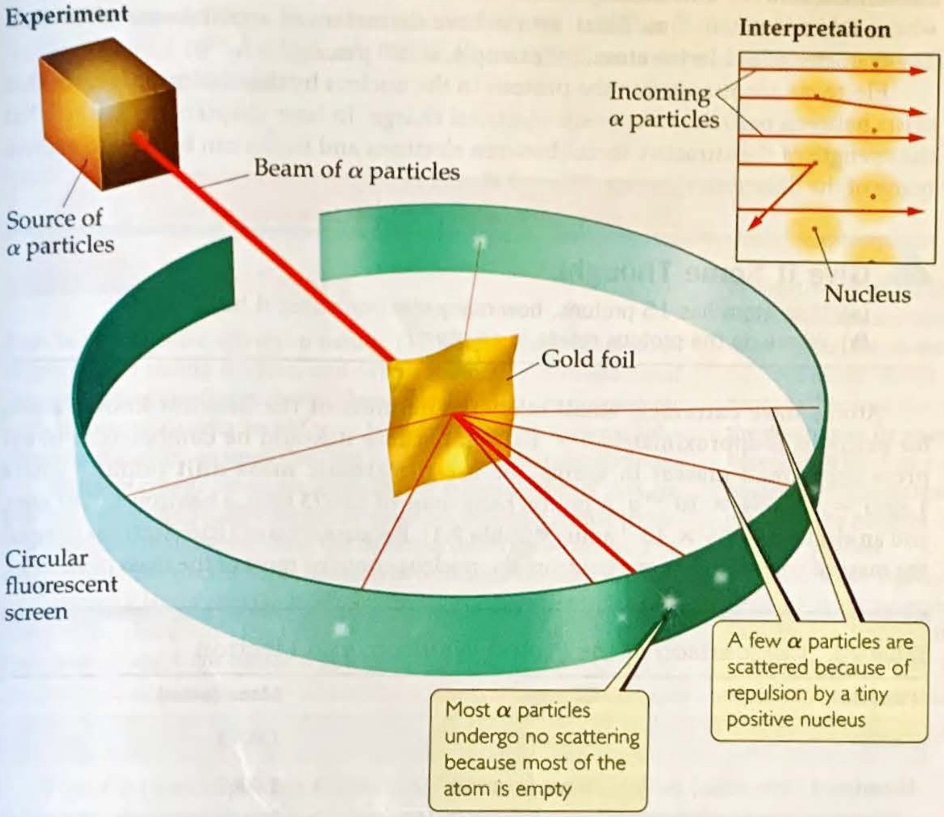
In 1910, Rutherford was studying the angles at which α particles were deflected, or *scattered*, as they passed through a thin sheet of gold foil (▶ **Figure 2.9**). He discovered that almost all the particles passed directly through the foil without deflection, with a few particles deflected about 1°, 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.

Rutherford explained the results by postulating the **nuclear model** of the atom, 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 α-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 α particle came close to a gold nucleus. In such encounters, the repulsion between the highly positive charge of the gold nucleus and the positive charge of the α particle was strong enough to deflect the particle, as shown in Figure 2.9.

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.

GO FIGURE

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



▲ **Figure 2.9** **Rutherford's α-Sacttering experiment.** When α 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 α particles are deflected at large angles. Although the nuclear atom has been depicted here as a yellow sphere, it is important to realize that most of the space around the nucleus contains only the low-mass electrons.

Give It Some Thought

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

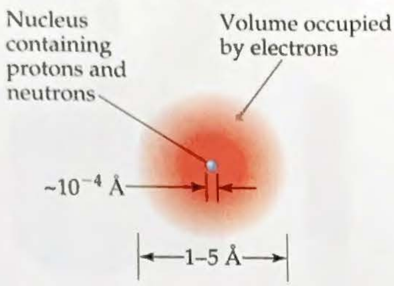
2.3 | The Modern View of Atomic Structure

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 -1.602×10^{-19} C. The charge of a proton is opposite in sign but equal in magnitude to that of an electron: $+1.602 \times 10^{-19}$ C. The quantity 1.602×10^{-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 an electron is 1− and that of a 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.*

GO FIGURE

What is the approximate diameter of the nucleus in units of pm?



▲ **Figure 2.10 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.

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 2.10**). Most atoms have diameters between 1×10^{-10} m (100 pm) and 5×10^{-10} m (500 pm). A convenient non-SI unit of length used for atomic dimensions is the **angstrom** (Å), where $1 \text{ Å} = 1 \times 10^{-10}$ 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.

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.

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 4×10^{-22} g. Because it would be cumbersome to express such small masses in grams, we use the **atomic mass unit** (amu),* where $1 \text{ amu} = 1.66054 \times 10^{-24}$ g. A proton has a mass of 1.0073 amu, a neutron 1.0087 amu, and an electron 5.486×10^{-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.

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−)	5.486×10^{-4}

SAMPLE EXERCISE 2.1 Atomic Size

The diameter of a U.S. 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 1

Which of the following factors determines the size of an atom?
(a) The volume of the nucleus; **(b)** the volume of space occupied by the electrons of the atom; **(c)** the volume of a single electron, multiplied by the number of electrons in the atom; **(d)** The total nuclear charge; **(e)** The total mass of the electrons surrounding the nucleus.

Practice Exercise 2

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?

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

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, 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!

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. Electric forces are of primary importance in determining the chemical properties of elements.

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*: 2.112**

Figure 2.10 incorporates the features we have just discussed. Electrons play the major role in chemical reactions. The significance of representing the region containing electrons as an indistinct cloud will become clear in later chapters when we consider the energies and spatial arrangements of the electrons. For now we have all the information we need to discuss many topics that form the basis of everyday uses of chemistry.

Atomic Numbers, Mass Numbers, and Isotopes

What makes an atom of one element different from an atom of another element? The atoms of each element have a *characteristic number of protons.* 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 front inside cover of the text.

Atoms of a given element can differ in the number of neutrons they contain and, consequently, in mass. For example, while most atoms of carbon have six neutrons, 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, whereas 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.

The atomic number is indicated 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) ↘

$^{12}_{6}C$ ⟵ Symbol of element

Atomic number (number of protons or electrons) ↗

Table 2.2 Some Isotopes of Carbon[a]

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

[a]Almost 99% of the carbon found in nature is ^{12}C.

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.

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. It is important to keep in mind that the isotopes of any given element are all alike chemically. A carbon dioxide molecule that contains a ^{13}C atom behaves for all practical purposes identically to one that contains a ^{12}C atom.

SAMPLE EXERCISE 2.2 Determining the Number of Subatomic Particles in Atoms

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

SOLUTION

(a) The superscript 197 is the mass number (protons + neutrons). According to the list of elements given on the front inside 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 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 1

Which of these atoms has the largest number of neutrons in the nucleus? **(a)** ^{148}Eu, **(b)** ^{157}Dy, **(c)** ^{149}Nd, **(d)** ^{162}Ho, **(e)** ^{159}Gd.

Practice Exercise 2

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

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 1

Which of the following is an incorrect representation for a neutral atom: **(a)** $^{6}_{3}Li$, **(b)** $^{13}_{6}C$, (c) $^{63}_{30}Cu$,**(d)** $^{30}_{15}P$,**(e)** $^{108}_{47}Ag$?

Practice Exercise 2

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

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 1.6735×10^{-24} g and the ^{16}O atom has a mass of 2.6560×10^{-23} g. As we noted in Section 2.3, it is convenient to use the **atomic mass unit** 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 a chemically unbound 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 summing (indicated by the Greek sigma, Σ) over the masses of its isotopes multiplied by 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 front inside 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 given as 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 1

The atomic weight of copper, Cu, is listed as 63.546. Which of the following statements are untrue?

(a) Not all the atoms of copper have the same number of electrons.
(b) All the copper atoms have 29 protons in the nucleus.
(c) The dominant isotopes of Cu must be ^{63}Cu and ^{64}Cu.
(d) Copper is a mixture of at least two isotopes.
(e) The number of electrons in the copper atoms is independent of atomic mass.

Practice Exercise 2

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.

A Closer Look

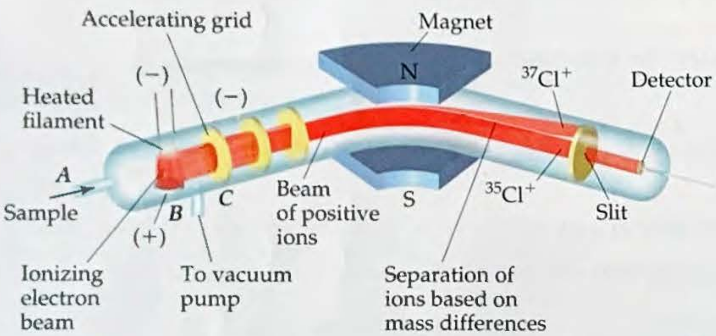
The Mass Spectrometer

The most accurate means for determining atomic weights is provided by the **mass spectrometer** (▼ **Figure 2.11**). 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, called *ions*, that are then accelerated toward a negatively charged grid (*C*). After the ions pass through the grid, they encounter two slits that allow only a narrow beam of ions to pass. This beam then passes between the poles of a magnet, which deflects the ions into a curved path. For ions with the same charge, the extent of deflection depends on mass—the more massive the ion, the less the deflection. The ions are thereby separated according to their masses. By changing the strength of the magnetic field or the accelerating voltage on the grid, ions of various masses can be selected to enter the detector.

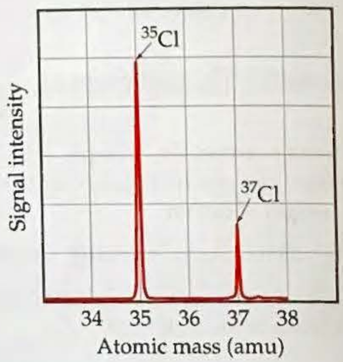
A graph of the intensity of the detector signal versus ion atomic mass is called a *mass spectrum* (▼ **Figure 2.12**). Analysis of a mass spectrum gives both the masses of the ions reaching the detector and their relative abundances, which are obtained from the signal intensities. Knowing the atomic 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: ***2.27, 2.38, 2.40, 2.88, 2.98, 2.99***



▲ **Figure 2.11 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 2.12 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.

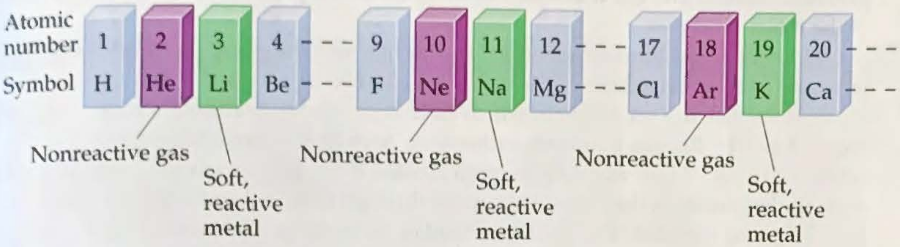
2.5 | The Periodic Table

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, respectively—as shown in ▶ **Figure 2.13**.

GO FIGURE

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



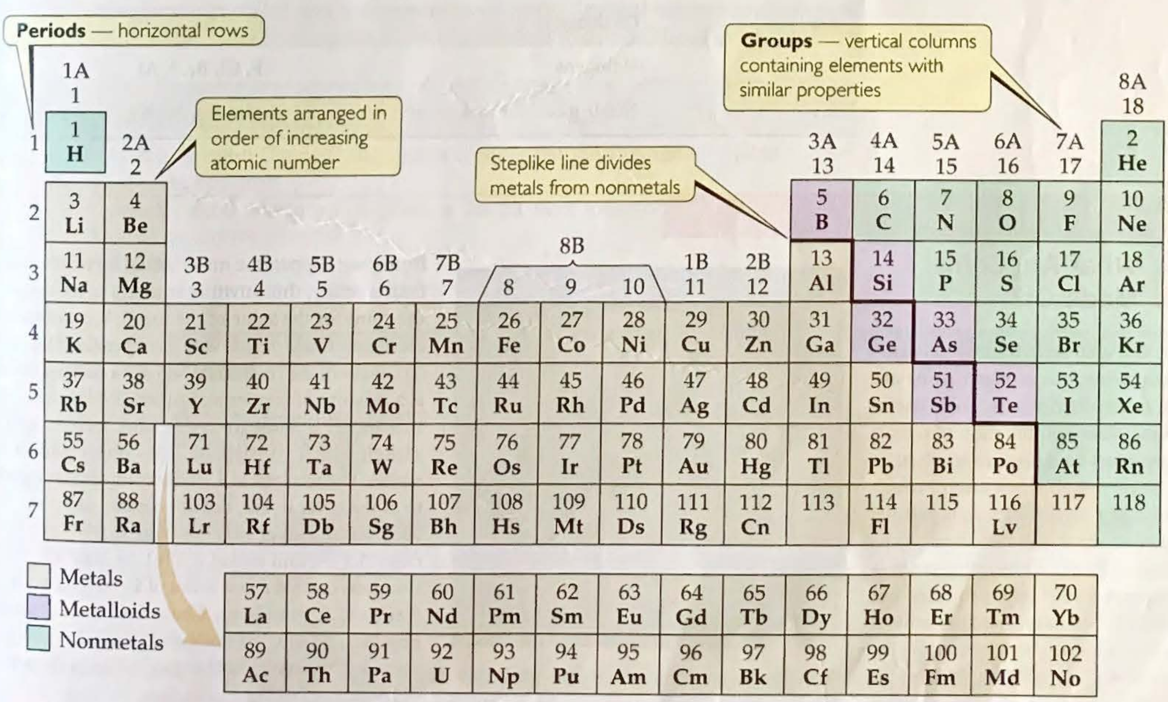
▲ **Figure 2.13 Arranging elements by atomic number reveals a periodic pattern of properties.** This pattern is the basis of the periodic table.

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 2.14**). 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.

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



▲ **Figure 2.14 Periodic table of elements.**

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 2.14. 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 2.14. 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 have been traditionally 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,

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

What Are Coins Made Of?

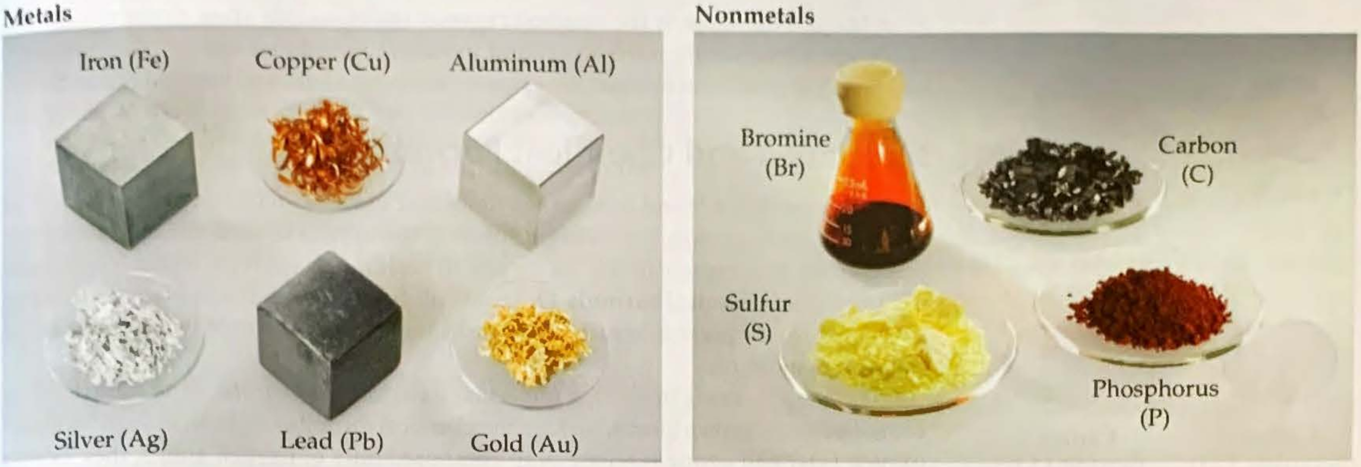
Copper, silver, and gold were traditionally employed to make coins, but modern coins are typically made from other metals. To be useful for coinage, a metal, or combination of metals (called an *alloy*), must be corrosion resistant. It must also be hard enough to withstand rough usage and yet be of a consistency that permits machines to accurately stamp the coins. Some metals that might otherwise make fine coins—for example, manganese (Mn)—are ruled out because they make the coins too hard to stamp. A third criterion is that the value of the metal in the coin should not be as great as the face value of the coin. For example, if pennies were made today from pure copper, the metal would be worth more than a penny, thus inviting smelters to melt down the coins for the value of the metal. Pennies today are largely made of zinc with a copper cladding.

A photo of a silver dollar from the president series.

One of the traditional alloys for making coins is a mixture of copper and nickel. Today only the U.S. nickel is made from this alloy, called cupronickel, which consists of 75% copper and 25% nickel. The modern U.S. dollar coin, often referred to as the silver dollar, doesn't contain any silver. It consists of copper (88.5%), zinc (6.0%), manganese (3.5%), and nickel (2.0%). In 2007 the U.S. Congress created a new series of $1 coins honoring former U.S. presidents. The coins have not been popular, and supplies stockpiled. The U.S. Treasury secretary suspended further production of the coins in December 2011.

GO FIGURE

Name two ways in which the metals shown in Figure 2.15 differ in general appearance from the nonmetals.



▲ **Figure 2.15 Examples of metals and nonmetals.**

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 2.14 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 metals in appearance (▲ Figure 2.15) 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 nonmetals. These elements are often referred to as **metalloids**.

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?

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

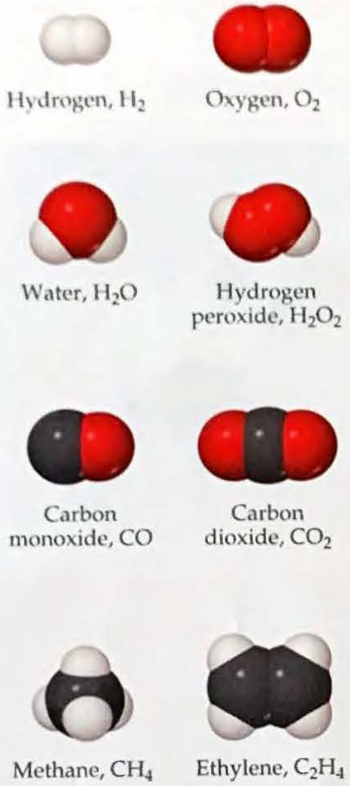
Elements in the same group of the periodic table are most likely to exhibit similar properties. We therefore expect Ca and Mg to be most alike because they are in the same group (2A, the alkaline earth metals).

Practice Exercise 1

A biochemist who is studying the properties of certain sulfur (S)-containing compounds in the body wonders whether trace amounts of another nonmetallic element might have similar behavior. To which element should she turn her attention? **(a)** O, **(b)** As, **(c)** Se, **(d)** Cr, **(e)** P.

Practice Exercise 2

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



▲ **Figure 2.16 Molecular models.** Notice how the chemical formulas of these simple molecules correspond to their compositions.

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 2.16**. 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.*

Molecular and Empirical Formulas

Chemical formulas that indicate the actual numbers of atoms in a molecule are called **molecular formulas**. (The formulas in Figure 2.16 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.

Give It Some Thought

Consider the following four formulas: SO_2, B_2H_6, CH, $C_4H_2O_2$. Which of these formulas could be (**a**) only an empirical formula, (**b**) only a molecular formula, (**c**) either a molecular or an empirical formula?

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 1

Tetracarbon dioxide is an unstable oxide of carbon with the following molecular structure:

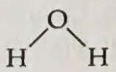
What are the molecular and empirical formulas of this substance? (a) C_2O_2, CO_2, (b) C_4O, CO, (c) CO_2, CO_2, (d) C_4O_2, C_2O, (e) C_2O, CO_2.

Practice Exercise 2

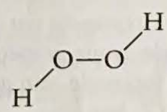
Give the empirical formula for *decaborane*, whose molecular formula is $B_{10}H_{14}$.

Picturing Molecules

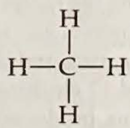
The molecular formula of a substance summarizes the composition of the substance but does not show how the atoms are joined in the molecule. A **structural formula** shows which atoms are attached to which, as in the following examples:



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.

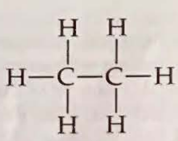
A structural formula usually does not depict the actual geometry of the molecule, that is, the actual angles at which atoms are joined. A structural formula can be written as a *perspective drawing* (**Figure 2.17**), however, to portray the three-dimensional shape.

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 a molecule (Figure 2.17). 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 a molecule would look like if the atoms were scaled up in size (Figure 2.17). 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?

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 1

Tetracarbon dioxide is an unstable oxide of carbon with the following molecular structure:

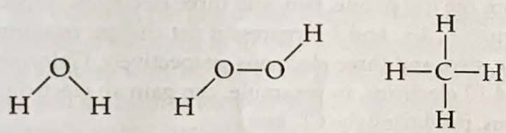
What are the molecular and empirical formulas of this substance? (a) C_2O_2, CO_2, (b) C_4O, CO, (c) CO_2, CO_2, (d) C_4O_2, C_2O, (e) C_2O, CO_2.

Practice Exercise 2

Give the empirical formula for *decaborane*, whose molecular formula is $B_{10}H_{14}$.

Picturing Molecules

The molecular formula of a substance summarizes the composition of the substance but does not show how the atoms are joined 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.

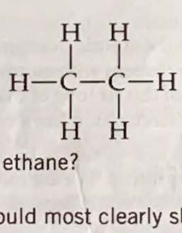
A structural formula usually does not depict the actual geometry of the molecule, that is, the actual angles at which atoms are joined. A structural formula can be written as a *perspective drawing* (**Figure 2.17**), however, to portray the three-dimensional shape.

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 a molecule (Figure 2.17). 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 a molecule would look like if the atoms were scaled up in size (Figure 2.17). 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

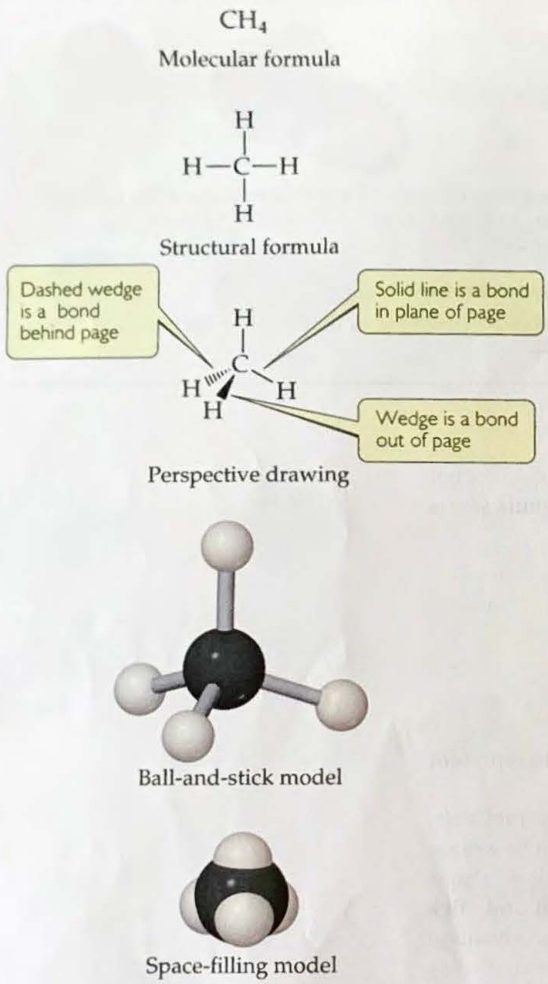
The structural formula for ethane is



(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?

GO FIGURE

Which model, the ball-and-stick or the space-filling, more effectively shows the angles between bonds around a central atom?

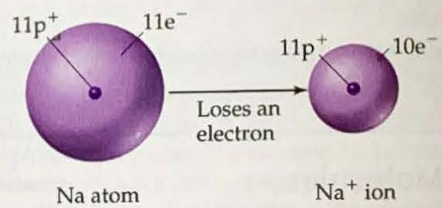


▲ **Figure 2.17 Different representations of the methane (CH_4) molecule.** Structural formulas, perspective drawings, ball-and-stick models, and space-filling models.

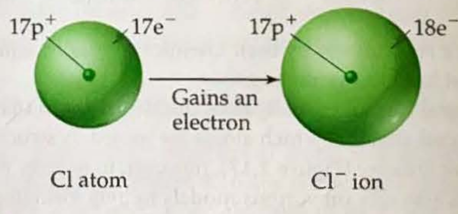
2.7 | Ions and Ionic Compounds

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+.



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; and **(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+ and is designated $^{48}Ti^{3+}$.

(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 1

In which of the following species is the number of protons less than the number of electrons? (a) Ti^{2+}, (b) P^{3-}, (c) Mn, (d) Se_4^{2-}, (e) Ce^{4+}.

Practice Exercise 2

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

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), which consist of atoms joined as in a molecule, but carrying a net positive or negative charge. Polyatomic ions will be discussed 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. The addition or removal of one or more electrons produces a charged species with behavior very different from that of its associated atom or group of atoms.

Predicting Ionic Charges

As noted in Table 2.3, the elements of group 8A are called the noble-gas elements. The noble gases are chemically nonreactive elements that form very few compounds. 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. We might deduce that atoms tend to acquire the electron arrangements of the noble gases because these 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. This simple observation will be helpful for now to account for the formation of ions. A deeper explanation awaits us in 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 barium and oxygen 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 1

Although it is helpful to know that many ions have the electron arrangement of a noble gas, many elements, especially among the metals, form ions that do not have a noble-gas electron arrangement. Use the periodic table, Figure 2.14, to determine which of the following ions has a noble-gas electron arrangement, and which do not. For those that do, indicate the noble-gas arrangement they match: (a) Ti^{4+}, (b) Mn^{2+}, (c) Pb^{2+}, (d) Te^{2-}, (e) Zn^{2+}.

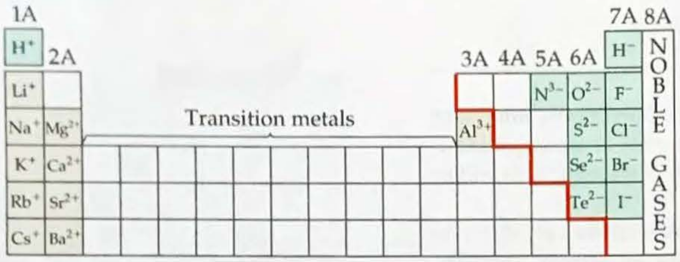
Practice Exercise 2

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

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 2.18** 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. (As noted in Practice Exercise 1 of Sample Exercise 2.8, many of the other groups do not lend themselves to such simple rules.)

GO FIGURE

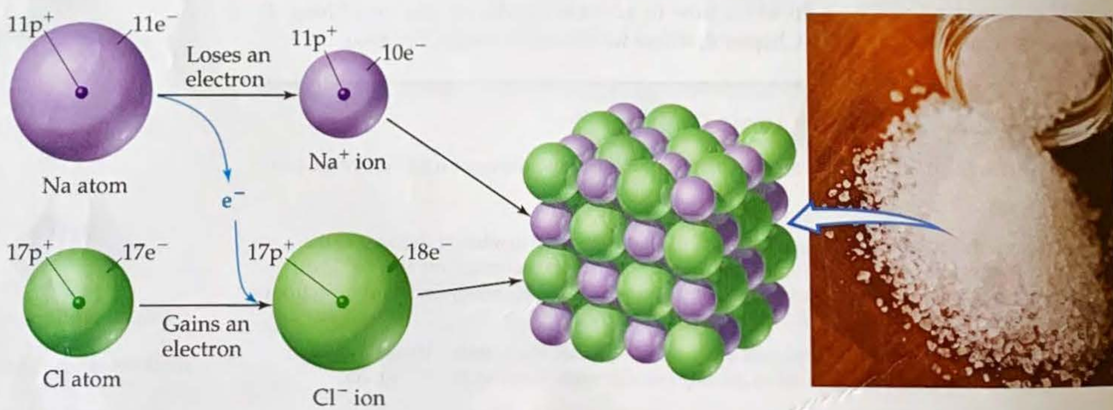
The most common ions for silver, zinc, and scandium are Ag^+, Zn^{2+}, and Sc^{3+}. Locate the boxes in which you would place these ions in this table. Which of these ions has the same number of electrons as a noble-gas element?



▲ **Figure 2.18 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.

Ionic Compounds

A great deal of chemical activity involves the transfer of electrons from one substance to another. ▼ **Figure 2.19** 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 charges 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.



▲ **Figure 2.19 Formation of an ionic compound.** 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. These ions are arranged in a lattice in solid sodium chloride, NaCl.

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.

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 1

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

Practice Exercise 2

Give a reason why each of the following statements is a safe prediction:

(a) Every compound of Rb with a nonmetal is ionic in character.
(b) Every compound of nitrogen with a halogen element is a molecular compound.
(c) The compound $MgKr_2$ does not exist.
(d) Na and K are very similar in the compounds they form with nonmetals.
(e) If contained in an ionic compound, calcium (Ca) will be in the form of the doubly charged ion, Ca^{2+}.

The ions in ionic compounds are arranged in three-dimensional structures, as Figure 2.19 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. Because chemical compounds are always electrically neutral, 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^- in NaCl, one Ba^{2+} to two Cl^- in $BaCl_2$, 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:

$$Mg^{2+} \quad N^{3-} \longrightarrow Mg_3N_2$$

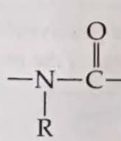
Give It Some Thought

Can you tell from the formula of a substance whether it is ionic or molecular in nature? Why or why not?

Chemistry and Life

Elements Required by Living Organisms

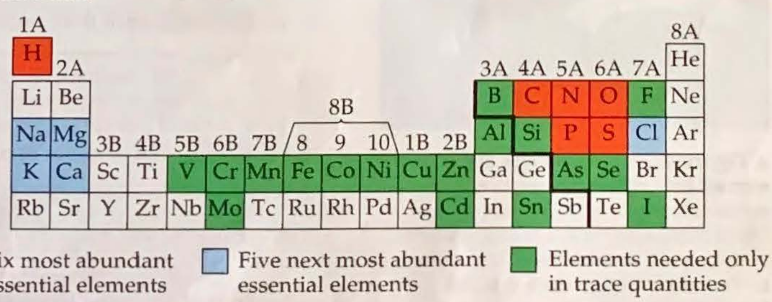
The elements essential to life are highlighted in color in ▼ Figure 2.20. More than 97% of the mass of most organisms is made up of just six of these elements—oxygen, carbon, hydrogen, nitrogen, phosphorus, and sulfur. Water is the most common compound in living organisms, accounting for at least 70% 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 carbon-based 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 other 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: 2.102



▲ **Figure 2.20 Elements essential to life.**

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 empirical formula $AlCl_3$.

(b) Two Al^{3+} ions are required to balance the charge of three O^{2-} ions. A 2:3 ratio is needed to balance the total positive charge of 6+ and the total negative charge of 6−. The empirical 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 1

For the following ionic compounds formed with S^{2-}, what is the empirical formula for the positive ion involved? **(a)** MnS, **(b)** Fe_2S_3, **(c)** MoS_2, **(d)** K_2S, **(e)** Ag_2S.

Practice Exercise 2

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-}.

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, one that conveys 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.

GO FIGURE

Is the difference in properties we see between the two substances in Figure 2.21 a difference in physical or chemical properties?

▲ **Figure 2.21 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.

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 2.21**).

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 (as indicated on the periodic table on the front inside cover of this book). 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.

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	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

Table 2.4 Common Cations[a]

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**

[a]The ions we use most often in this course are in boldface. Learn them first.

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 in this text 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 named without a Roman numeral, calcium oxide?

(b) What does the *-ium* ending on the name *ammonium* tell you about the composition of the ion?

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

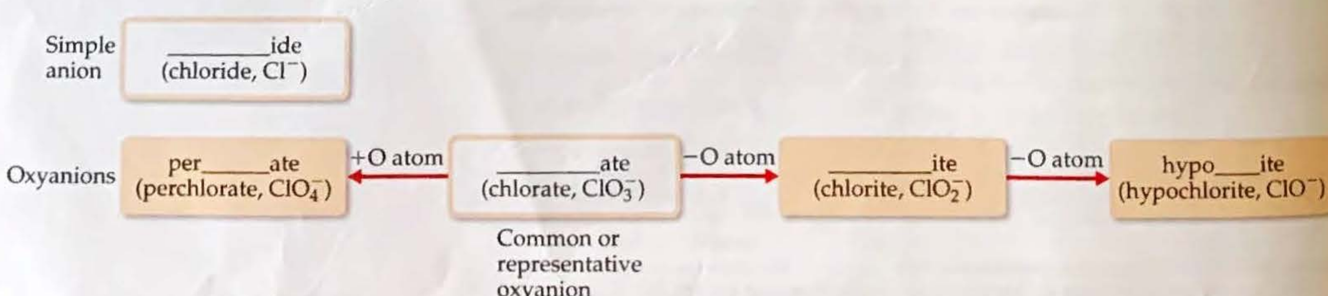
Prefixes are used when the series of oxyanions of an element extends to four members, as with the halogens. The prefix *per-* indicates one more O atom than the oxyanion ending in *-ate*; *hypo-* indicates one O atom fewer than the oxyanion ending in *-ite*:

ClO_4^-	perchlorate ion (one more O atom than chlorate)
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 2.22.

GO FIGURE

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



▲ **Figure 2.22 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 three 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 four O atoms in period 3.

Charges increase right to left.

▲ **Figure 2.23 Common oxyanions.** The composition and charges of common oxyanions are related to their location in the periodic table.

Give It Some Thought

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

▲ **Figure 2.23** 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 2.23, 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 2.23 ends in *-ate*, the ClO_4^- ion also has a *per-* prefix.

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 2.23.

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 1

Which of the following oxyanions is incorrectly named? **(a)** ClO_2^-, chlorate; **(b)** IO_4^-, periodate; **(c)** SO_3^{2-}, sulfite; **(d)** IO_3^-, iodate; **(e)** SeO_4^{2-}, selenate.

Practice Exercise 2

The formula for the bromate ion is analogous to that for the chlorate ion. Write the formula for the hypobromite and bromite ions.

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

Table 2.5 Common Anions[a]

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**

[a]The ions we use most often are in boldface. Learn them first.

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 1

Which of the following ionic compounds is incorrectly named? **(a)** $Zn(NO_3)_2$, zinc nitrate; **(b)** $TeCl_4$, tellurium(IV) chloride; **(c)** Fe_2O_3, diiron oxide; **(d)** BaO, barium oxide; **(e)** $Mn_3(PO_4)_2$, manganese (II) phosphate.

Practice Exercise 2

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

Give It Some Thought

Calcium bicarbonate is also called calcium hydrogen carbonate. **(a)** Write the formula for this compound, **(b)** predict the formulas for potassium bisulfate and lithium dihydrogen phosphate.

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.

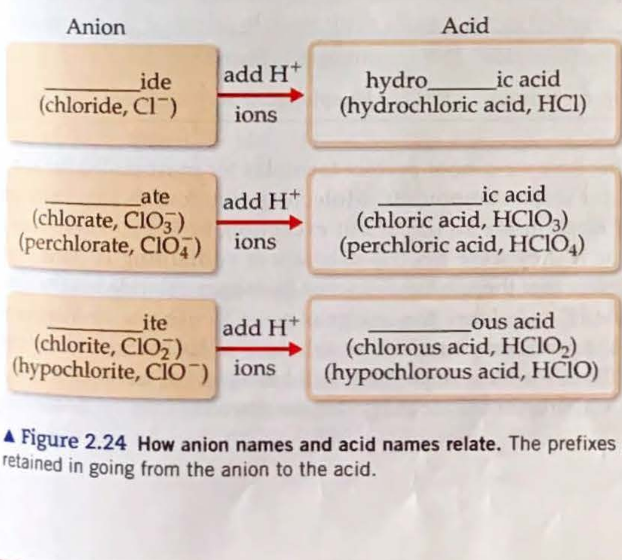
An acid is composed of an anion connected to enough H^+ ions to neutralize, or balance, the anion's charge. Thus, the SO_4^{2-} ion requires two H^+ ions, forming H_2SO_4. The name of an acid is related to the name of its anion, as summarized in ▼ **Figure 2.24.**

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)



▲ **Figure 2.24 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 2.13 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 1

Which of the following acids are incorrectly named? For those that are, provide a correct name or formula. **(a)** hydrocyanic acid, HCN; **(b)** nitrous acid, HNO_3; **(c)** perbromic acid, $HBrO_4$; **(d)** iodic acid, HI; **(e)** selenic acid, $HSeO_4$.

Practice Exercise 2

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

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 one closer to the bottom of the table is named first.*
3. *The name of the second element is given an* -ide *ending.*
4. *Greek prefixes* (◄ **Table 2.6**) indicate the number of atoms of each element. (Exception: 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.

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*

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. The distinction, which is important, will be explained in Section 4.1.) Similarly, the formula for hydrogen sulfide is H_2S because two H^+ ions are needed to balance the charge on S^{2-}.