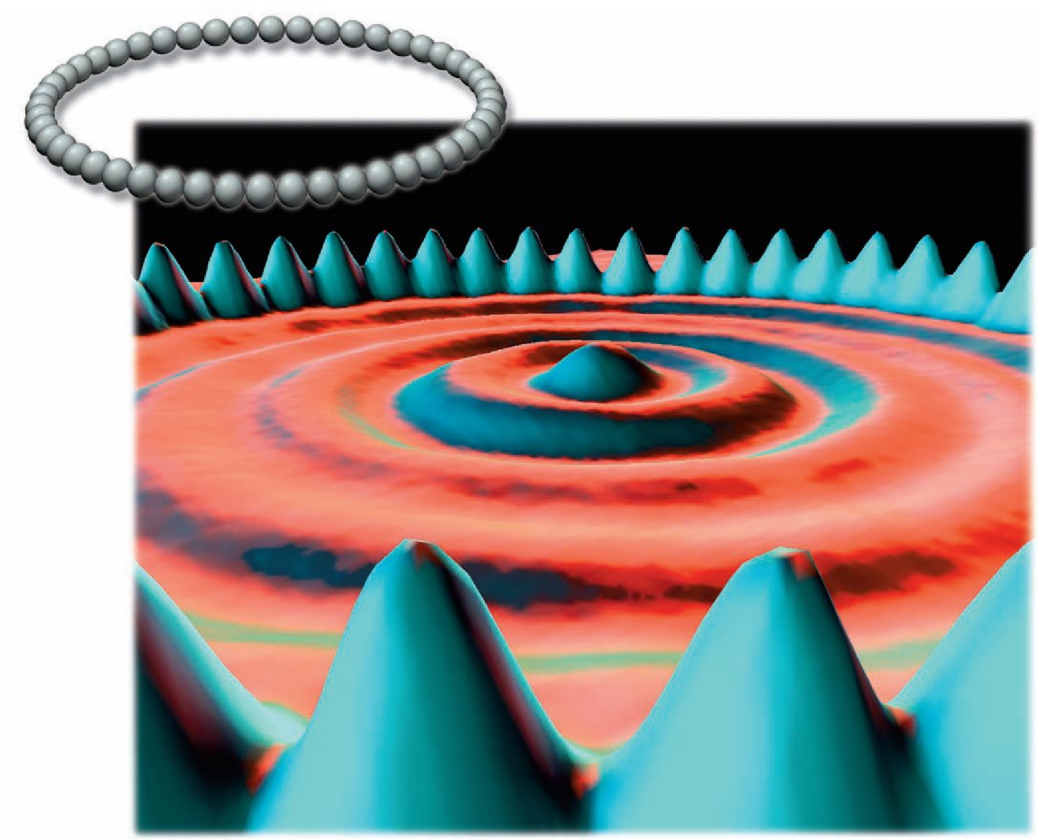
CHAPTER

2 ATOMS, MOLECULES, AND IONS



A CIRCLE OF INIDIVUAL IRON ATOMS on a copper surface, as viewed by a technique known as scanning tunneling microscopy (STM). The image is artificially colored to enhance it. The shapes of the iron atoms in the STM image are distorted, and the atoms of the copper surface are not revealed.

WHAT'S AHEAD

LOOK AROUND YOU. Notice the great variety of colors, textures, and other properties in the materials that surround you—the colors in a garden scene, the texture of the fabric in your clothes, the solubility of sugar in a cup of coffee, the transparency of a window. The materials in our world exhibit a striking and seemingly infinite variety.

We can classify properties in different ways, but how do we understand and explain them? What makes diamonds transparent and hard, while 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.

Remarkably, the diversity of these properties we see around us results from only about 100 different elements and therefore 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 how do atoms combine with one another? What rules govern the ways in which atoms can 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?

The chapter-opening photograph is an image of a circle of 48 iron atoms arranged on a copper metal surface. The diameter of the circle is about 1/20,000 the diameter of a human hair. Atoms are indeed very tiny entities.

This very striking image reveals the power of modern experimental methods to identify individual atoms, but it does not reveal the structures of the atoms themselves. Fortunately, we can use a variety of experimental techniques to probe the atom to gain a clearer understanding of what it is like. In this chapter we begin to explore the fascinating world of atoms that we discover by such experiments. We will examine the basic structure of the atom and briefly 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 have speculated about the nature of the fundamental "stuff" from which the world is made. Democritus (460–370 BC) and other early Greek philosophers thought that the material world must be made up of tiny indivisible particles that they called *atomos*, meaning "indivisible or uncuttable." Later, Plato and Aristotle formulated the notion that there can be no ultimately indivisible particles. 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, when scientists tried to explain the properties of gases. Air is composed of something invisible and in constant motion; we can feel the motion of the wind against us, for example. It is natural to think of tiny invisible particles as giving rise to these familiar effects. Isaac Newton (1642–1727), the most famous scientist of his time, favored the idea of atoms. But thinking of atoms as invisible particles in air is very different from thinking of atoms as the fundamental building blocks of elements.

▲ **Figure 2.1 John Dalton (1766–1844).** Dalton was the son of a poor English weaver. He began teaching at the age of 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.

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 an English schoolteacher, John Dalton (Figure 2.1 ◄), during the period from 1803–1807. Dalton's atomic theory involved the following postulates:

1. Each element is composed of extremely small particles called atoms.
2. All atoms of a given element are identical to one another in mass and other properties, but the atoms of one element are different from the atoms of all other elements.
3. The atoms of one element cannot be changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions.
4. Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

According to Dalton's atomic theory, **atoms** are the smallest particles of an element that retain the chemical identity of the element. (Section 1.1) As noted in the postulates of Dalton's theory, an element is composed of only one kind of atom. A compound, in contrast, contains atoms of two or more elements.

Dalton's theory explains several simple laws of chemical combination that were known during his time. One of these laws was the *law of constant composition*: In a given compound, the relative numbers and kinds of atoms are constant. (Section 1.2) This law is the basis of Dalton's Postulate 4. Another fundamental chemical law was the *law of conservation of mass* (also known as the *law of conservation of matter*): The total mass of materials present after a chemical reaction is the same as the total mass present before the reaction. This law is the basis for Postulate 3. Dalton proposed that atoms always retain their identities and that atoms taking part in a chemical reaction rearrange to give new chemical combinations.

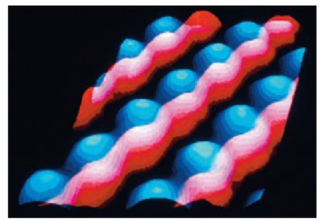
A good theory should explain the known facts and predict new ones. Dalton used his theory to deduce the *law of multiple proportions*: If two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in the ratio of small whole numbers. We can illustrate this law by considering the substances 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. In other words, the ratio of the mass of oxygen per gram of hydrogen in the two compounds is 2 : 1. Using the atomic theory, we can conclude that hydrogen peroxide contains twice as many atoms of oxygen per hydrogen atom as does water.

GIVE IT SOME THOUGHT

One compound of carbon and oxygen contains 1.333 g of oxygen per gram of carbon, whereas a second compound contains 2.666 g of oxygen per gram of carbon. **(a)** What chemical law do these data illustrate? **(b)** If the first compound has an equal number of oxygen and carbon atoms, what can we conclude about the composition of the second compound?

2.2 THE DISCOVERY OF ATOMIC STRUCTURE

Dalton reached his conclusion about atoms based on chemical observations in the macroscopic world of the laboratory. Neither he nor those who followed him during the century after his work was published had direct evidence for the existence of atoms. Today, however, we can use powerful instruments to measure the properties of individual atoms and even provide images of them (Figure 2.2▶).



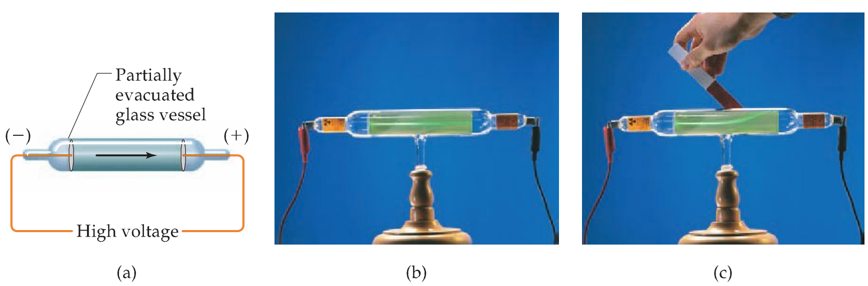
▲ Figure 2.2 **An image of the surface of the semiconductor GaAs (gallium arsenide).** This image was obtained by a technique called scanning tunneling microscopy. The color was added to the image by computer to distinguish the gallium atoms (blue spheres) from the arsenic atoms (red spheres).

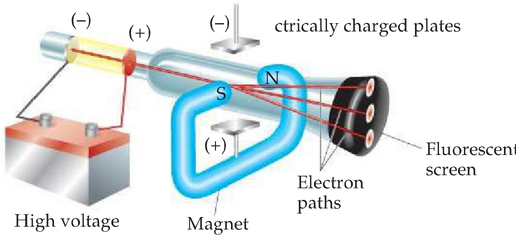
As scientists began to develop methods for more detailed probing of the nature of matter, the atom, which was supposed to be indivisible, began to show signs of a more complex structure: We now know that the atom is composed of still smaller **subatomic particles**. Before we summarize the current model of atomic structure, we will 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 a simple statement of the behavior of charged particles: *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 partially evacuated tubes (tubes that had been pumped almost empty of air), such as those shown in Figure 2.3▼. When a high voltage was applied to the

▼ Figure 2.3 **Cathode-ray tube.** (a) In a cathode-ray tube, electrons move from the negative electrode (cathode) to the positive electrode (anode). (b) A photo of a cathode-ray tube containing a fluorescent screen to show the path of the cathode rays. (c) The path of the cathode rays is deflected by the presence of a magnet.





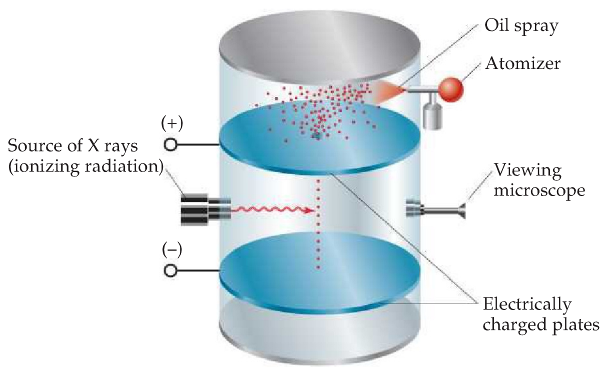
▶ **Figure 2.4 Cathode-ray tube with perpendicular magnetic and electric fields.** The cathode rays (electrons) originate from the negative plate on the left and are accelerated toward the positive plate on the right, which has a hole in its center. A narrow beam of electrons passes through the hole and is then deflected by the magnetic and electric fields. The three paths result from different strengths of the magnetic and electric fields. The charge-to-mass ratio of the electron can be determined by measuring the effects that the magnetic and electric fields have on the direction of the beam.

electrodes in the tube, radiation was produced. This radiation, called **cathode rays**, originated from the negative electrode, or cathode. Although the rays themselves could not be seen, their movement was detected because the rays cause certain materials, including glass, to *fluoresce*, or to give off light.

Scientists held conflicting views about the nature of the cathode rays. It was not initially clear whether the rays were an invisible stream of particles or a new form of radiation. 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 [Figure 2.3(c)]. The British scientist J. J. Thomson observed many properties of the cathode rays, including the fact that they are the same regardless of the identity of the cathode material. In a paper published in 1897, Thomson summarized his observations and concluded that cathode rays are streams of negatively charged particles. Thomson's paper is generally accepted as the "discovery" of what later became known as the *electron*.

Thomson constructed a cathode-ray tube having a fluorescent screen at one end, such as that shown in Figure 2.4▲, so that he could quantitatively measure the effects of electric and magnetic fields on the thin stream of electrons passing through a hole in the positively charged electrode. These measurements 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.*

Once the charge-to-mass ratio of the electron was known, measuring either the charge or the mass of an electron would yield the value of the other quantity. In 1909, Robert Millikan (1868–1953) of the University of Chicago succeeded in measuring the charge of an electron by performing a series of experiments described in Figure 2.5▼. He then calculated the mass of the electron by using



▶ **Figure 2.5 Millikan's oil-drop experiment.** A representation of the apparatus Millikan used to measure the charge of the electron. Small drops of oil, which had picked up extra electrons, were allowed to fall between two electrically charged plates. Millikan monitored the drops, measuring how the voltage on the plates affected their rate of fall. From these data he calculated the charges on the drops. His experiment showed that the charges were always integral multiples of 1.602×10^{-19} C, which he deduced was the charge of a single electron.

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

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 presently accepted value for the mass of the electron, 9.10938×10^{-28} g. This mass is about 2000 times smaller than that of hydrogen, the lightest atom.

▲ **Figure 2.6 Marie Sklodowska Curie (1867–1934).** When M. Curie presented her doctoral thesis, it was described as the greatest single contribution of any doctoral thesis in the history of science. Among other things, Curie discovered two new elements, polonium and radium. In 1903 Henri Becquerel, M. Curie, and her husband, Pierre, were jointly awarded the Nobel Prize in physics. In 1911 M. Curie won a second Nobel Prize, this time in chemistry.

Radioactivity

In 1896 the French scientist Henri Becquerel (1852–1908) was studying a uranium compound when he discovered that it 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.

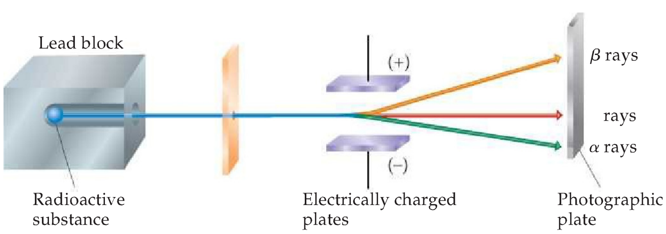
Further study of the nature of radioactivity, principally by the British scientist Ernest Rutherford (Figure 2.7▶), revealed three types of radiation: alpha (α), beta (β), and gamma (γ) radiation. Each type differs in its response to an electric field, as shown in Figure 2.8▼. The paths of both α and β radiation are bent by the electric field, although in opposite directions; γ radiation is unaffected.

Rutherford showed that both α and β rays consist of fast-moving particles, which were called α and β 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 positive charge and are attracted toward 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. We will discuss radioactivity in greater detail in Chapter 21.

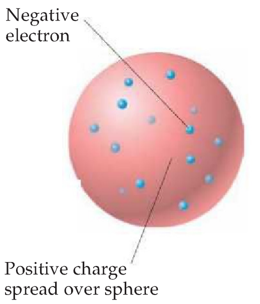
▲ **Figure 2.7 Ernest Rutherford (1871–1937).** Rutherford, whom Einstein called "the second Newton," was born and educated in New Zealand. In 1895 he was the first overseas student ever to be awarded a position at the Cavendish Laboratory at Cambridge University in England, where he worked with J. J. Thomson. In 1898 he joined the faculty of McGill University in Montreal. While at McGill, Rutherford did his research on radioactivity that led to his being awarded the 1908 Nobel Prize in chemistry. In 1907 Rutherford moved back to England to be a faculty member at Manchester University, where in 1910 he performed his famous α-particle scattering experiments that led to the nuclear model of the atom. In 1992 his native New Zealand honored Rutherford by putting his likeness, along with his Nobel Prize medal, on their $100 currency note.

The Nuclear Atom

With the 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 the mass of an atom, 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, as shown in



▲ **Figure 2.8 Behavior of alpha (α), beta (β), and gamma (γ) rays in an electric field.** The α rays consist of positively charged particles and are therefore attracted to the negatively charged plate. The β rays consist of negatively charged particles and are attracted to the positively charged plate. The γ rays, which carry no charge, are unaffected by the electric field.



▲ **Figure 2.9 J. J. Thomson's "plum-pudding" model of the atom.** Thomson pictured the small electrons to be embedded in the atom much like raisins in a pudding or seeds in a watermelon. Ernest Rutherford proved this model wrong.

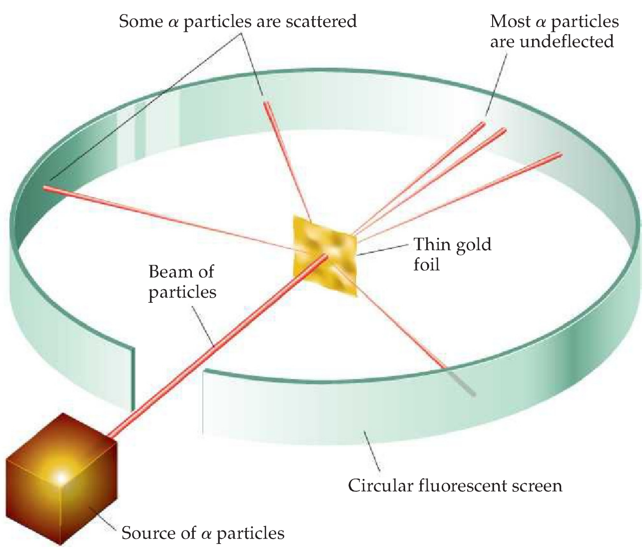
Figure 2.9◀. This so-called "plum-pudding" model, named after a traditional English dessert, was very short lived.

In 1910, Rutherford and his coworkers performed an experiment that disproved Thomson's model. Rutherford was studying the angles at which α particles were deflected, or *scattered*, as they passed through a thin gold foil only a few thousand atoms thick (Figure 2.10▼). Rutherford and his coworkers discovered that almost all the α particles passed directly through the foil without deflection. A few particles were found to be deflected by approximately 1 degree, consistent with Thomson's plum-pudding model. Just for the sake of completeness, Rutherford suggested that Ernest Marsden, an undergraduate student working in the laboratory, look for evidence of scattering at large angles. To everyone's surprise, a small amount of scattering was observed at large angles. Some particles were even 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.

By 1911, Rutherford was able to explain these observations. He postulated that most of the mass of each gold atom in his foil and all of its positive charge reside in a very small, extremely dense region, which he called the **nucleus**. He postulated further that most of the total volume of an atom is empty space in which electrons move around the nucleus. In the α-scattering experiment, most α particles passed directly through the foil because they did not encounter the minute nucleus of any gold atom; they merely passed through the empty space making up the greatest part of all the atoms in the foil. Occasionally, however, an α particle came close to a gold nucleus. The repulsion between the highly charged gold nucleus and the α particle was strong enough to deflect the less massive α particle, as shown in Figure 2.11▶.

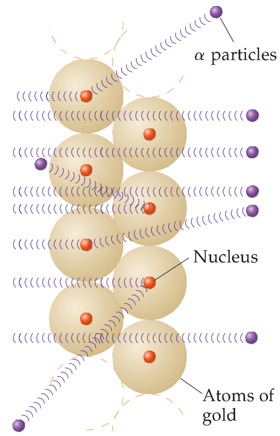
Subsequent experimental studies led to the discovery of both positive particles (*protons*) and neutral particles (*neutrons*) in the nucleus. Protons were discovered in 1919 by Rutherford. In 1932 British scientist James Chadwick (1891–1972) discovered neutrons. We examine these particles more closely in Section 2.3.

▶ **Figure 2.10 Rutherford's experiment on the scattering of α particles.** The red lines represent the paths of the α particles. When the incoming beam strikes the gold foil, most particles pass straight through the foil, but some are scattered.



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?



▲ **Figure 2.11 Rutherford's model explaining the scattering of α particles.** The gold foil is several thousand atoms thick. Because most of the volume of each atom is empty space, most α particles pass through the foil without deflection. When an α particle passes very close to a gold nucleus, however, it is repelled, causing its path to be altered.

2.3 THE MODERN VIEW OF ATOMIC STRUCTURE

Since the time of Rutherford, physicists have learned much about the detailed composition of atomic nuclei. In the course of these discoveries, the list of particles that make up nuclei has grown long and continues to increase. As chemists, however, we can take a very simple view of the atom because only three subatomic particles—the **proton, neutron,** and **electron**—have a bearing on chemical behavior.

The charge of an electron is -1.602×10^{-19} C, and that of a proton is $+1.602 \times 10^{-19}$ 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 the electron is 1−, and that of the proton is 1+. Neutrons are uncharged and are therefore 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 together in the nucleus of the atom, which, as Rutherford proposed, is extremely small. The vast majority of an atom's volume is the space in which the electrons reside. 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 between 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 instead the **atomic mass unit**, or amu.* One amu equals 1.66054×10^{-24} g. The masses of the proton and neutron are very nearly equal, and both are much greater than that of the electron: A proton has a mass of 1.0073 amu, a neutron 1.0087 amu, and an electron 5.486×10^{-4} amu. Because it would take 1836 electrons to equal the mass of 1 proton, the nucleus contains most of the mass of an atom. Table 2.1 ▼ summarizes the charges and masses of the subatomic particles. We will have more to say about atomic masses in Section 2.4.

Atoms are also extremely small. Most atoms have diameters between 1×10^{-10} m and 5×10^{-10} m, or 100–500 pm. A convenient, although non-SI, unit of length used to express atomic dimensions is the **angstrom** (Å). One

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}

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

angstrom equals 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. Both picometers and angstroms are commonly used to express the dimensions of atoms and molecules.

SAMPLE EXERCISE 2.1 | Illustrating the Size of an Atom

The diameter of a US penny is 19 mm. The diameter of a silver atom, by comparison, is only 2.88 Å. How many silver atoms could be arranged side by side in a straight line across the diameter of a penny?

SOLUTION

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

$$\text{Ag atoms} = (19\ \cancel{\text{mm}})\left(\frac{10^{-3}\ \cancel{\text{m}}}{1\ \cancel{\text{mm}}}\right)\left(\frac{1\ \cancel{\text{Å}}}{10^{-10}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{Ag atom}}{2.88\ \cancel{\text{Å}}}\right) = 6.6 \times 10^{7}\ \text{Ag atoms}$$

That is, 66 million silver atoms could sit side by side across a penny!

PRACTICE EXERCISE

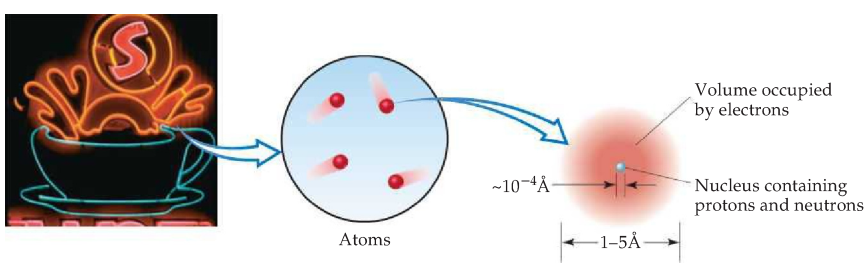
The diameter of a carbon atom is 1.54 Å. **(a)** Express this diameter in picometers. **(b)** How many carbon atoms could be aligned side by side in a straight line across the width of a pencil line that is 0.20 mm wide?
Answers: **(a)** 154 pm, **(b)** 1.3×10^{6} C atoms

The diameters of atomic nuclei are 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 incredible 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! Astrophysicists have suggested that the interior of a collapsed star may approach this density.

An illustration of the atom that incorporates the features we have just discussed is shown in Figure 2.12▼. The electrons, which take up most of the volume of the atom, 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? For example, how does an atom of carbon differ from an atom of oxygen? The significant difference is in their subatomic compositions. The atoms of each element have a characteristic number of protons. Indeed, the number



▼ **Figure 2.12 The structure of the atom.** Neon gas is composed of atoms. The nucleus, which contains protons and neutrons, is the location of virtually all the mass of the atom. The rest of the atom is the space in which the light, negatively charged electrons reside.

A Closer Look BASIC FORCES

There are four basic forces 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 and magnetic forces are intimately related. Electric forces are of fundamental importance 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 a stronger attractive force, called the *strong nuclear force*, did not keep them together. This force acts between subatomic particles, as in the nucleus. At this distance, the strong nuclear force is stronger than the 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 Exercises: 2.83(b) and 2.89

of protons in the nucleus of 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, whereas 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, and the superscript, called the **mass number**, is the total number of protons plus neutrons in the atom:

Mass number (number of protons plus neutrons)

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

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, atoms that contain six protons and eight neutrons have a mass number of 14 and are represented as $^{14}_{6}C$ or ^{14}C and referred to as carbon-14.

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

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.

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, the sum of the number of protons plus the number of neutrons. According to the list of elements given inside the front cover, gold has an atomic number of 79. Consequently, an atom of ^{197}Au has 79 protons, 79 electrons, and $197 - 79 = 118$ neutrons. **(b)** The atomic number of strontium (listed inside the 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 of them. **(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 will discuss the mass scale used for atoms and introduce the concept of *atomic weights*. In Section 3.3 we will extend these concepts to show how atomic masses are used to determine the masses of compounds and *molecular weights*.

The Atomic Mass Scale

Although scientists of the nineteenth century knew nothing about subatomic particles, they 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* (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.

Average Atomic Masses

Most elements occur in nature as mixtures of isotopes. We can determine the *average atomic mass* of an element by using the masses of its various isotopes and their relative abundances. 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. We calculate the average atomic mass of carbon from the fractional abundance of each isotope and the mass of that isotope:

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

The average atomic mass of each element (expressed in atomic mass units) is also known as its **atomic weight**. Although the term *average atomic mass* is more proper, the term *atomic weight* is more common. 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, which has an atomic mass of 34.969 amu, and 24.22% ^{37}Cl, which has an atomic mass of 36.966 amu. Calculate the average atomic mass (that is, the atomic weight) of chlorine.

SOLUTION

We can calculate the average atomic mass 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

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

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

PRACTICE EXERCISE

Three isotopes of silicon occur in nature: ^{28}Si (92.23%), which has an atomic mass of 27.97693 amu; ^{29}Si (4.68%), which has an atomic mass of 28.97649 amu; and ^{30}Si (3.09%), which has an atomic mass of 29.97377 amu. Calculate the atomic weight of silicon.
Answer: 28.09 amu

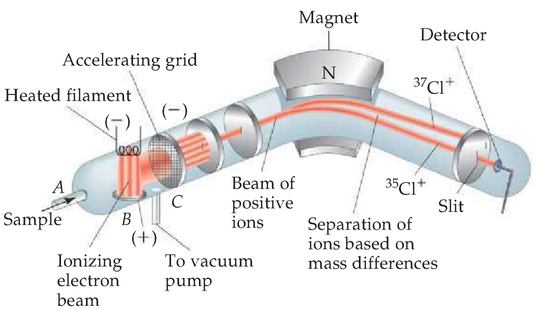
A Closer Look THE MASS SPECTROMETER

The most direct and accurate means for determining atomic and molecular weights is provided by the **mass spectrometer** (Figure 2.13▼). 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, mostly with a 1+ charge. These charged particles are accelerated toward a negatively charged wire 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, much as electrons are deflected by a magnetic field (Figure 2.4). For charged 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 negatively charged grid, charged particles of various masses can be selected to enter the detector at the end of the instrument.

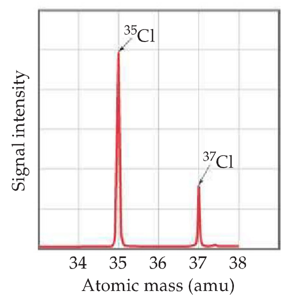
A graph of the intensity of the detector signal versus particle atomic mass is called a *mass spectrum*. The mass spectrum of chlorine atoms, shown in Figure 2.14▼, reveals the presence of two isotopes. Analysis of a mass spectrum gives both the masses of the charged particles reaching the detector and their relative abundances. The abundances 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, 2.93, and 2.94



▲ **Figure 2.13 A mass spectrometer.** Cl atoms are introduced on the left side of the spectrometer and are ionized to form Cl^+ ions, which are then directed through a magnetic field. The paths of the ions of the two isotopes of Cl diverge as they pass through the magnetic field. As drawn, the spectrometer is tuned to detect $^{35}Cl^+$ ions. The heavier $^{37}Cl^+$ ions are not deflected enough for them to reach the detector.

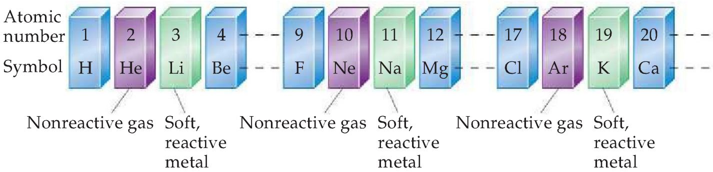


▲ **Figure 2.14 Mass spectrum of atomic chlorine.** The fractional abundances of the ^{35}Cl and ^{37}Cl isotopes of chlorine are indicated by the relative signal intensities of the beams reaching the detector of the mass spectrometer.

2.5 THE PERIODIC TABLE

Dalton's atomic theory set the stage for a vigorous growth in chemical experimentation during the early 1800s. As the body of chemical observations grew and the list of known elements expanded, attempts were made to find regular 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 very strong similarities to one another. The elements lithium (Li), sodium (Na), and potassium (K) are all soft, very reactive metals,



◀ **Figure 2.15** **Arranging the elements by atomic number reveals a periodic pattern of properties.** This periodic pattern is the basis of the periodic table.

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 2.15▲.

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**. The periodic table is shown in Figure 2.16▼ and is also given on the front inside cover of the text. For each element in the table, the atomic number and atomic symbol are given. The atomic weight is often given as well, as in the following 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.

	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	113	114	115	116		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 2.16** **Periodic table of the elements.** Different colors are used to show the division of the elements into metals, metalloids, and nonmetals.

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, which begin with lithium (Li) and sodium (Na), respectively, 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 (those with atomic numbers 57–70) appear at the bottom of the table. The seventh and last period is incomplete, but it also has 14 of its members placed in a row at the bottom of the table.

The vertical columns of the periodic table are called **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 2.16. 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 the lower set of labels at the top of the table in Figure 2.16. We will use the traditional North American convention with Arabic numerals.

Elements that belong to the same 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. As their name suggests, the coinage metals are used throughout the world to make coins. Many other groups in the periodic table also have names, as listed in Table 2.3▼.

We will learn in Chapters 6 and 7 that the elements in a group of the periodic table 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 the behaviors of elements and to aid in remembering many facts. You will find it helpful to refer to the periodic table frequently when studying the remainder of this chapter.

▲ **Figure 2.17 Some familiar examples of metals and nonmetals.** The nonmetals (from bottom left) are sulfur (yellow powder), iodine (dark, shiny crystals), bromine (reddish brown liquid and vapor in glass vial), and three samples of carbon (black charcoal powder, diamond, and graphite in the pencil lead). The metals are in the form of an aluminum wrench, copper pipe, lead shot, silver coins, and gold nuggets.

Except for hydrogen, all the elements on the left side and in the middle of the periodic table are **metallic elements**, or **metals**. The majority of elements are metallic; they all share characteristic properties, such as luster and high electrical and heat conductivity. All metals, with the exception of mercury (Hg), are solids at room temperature. The metals are separated from the **nonmetallic elements**, or **nonmetals**, by a diagonal steplike line that runs from boron (B) to astatine (At), as shown in Figure 2.16. Hydrogen, although on the left side of the periodic 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 2.17◀) and in other physical properties. Many of the elements that lie along the line that separates metals from nonmetals, such as antimony (Sb), have properties that fall between those of metals and those of nonmetals. These elements are often referred to as **metalloids**.

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. Seaborg (Figure 2.18 ▶) became a faculty member in the chemistry department at the University of California, Berkeley in 1937. In 1940 he and his colleagues Edwin McMillan, Arthur Wahl, and Joseph Kennedy 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.

During the period 1944 through 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 U.S. 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. Another team at Berkeley corroborated that discovery in 1993. 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," with a proposed symbol of Sg. After several years of controversy about whether an element should be named after a living person, the IUPAC officially adopted the name seaborgium in 1997. Seaborg became the first person to have an element named after him while he was still alive.

◀ **Figure 2.18 Glenn Seaborg (1912–1999).** The photograph shows Seaborg at Berkeley in 1941 using a Geiger counter to try to detect radiation produced by plutonium. Geiger counters will be discussed in Section 21.5.

Related Exercise: 2.96

GIVE IT SOME THOUGHT

Chlorine is a halogen. Locate this element in the periodic table. **(a)** What is its symbol? **(b)** In what period and in what group is the element located? **(c)** What is its atomic number? **(d)** Is chlorine a metal or nonmetal?

SAMPLE EXERCISE 2.5 | Using the Periodic Table

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

SOLUTION

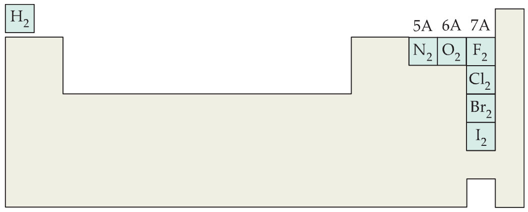
Elements that are in the same group of the periodic table are most likely to exhibit similar chemical and physical properties. We therefore expect that Ca and Mg should be most alike because they are in the same group (2A, the alkaline earth metals).

PRACTICE EXERCISE

Locate Na (sodium) and Br (bromine) on the periodic table. Give the atomic number of each, and label each a metal, metalloid, or nonmetal.
Answer: Na, atomic number 11, is a metal; Br, atomic number 35, is a nonmetal.

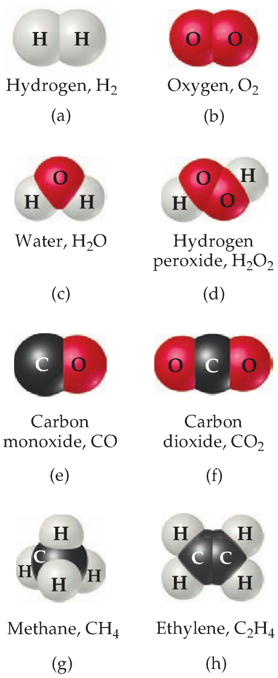
2.6 MOLECULES AND MOLECULAR COMPOUNDS

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, both of which are formed from atoms. We examine molecules here and ions in Section 2.7.



▶ **Figure 2.19 Diatomic molecules.** Seven common elements exist as diatomic molecules at room temperature.

A **molecule** is an assembly of two or more atoms tightly bound together. The resultant "package" of atoms behaves in many ways as a single, distinct object, just as a cell phone composed of many parts can be recognized as a single object. We will discuss the forces that hold the atoms together (the chemical bonds) in Chapters 8 and 9.



▲ **Figure 2.20 Molecular models of some simple molecules.** Notice how the chemical formulas of these substances correspond to their compositions.

Molecules and Chemical Formulas

Many elements are found in nature in molecular form; that is, two or more of the same type of atom are bound together. For example, the oxygen normally found in air consists of molecules that contain two oxygen atoms. We represent this molecular form of oxygen by the **chemical formula** O_2 (read "oh two"). The subscript in the formula tells us that two oxygen atoms are present in each molecule. A molecule that is 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 for this substance 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. Their locations in the periodic table are shown in Figure 2.19▲. When we speak of the substance hydrogen, we mean H_2 unless we explicitly indicate otherwise. Likewise, when we speak of oxygen, nitrogen, or any of the halogens, we are referring to O_2, N_2, F_2, Cl_2, Br_2, or I_2. Thus, the properties of oxygen and hydrogen listed in Table 1.3 are those of O_2 and H_2. Other, less common forms of these elements behave much differently.

Compounds that are composed of molecules contain more than one type of atom and are called **molecular compounds**. A molecule of water, for example, consists of two hydrogen atoms and one oxygen atom and is therefore represented by the chemical formula H_2O. Lack of a subscript on the O indicates one atom of O per water molecule. Another compound composed of these same elements (in different relative proportions) is hydrogen peroxide, H_2O_2. The properties of hydrogen peroxide are very different from the properties of water.

Several common molecules are shown in Figure 2.20◀. Notice how the composition of each compound is given by its chemical formula. Notice also that these substances are composed only of nonmetallic elements. *Most molecular substances that we will encounter contain only nonmetals.*

Molecular and Empirical Formulas

Chemical formulas that indicate the actual numbers and types of atoms in a molecule are called **molecular formulas**. (The formulas in Figure 2.20 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.

Molecular formulas provide more information about molecules than do empirical formulas. 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, such as the most common forms of elemental carbon, that do not exist as isolated molecules. For these substances, we must rely on empirical formulas. Thus, all the common forms of elemental carbon are represented by the element's chemical symbol, C, which is the empirical formula for all the forms.

SAMPLE EXERCISE 2.6 | Relating Empirical and Molecular Formulas

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

SOLUTION

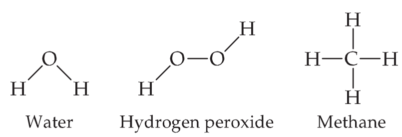
(a) The subscripts of an empirical formula are the smallest whole-number ratios. The smallest ratios are obtained by dividing each subscript by the largest common factor, in this case 6. The resultant empirical formula for glucose is CH_2O.
(b) Because the subscripts in N_2O are already the lowest integral numbers, the empirical formula for nitrous oxide is the same as its molecular formula, N_2O.

PRACTICE EXERCISE

Give the empirical formula for the substance called *diborane*, whose molecular formula is B_2H_6.
Answer: BH_3

Picturing Molecules

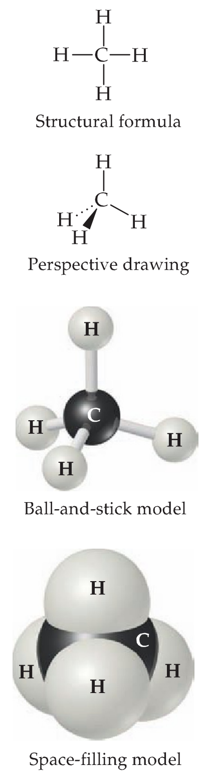
The molecular formula of a substance summarizes the composition of the substance but does not show how the atoms come together to form the molecule. The **structural formula** of a substance shows which atoms are attached to which within the molecule. For example, the structural formulas for water, hydrogen peroxide, and methane (CH_4) can be written as follows:



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*, however, to give some sense of three-dimensional shape, as shown in Figure 2.21 ▶.

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 within the molecule (Figure 2.21). In a ball-and-stick



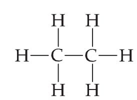
▲ **Figure 2.21 Different representations of the methane (CH_4) molecule.** Structural formulas, perspective drawings, ball-and-stick models, and space-filling models each help us visualize the ways atoms are attached to each other in molecules. In the perspective drawing, solid lines represent bonds in the plane of the paper, the solid wedge represents a bond that extends out from the plane of the paper, and dashed lines represent bonds behind the paper.

model, balls of the same size may represent all atoms, or the relative sizes of the balls may reflect the relative sizes of the atoms. 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 2.21). 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 their colors, but they may also be labeled with the element's symbol.

GIVE IT SOME THOUGHT

The structural formula for the substance ethane is shown here:

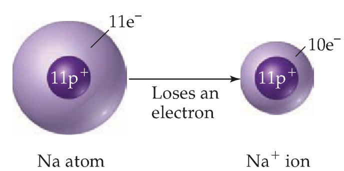


(a) What is the molecular formula for ethane? **(b)** What is its empirical formula? **(c)** Which kind of molecular model would most clearly show the angles between atoms?

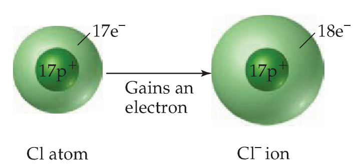
2.7 IONS AND IONIC COMPOUNDS

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 a neutral atom, a charged particle called an **ion** is formed. An ion with a positive charge is called a **cation** (pronounced CAT-ion); a negatively charged ion is called 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, whereas nonmetal atoms tend to gain electrons to form anions.

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

Give the chemical symbol, including mass number, for each of the following ions: **(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 (22) is the atomic number of the element. By referring to a periodic table or list of elements, we see that the element with atomic number 22 is titanium (Ti). The mass number of this isotope of titanium is $22 + 26 = 48$ (the sum of the protons and neutrons). Because the ion has three more protons than electrons, it has a net charge of 3+. Thus, the symbol for the ion is $^{48}Ti^{3+}$.

(b) By referring to a periodic table or a table of elements, we see that sulfur (S) has an atomic number of 16. Thus, each atom or ion of sulfur must contain 16 protons. We are told that the ion also has 16 neutrons, meaning the mass number of the ion is $16 + 16 = 32$. Because the ion has 16 protons and 18 electrons, its net charge is 2−. Thus, the symbol for the ion 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 will consider further examples of 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. The difference is like the change from Dr. Jekyll to Mr. Hyde: 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 the 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. The members of the noble-gas family are chemically very 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 the neutral neon atom (atomic number 10). Similarly, when chlorine gains an electron, it ends up with 18, the same number of electrons as in argon (atomic number 18). 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 the Charges of Ions**

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

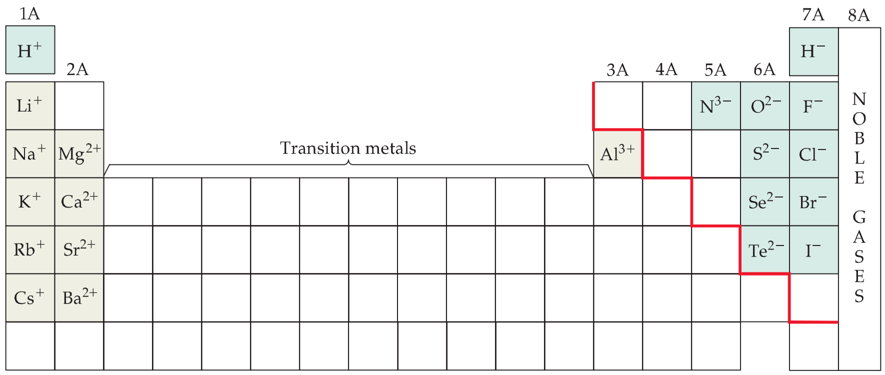
SOLUTION

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 of its 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, thereby 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−



▲ **Figure 2.22 Charges of some common ions.** Notice that the steplike line that divides metals from nonmetals also separates cations from anions. Hydrogen forms both 1+ and 1− ions.

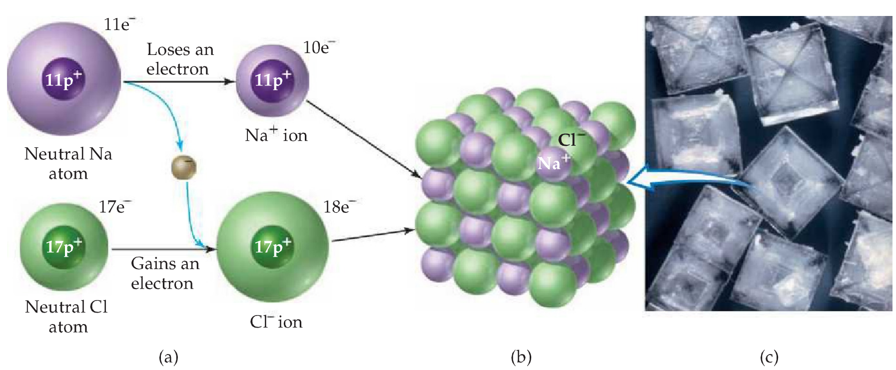
The periodic table is very useful for remembering the charges of ions, especially those of the elements on the left and right sides of the table. As Figure 2.22▲ shows, the charges of these ions relate in a simple way to their positions in the table. On the left side of the table, for example, the group 1A elements (the alkali metals) form 1+ ions, and the group 2A elements (the alkaline earths) form 2+ ions. On the other side of the table the group 7A elements (the halogens) form 1− ions, and the group 6A elements form 2− ions. As we will see later in the text, 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. As we just saw, ions form when one or more electrons transfer from one neutral atom to another. Figure 2.23▼ shows that when elemental sodium is allowed to react with elemental chlorine, an electron transfers from a neutral sodium atom to a neutral chlorine atom. We are left with 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 that contains both positively and negatively charged ions.

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, whereas anions are nonmetal ions. Consequently, *ionic compounds*

▼ **Figure 2.23 The formation of an ionic compound.** (a) The transfer of an electron from a neutral Na atom to a neutral 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.



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 the following compounds would you expect to be ionic: N_2O, Na_2O, $CaCl_2$, SF_4?

SOLUTION

We would predict that Na_2O and $CaCl_2$ are ionic compounds because they are composed of a metal combined with a nonmetal. The other two compounds, composed entirely of nonmetals, are predicted (correctly) to be molecular compounds.

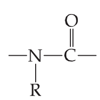
PRACTICE EXERCISE

Which of the following 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. The arrangement of Na^+ and Cl^- ions in NaCl is shown in Figure 2.23. Because there is no discrete molecule of NaCl, we are able to write only an empirical formula for this substance. In fact, only empirical formulas can be written for most ionic compounds.

Chemistry and Life — ELEMENTS REQUIRED BY LIVING ORGANISMS

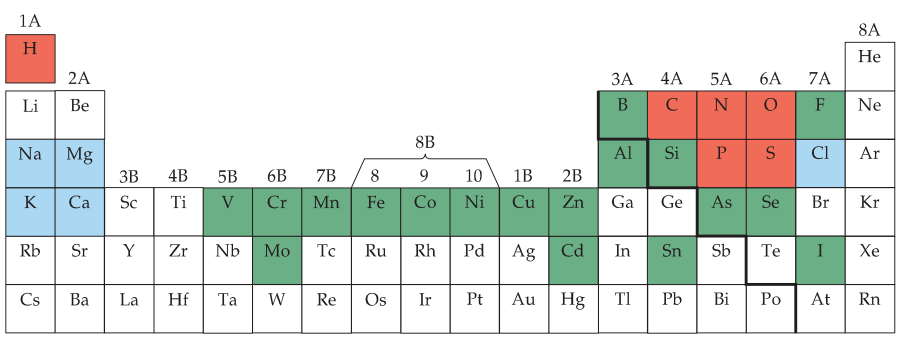
Figure 2.24▼ shows the elements that are essential for life. More than 97% of the mass of most organisms comprises just six elements—oxygen, carbon, hydrogen, nitrogen, phosphorus, and sulfur. Water (H_2O) is the most common compound in living organisms, accounting for at least 70% of the mass of most cells. Carbon is the most prevalent element (by mass) in the solid components of cells. Carbon atoms are found in a vast variety of organic molecules, in which the carbon atoms are bonded to other carbon atoms or to atoms of other elements, principally H, O, N, P, and S. All proteins, for example, contain the following group of atoms that occurs repeatedly within 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 that are required by all organisms: Ca^{2+}, Cl^-, Mg^{2+}, K^+, and Na^+. Calcium ions, for example, are necessary for the formation of bone and for the transmission of signals in the nervous system, such as those that trigger the contraction of cardiac muscles, causing the heart to beat. 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.

▼ **Figure 2.24 Biologically essential elements.** The elements that are essential for life are indicated by colors. Red denotes the six most abundant elements in living systems (hydrogen, carbon, nitrogen, oxygen, phosphorus, and sulfur). Blue indicates the five next most abundant elements. Green indicates the elements needed in only trace amounts.



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 the patterns, rules, and generalizations free us from learning (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 the topics relate to one another, how ideas connect together. They therefore 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: Begin to notice the structure of the topic you are studying. Pay attention to the 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 seen in Table 2.2) and how the periodic table aids us in remembering the charges of ions (as seen in Figure 2.22). You may surprise yourself by observing patterns that are not even explicitly spelled out yet. Perhaps you have even noticed certain trends in chemical formulas. 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 you have already prepared yourself for some topics we will address in later chapters.

We can readily write the empirical formula for an ionic compound if we know the charges of the ions of which the compound is composed. 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 will be 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

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

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

What are the empirical formulas of the compounds 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. Thus, the formula is $AlCl_3$.

(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−). Thus, the formula is Al_2O_3.

(c) Two NO_3^- ions are needed to balance the charge of one Mg^{2+}. Thus, the formula is $Mg(NO_3)_2$. In this case the formula for the entire 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 formulas for the compounds formed by the following ions: **(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

To obtain information about a particular substance, you must know its chemical formula and name. The names and 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 now more than 19 million known chemical substances. Naming them all would be a hopelessly complicated task if each had a special 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 individual, traditional names (so-called "common" names). For most substances, however, we rely on a systematic 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, usually in combination with hydrogen, oxygen, nitrogen, or sulfur. All others are *inorganic compounds*. Early chemists associated organic compounds with plants and animals, and they associated inorganic compounds with the nonliving portion of our world. Although this distinction between living and nonliving matter 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 inorganic compounds, and in Section 2.9 we will introduce the names of some simple organic compounds. Among inorganic compounds, we will consider three categories: 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 positive ions, and the nonmetals form the negative ions. Let's examine the naming of positive ions, then the naming of negative ones. After that, we will consider how to put the names of the ions together to identify the complete ionic compound.

1. Positive Ions (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

Ions formed from a single atom are called *monatomic ions*.

(b) *If a metal can form different cations, 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 exhibit different properties, such as different colors (Figure 2.25▶).

Most of the metals that can form more than one cation are *transition metals*, elements that occur in the middle block of elements, from group 3B to group 2B in the periodic table. The charges of these ions are indicated by Roman numerals. The metals that form only one cation are those of group 1A (Na^+, K^+, and Rb^+) and group 2A (Mg^{2+}, Ca^{2+}, Sr^{2+}, and Ba^{2+}), as well as Al^{3+} (group 3A) and two transition-metal ions: Ag^+ (group 1B) and Zn^{2+} (group 2B). Charges are not expressed explicitly 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 two differently charged ions of a metal is to apply the ending *-ous* or *-ic*.

▲ **Figure 2.25 Ions of the same element with different charges exhibit different properties.** Compounds containing ions of the same element but with different charge can be very different in appearance and properties. Both substances shown are complex compounds of iron that also contain K^+ and CN^- ions. The substance on the left is potassium ferrocyanide, which contains Fe(II) bound to CN^- ions. The substance on the right is potassium ferricyanide, which contains Fe(III) bound to CN^- ions. Both substances are used extensively in blueprinting and other dyeing processes.

These endings represent the lower and higher charged ions, respectively. They are 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. They are both polyatomic. The vast majority of cations are monatomic metal ions.

The names and formulas of some common cations are shown in Table 2.4▼; they are also included in a table of common ions in the back inside cover of the text. The ions listed on the left side in Table 2.4 are the monatomic ions that do not have variable charges. Those listed on the right side are either polyatomic cations or cations with variable charges. The Hg_2^{2+} ion is unusual because this metal ion 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

Why is CrO named using a Roman numeral, chromium(II) oxide, whereas CaO is named without a Roman numeral in the name, calcium oxide?

2. Negative Ions (Anions)

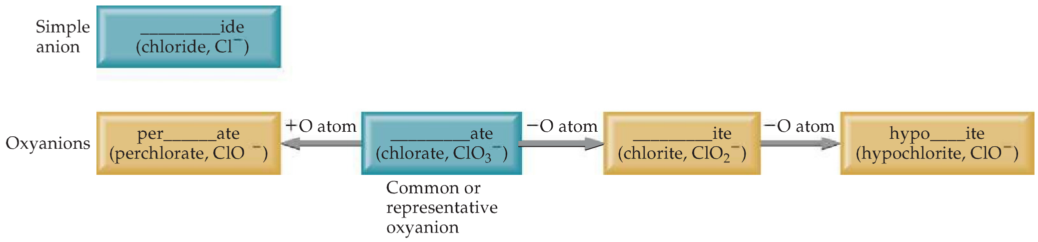
(a) *The names of monatomic anions are formed by replacing the ending of the name of the element with* -ide:

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

TABLE 2.4 ■ Common Cations*

Charge	Formula	Name	Formula	Name
1+	$\mathbf{H^+}$	**Hydrogen ion**	$\mathbf{NH_4^+}$	**Ammonium ion**
	Li^+	Lithium ion	Cu^+	Copper(I) or cuprous ion
	$\mathbf{Na^+}$	**Sodium ion**		
	$\mathbf{K^+}$	**Potassium ion**		
	Cs^+	Cesium ion		
	$\mathbf{Ag^+}$	**Silver ion**		
2+	$\mathbf{Mg^{2+}}$	**Magnesium ion**	Co^{2+}	Cobalt(II) or cobaltous ion
	$\mathbf{Ca^{2+}}$	**Calcium ion**	$\mathbf{Cu^{2+}}$	**Copper(II)** or cupric ion
	Sr^{2+}	Strontium ion	$\mathbf{Fe^{2+}}$	**Iron(II)** or ferrous ion
	Ba^{2+}	Barium ion	Mn^{2+}	Manganese(II) or manganous ion
	$\mathbf{Zn^{2+}}$	**Zinc ion**	Hg_2^{2+}	Mercury(I) or mercurous ion
	Cd^{2+}	Cadmium ion	$\mathbf{Hg^{2+}}$	**Mercury(II)** or mercuric ion
			Ni^{2+}	Nickel(II) or nickelous ion
			$\mathbf{Pb^{2+}}$	**Lead(II)** or plumbous ion
			Sn^{2+}	Tin(II) or stannous ion
3+	$\mathbf{Al^{3+}}$	**Aluminum ion**	Cr^{3+}	Chromium(III) or chromic ion
			$\mathbf{Fe^{3+}}$	**Iron(III)** or ferric ion

*The most common ions are in boldface.



▲ **Figure 2.26** **Summary of the procedure for naming anions.** The root of the name (such as "chlor" for chlorine) goes in the blank.

A few simple 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* -ate *or* -ite. These anions are called **oxyanions**. The ending *-ate* is used for the most common oxyanion of an element. The ending *-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*; the prefix *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 2.26▲.

GIVE IT SOME THOUGHT

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

Students often have a hard time remembering the number of oxygen atoms in the various oxyanions and the charges of these ions. Figure 2.27▼ lists the oxyanions of C, N, P, S, and Cl that contain the maximum number of O atoms. The periodic pattern seen in these formulas can help you remember them. Notice that C and N, which are in the second period of the periodic table, have only three O atoms each, whereas P, S, and Cl, which are in the third period, have four O atoms each. If we begin at the lower right side of the figure, with Cl, we see that the charges increase from right to left, from 1− for Cl (ClO_4^-) to 3− for P (PO_4^{3-}). In the second period the charges also increase from right to left, from 1− for N (NO_3^-) to 2− for C (CO_3^{2-}). Each anion

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

◀ **Figure 2.27** **Common oxyanions.** The composition and charges of common oxyanions are related to their location in the periodic table.

shown in Figure 2.27 has a name ending in *-ate*. The ClO_4^- ion also has a *per-* prefix. If you know the rules summarized in Figure 2.26 and the names and formulas of the five oxyanions in Figure 2.27, you can deduce the names for the other oxyanions of these elements.

GIVE IT SOME THOUGHT

Predict the formulas for the borate ion and silicate ion, assuming that they contain a single B and Si atom, respectively, and follow the trends shown in Figure 2.27▲.

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 members of 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 perbromate ions.
Answer: BrO^- and BrO_4^-

(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^+ reduces the negative charge of the parent anion by one. An older method for naming some of these ions is to use 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

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 most common ions are in boldface.

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 following compounds: **(a)** K_2SO_4, **(b)** $Ba(OH)_2$, **(c)** $FeCl_3$.

SOLUTION

Each compound is ionic and is named using the guidelines we have already discussed. In naming ionic compounds, it is important to recognize polyatomic ions and to determine the charge of cations with variable charge.

(a) The cation in this compound is K^+, and the anion is $SO_4{}^{2-}$. (If you thought the compound contained S^{2-} and O^{2-} ions, you failed to recognize the polyatomic sulfate ion.) Putting together the names of the ions, we have the name of the compound, potassium sulfate.
(b) In this case the compound is composed of Ba^{2+} and OH^- ions. Ba^{2+} is the barium ion and OH^- is the hydroxide ion. Thus, the compound is called 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 Cl^- ions, the cation must be Fe^{3+} which is the iron(III), or ferric, ion. The Cl^- ion is the chloride ion. Thus, the compound is iron(III) chloride or ferric chloride.

PRACTICE EXERCISE

Name the following 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

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

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

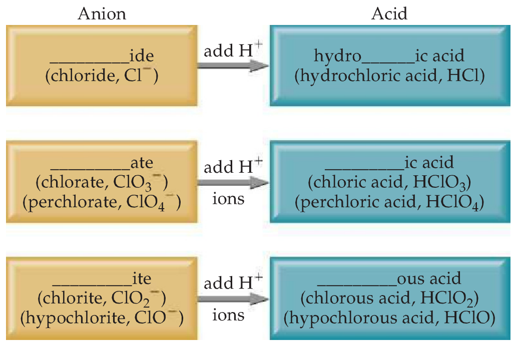
SOLUTION

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 the empirical formula of the compound, K_2S.
(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 formula for **(a)** magnesium sulfate, **(b)** silver sulfide, **(c)** lead(II) nitrate.
Answers: **(a)** $MgSO_4$, **(b)** Ag_2S, **(c)** $Pb(NO_3)_2$



▶ **Figure 2.28 Relating names of anions and acids.** Summary of the way in which anion names and acid names are related. The prefixes *per-* and *hypo-* are retained in going from the anion to the acid.

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 2.28▲.

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, *as in the following examples*:

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. These rules are illustrated by the oxyacids of chlorine:

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

SAMPLE EXERCISE 2.14 | Relating the Names and Formulas of Acids

Name the following 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 is usually written first.* An exception to this rule occurs in the case of compounds that contain oxygen. Oxygen is always written last except when combined with fluorine.
2. *If both elements are in the same group in the periodic table, the one having the higher atomic number 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 (such as *oxide*), 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

It is important to realize that you cannot predict the formulas of most molecular substances in the same way that you predict the formulas of ionic compounds. For this reason, we name molecular compounds using prefixes that explicitly indicate their composition. 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 following compounds: **(a)** SO_2, **(b)** PCl_5, **(c)** N_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)** dinitrogen trioxide.

PRACTICE EXERCISE

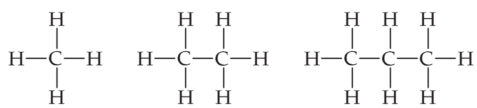
Give the chemical formula for **(a)** silicon tetrabromide, **(b)** disulfur dichloride.
Answers: **(a)** $SiBr_4$, **(b)** S_2Cl_2

2.9 SOME SIMPLE ORGANIC COMPOUNDS

The study of compounds of carbon is called **organic chemistry**, and as noted earlier in the chapter, compounds that contain carbon and hydrogen, often in combination with oxygen, nitrogen, or other elements, are called *organic compounds*. We will examine organic compounds and organic chemistry in some detail in Chapter 25. You will see a number of organic compounds throughout this text; many of them have practical applications or are relevant to the chemistry of biological systems. Here we present a very brief introduction to some of the simplest organic compounds to provide you with a sense of what these molecules look like and how they are named.

Alkanes

Compounds that contain only carbon and hydrogen are called **hydrocarbons**. In the most basic class of hydrocarbons, each carbon atom is bonded to four other atoms. These compounds are called **alkanes**. The three simplest alkanes, which contain one, two, and three carbon atoms, respectively, are methane (CH_4), ethane (C_2H_6), and propane (C_3H_8). The structural formulas of these three alkanes are as follows:



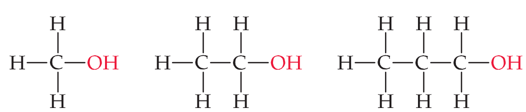
Methane Ethane Propane

We can make longer alkanes by adding additional carbon atoms to the "skeleton" of the molecule.

Although the 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 carbon atoms is called butane. For alkanes with five or more carbon atoms, the names are derived from prefixes such as those in Table 2.6. An alkane with eight carbon atoms, for example, is called *octane* (C_8H_{18}), where the *octa-* prefix for eight is combined with the *-ane* ending for an alkane. Gasoline consists primarily of octanes, as will be discussed in Chapter 25.

Some Derivatives of Alkanes

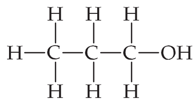
Other classes of organic compounds are obtained when hydrogen atoms of alkanes are replaced with *functional groups*, which are specific groups of atoms. An **alcohol**, for example, is obtained by replacing an H atom of an alkane with an —OH group. The name of the alcohol is derived from that of the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



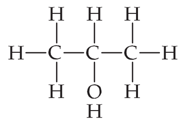
Methanol Ethanol 1-Propanol

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 properties 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 2-propanol (also known as isopropyl alcohol) is obtained if the OH functional group is attached to the middle carbon atom:



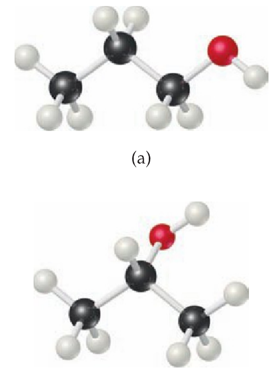
1-Propanol



2-Propanol

Ball-and-stick models of these two molecules are presented in Figure 2.29 ▶.

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.



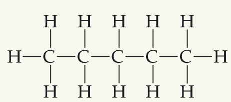
▲ **Figure 2.29 The two forms of propanol (C_3H_7OH).** (a) 1-Propanol, in which the OH group is attached to one of the end carbon atoms, and (b) 2-propanol, in which the OH group is attached to the middle carbon atom.

SAMPLE EXERCISE 2.16 | Writing Structural and Molecular Formulas for Hydrocarbons

Consider the alkane called *pentane*. **(a)** Assuming that the carbon atoms are in a straight line, write a structural formula for pentane. **(b)** What is the molecular formula for pentane?

SOLUTION

(a) Alkanes contain only carbon and hydrogen, and each carbon atom is attached to four other atoms. Because the name pentane contains the prefix *penta-* for five (Table 2.6), we can assume that pentane contains five carbon atoms bonded in a chain. If we then add enough hydrogen atoms to make four bonds to each carbon atom, we obtain the following structural formula:



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 can determine the molecular formula by counting the atoms present. Thus, *n*-pentane has the formula C_5H_{12}.

PRACTICE EXERCISE

Butane is the alkane with four carbon atoms. **(a)** What is the molecular formula of butane? **(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 REVIEW

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 α particles showed 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, 1.602×10^{-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 periodic **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*.

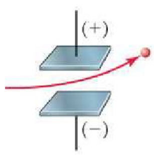
Section 2.9 **Organic chemistry** is the study of compounds that contain carbon. The simplest class of organic molecules is the **hydrocarbons**, which contain only carbon and hydrogen. Hydrocarbons in which each carbon atom is attached to four other atoms are called **alkanes**. Alkanes have names that 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.

KEY SKILLS

- Describe the basic postulates of Dalton's atomic theory.
- Describe the key experiments that led to the discovery of electrons and to the nuclear model of the atom.
- Describe the structure of the atom in terms of protons, neutrons, and electrons.
- Describe the electric charge and relative masses of protons, neutrons, and electrons.
- Use chemical symbols together with atomic number and mass number to express the subatomic composition of isotopes.
- Understand atomic weights and how they relate to the masses of individual atoms and to their natural abundances.
- Describe how elements are organized in the periodic table by atomic number and by similarities in chemical behavior, giving rise to periods and groups.
- Describe the locations of metals and nonmetals in the periodic table.
- Distinguish between molecular substances and ionic substances in terms of their composition.
- Distinguish between empirical formulas and molecular formulas.
- Describe how molecular formulas and structural formulas are used to represent the compositions of molecules.
- 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 ions.
- Write the empirical formulas of ionic compounds, given their charges.
- Write the name of an ionic compound given its chemical formula, or write the chemical formula given its name.
- Name or write chemical formulas for binary inorganic compounds and for acids.
- Identify organic compounds and name simple alkanes and alcohols.

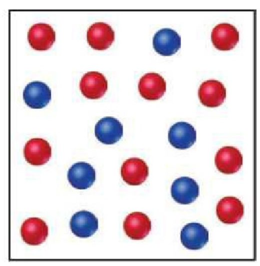
VISUALIZING CONCEPTS

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

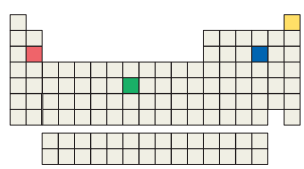


(a) Why does the path of the charged particle bend? **(b)** What is the sign of the electrical charge on the particle? **(c)** As the charge on the plates is increased, would you expect the bending to increase, decrease, or stay the same? **(d)** As the mass of the particle is increased while the speed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

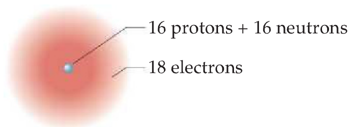
2.2 The following diagram is a representation of 20 atoms of a fictitious element, which we will call nevadium (Nv). The red spheres are ^{293}Nv, and the blue spheres are ^{295}Nv. **(a)** Assuming that this sample is a statistically representative sample of the element, calculate the percent abundance 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]



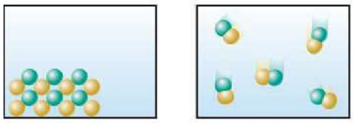
2.3 Four of the boxes in the following periodic table are colored. Which of these are metals and which are nonmetals? Which one is an alkaline earth metal? Which one is a noble gas? [Section 2.5]



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]

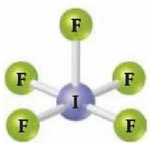


2.5 Which of the following diagrams most likely represents an ionic compound, and which represents a molecular one? Explain your choice. [Sections 2.6 and 2.7]

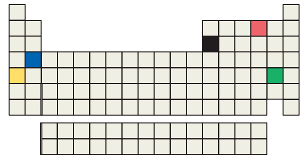


(i) (ii)

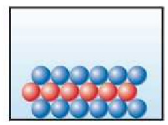
2.6 Write the chemical formula for the following compound. Is the compound ionic or molecular? Name the compound. [Sections 2.6 and 2.8]



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]



EXERCISES

Atomic Theory and the Discovery of Atomic Structure

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 **(a)** Figure 2.5 shows the apparatus used in the Millikan oil-drop experiment with the positively charged plate above the negatively charged plate. What do you think would be the effect on the rate of oil drops descending if the charges on the plates were reversed (negative above positive)? **(b)** In his original series of experiments, Millikan measured the charge on 58 separate oil drops. Why do you suppose he chose so many drops before reaching his final conclusions?

2.16 Millikan determined the charge on the electron by studying the static charges on oil drops falling in an electric field. 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	1.60×10^{-19}
B	3.15×10^{-19}
C	4.81×10^{-19}
D	6.31×10^{-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

2.17 The radius of an atom of krypton (Kr) is about 1.9 Å. **(a)** Express this distance in nanometers (nm) and in picometers (pm). **(b)** How many krypton 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 Kr atom?

2.18 An atom of tin (Sn) has a diameter of about 2.8×10^{-8} cm. **(a)** What is the radius of a tin atom in angstroms (Å) and in meters (m)? **(b)** How many Sn atoms would have to be placed side by side to span a distance of 6.0 μm? **(c)** If you assume that the tin 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		44			92
Neutrons		57	49		
Electrons			38	47	
Mass no.				108	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 2.13 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 2.14) 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

2.37 For each of the following elements, write its chemical symbol, locate it in the periodic table, 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; indicate whether it is a metal, metalloid, or nonmetal; and give the name of the element: **(a)** Ca, **(b)** Ti, **(c)** Ga, **(d)** Th, **(e)** Pt, **(f)** Se, **(g)** Kr.

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, the other is a viscous liquid. How is it possible for two substances with the same empirical formula to have markedly different properties?

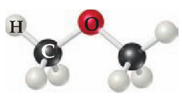
2.43 Write the empirical formula corresponding to each of the following molecular formulas: **(a)** Al_2Br_6, **(b)** C_8H_{10}, **(c)** $C_4H_8O_2$, **(d)** P_4O_{10}, **(e)** $C_6H_4Cl_2$, **(f)** $B_3N_3H_6$.

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.

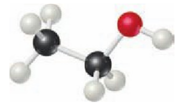
2.45 How many hydrogen atoms are in each of the following: **(a)** C_2H_5OH, **(b)** $Ca(CH_3COO)_2$, **(c)** $(NH_4)_3PO_4$?

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 $(NH_4)_2HPO_4$?

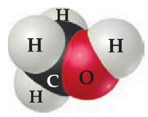
2.47 Write the molecular and structural formulas for the compounds represented by the following molecular models:



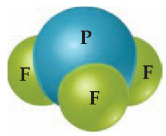
(a)



(b)

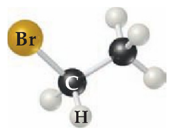


(c)

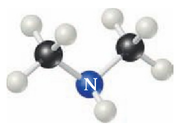


(d)

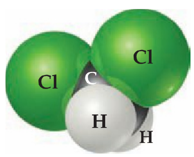
2.48 Write the molecular and structural formulas for the compounds represented by the following models:



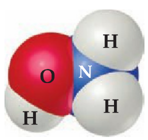
(a)



(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		35	49	
Neutrons		45	66	118
Electrons			46	76
Net charge		1−		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 silver in its compounds is 1+. Indicate the chemical formulas you would expect for compounds formed between Ag and **(a)** iodine, **(b)** sulfur, **(c)** fluorine.

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)** Cu^{2+} and Br^-, **(b)** Fe^{3+} and O^{2-}, **(c)** Hg_2^{2+} and CO_3^{2-}, **(d)** Ca^{2+} and AsO_4^{3-}, **(e)** NH_4^+ and CO_3^{2-}.

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

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)** MgO, **(b)** $AlCl_3$, **(c)** Li_3PO_4, **(d)** $Ba(ClO_4)_2$, **(e)** $Cu(NO_3)_2$, **(f)** $Fe(OH)_2$, **(g)** $Ca(C_2H_3O_2)_2$, **(h)** $Cr_2(CO_3)_3$, **(i)** K_2CrO_4, **(j)** $(NH_4)_2SO_4$.

2.66 Name the following ionic compounds: **(a)** K_2O, **(b)** $NaClO_2$, **(c)** $Sr(CN)_2$, **(d)** $Co(OH)_2$, **(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)** hydrobromic acid, **(b)** hydrosulfuric acid, **(c)** nitrous acid, **(d)** H_2CO_3, **(e)** $HClO_3$, **(f)** $HC_2H_3O_2$.

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 end carbon atoms.

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

ADDITIONAL EXERCISES

2.79 Describe a major contribution to science made by each of the following scientists: **(a)** Dalton, **(b)** Thomson, **(c)** Millikan, **(d)** Rutherford.

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

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). He obtains the following data for four of the drops:

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

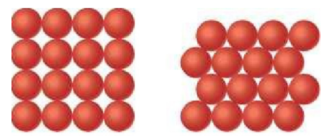
(a) If all the droplets were the same size, which would fall most slowly through the apparatus? **(b)** From these data, what is the best choice for the charge of the electron in warmombs? **(c)** Based on your answer to part (b), how many electrons are there on each of the droplets? **(d)** What is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

2.82 The natural abundance of ^{3}He is 0.000137%. **(a)** How many protons, neutrons, and electrons are in an atom of ^{3}He? **(b)** Based on the sum of the masses of their subatomic particles, which is expected to be more massive, an atom of ^{3}He or an atom of ^{3}H (which is also called *tritium*)? **(c)** Based on your answer for part (b), what would need to be the precision of a mass spectrometer that is able to differentiate between peaks that are due to $^3He^+$ and $^3H^+$?

2.83 An α particle is the nucleus of an ^{4}He atom. **(a)** How many protons and neutrons are in an α particle? **(b)** What force holds the protons and neutrons together in the α particle? **(c)** What is the charge on an α particle in units of electronic charge? **(d)** The charge-to-mass ratio of an α particle is 4.8224×10^4 C/g. Based on the charge on the particle, calculate its mass in grams and in amu. **(e)** By using the data in Table 2.1, compare your answer for part (d) with the sum of the masses of the individual subatomic particles. Can you explain the difference in mass? (If not, we will discuss such mass differences further in Chapter 21.)

2.84 A cube of gold that is 1.00 cm on a side has a mass of 19.3 g. A single gold atom has a mass of 197.0 amu. **(a)** How many gold atoms are in the cube? **(b)** From the information given, estimate the diameter in Å of a single gold atom. **(c)** What assumptions did you make in arriving at your answer for part (b)?

2.85 The diameter of a rubidium atom is 4.95 Å. We will consider two different ways of placing the atoms on a surface. In arrangement A, all the atoms are lined up with one another. Arrangement B is called a *close-packed* arrangement because the atoms sit in the "depressions" formed by the previous row of atoms:



(a) (b)

(a) Using arrangement A, how many Rb atoms could be placed on a square surface that is 1.0 cm on a side? **(b)** How many Rb atoms could be placed on a square surface that is 1.0 cm on a side, using arrangement B? **(c)** By what factor has the number of atoms on the surface increased in going to arrangement B from arrangement A? If extended to three dimensions, which arrangement would lead to a greater density for Rb metal?

2.86 **(a)** Assuming the dimensions of the nucleus and atom shown in Figure 2.12, what fraction of the *volume* of the atom is taken up by the nucleus? **(b)** Using the mass of the proton from Table 2.1 and assuming its diameter is 1.0×10^{-15} m, calculate the density of a proton in g/cm^3.

2.87 Identify the element represented by each of the following symbols and give the number of protons and neutrons in each: **(a)** $^{74}_{33}X$, **(b)** $^{127}_{53}X$, **(c)** $^{152}_{63}X$, **(d)** $^{209}_{83}X$.

2.88 The element oxygen has three naturally occurring isotopes, with 8, 9, and 10 neutrons in the nucleus, respectively. **(a)** Write the full chemical symbols for these three isotopes. **(b)** Describe the similarities and differences between the three kinds of atoms of oxygen.

2.89 Use Coulomb's law, $F = kQ_1Q_2/d^2$, to calculate the electric force on an electron ($Q = -1.6 \times 10^{-19}$ C) exerted by a single proton if the particles are 0.53×10^{-10} m apart. The constant k in Coulomb's law is $9.0 \times 10^9\ N \cdot m^2/C^2$. (The unit abbreviated N is the Newton, the SI unit of force.)

2.90 The element lead (Pb) consists of four naturally occurring isotopes with atomic masses 203.97302, 205.97444, 206.97587, and 207.97663 amu. The relative abundances of these four isotopes are 1.4, 24.1, 22.1, and 52.4%, respectively. From these data, calculate the atomic weight of lead.

2.91 Gallium (Ga) consists of two naturally occurring isotopes with masses of 68.926 and 70.925 amu. **(a)** How many protons and neutrons are in the nucleus of each isotope? Write the complete atomic symbol for each, showing the atomic number and mass number. **(b)** The average atomic mass of Ga is 69.72 amu. Calculate the abundance of each isotope.

2.92 Using a suitable reference such as the *CRC Handbook of Chemistry and Physics* or http://www.webelements.com, look up the following information for nickel: **(a)** the number of known isotopes, **(b)** the atomic masses (in amu) and the natural abundance of the five most abundant isotopes.

2.93 There are two different isotopes of bromine atoms. Under normal conditions, elemental bromine consists of Br_2 molecules (Figure 2.19), and the mass of a Br_2 molecule is the sum of the masses of the two atoms in the molecule. The mass spectrum of Br_2 consists of three peaks:

Mass (amu)	Relative Size
157.836	0.2569
159.834	0.4999
161.832	0.2431

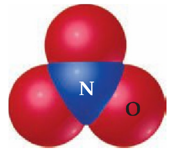
(a) What is the origin of each peak (of what isotopes does each consist)? **(b)** What is the mass of each isotope? **(c)** Determine the average molecular mass of a Br_2 molecule. **(d)** Determine the average atomic mass of a bromine atom. **(e)** Calculate the abundances of the two isotopes.

2.94 It is common in mass spectrometry to assume that the mass of a cation is the same as that of its parent atom. **(a)** Using data in Table 2.1, determine the number of significant figures that must be reported before the difference in mass of ^{1}H and $^{1}H^{+}$ is significant. **(b)** What percentage of the mass of an ^{1}H atom does the electron represent?

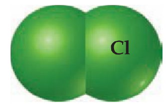
2.95 From the following list of elements—Ar, H, Ga, Al, Ca, Br, Ge, K, O—pick the one that best fits each description. Use each element only once: **(a)** an alkali metal, **(b)** an alkaline earth metal, **(c)** a noble gas, **(d)** a halogen, **(e)** a metalloid, **(f)** a nonmetal listed in group 1A, **(g)** a metal that forms a 3+ ion, **(h)** a nonmetal that forms a 2− ion, **(i)** an element that resembles aluminum.

2.96 The first atoms of seaborgium (Sg) were identified in 1974. The longest-lived isotope of Sg has a mass number of 266. **(a)** How many protons, electrons, and neutrons are in an ^{266}Sg atom? **(b)** Atoms of Sg are very unstable, and it is therefore difficult to study this element's properties. Based on the position of Sg in the periodic table, what element should it most closely resemble in its chemical properties?

2.97 From the molecular structures shown here, identify the one that corresponds to each of the following species: **(a)** chlorine gas; **(b)** propane, **(c)** nitrate ion; **(d)** sulfur trioxide; **(e)** methyl chloride, CH_3Cl.



(i)



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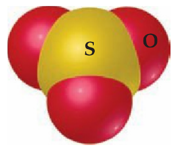

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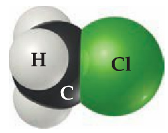

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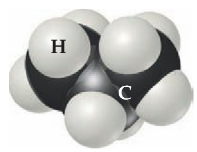


(v)

2.98 Name each of the following oxides. Assuming that the compounds are ionic, what charge is associated with the metallic element in each case? **(a)** NiO, **(b)** MnO_2, **(c)** Cr_2O_3, **(d)** MoO_3.

2.99 Iodic acid has the molecular formula HIO_3. Write the formulas for the following: **(a)** the iodate anion, **(b)** the periodate anion, **(c)** the hypoiodite anion, **(d)** hypoiodous acid, **(e)** periodic acid.

2.100 Elements in the same group of the periodic table often form oxyanions with the same general formula. The anions are also named in a similar fashion. Based on these observations, suggest a chemical formula or name, as appropriate, for each of the following ions: **(a)** BrO_4^{-}, **(b)** SeO_3^{2-}, **(c)** arsenate ion, **(d)** hydrogen tellurate ion.

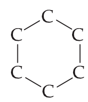
2.101 Carbonic acid occurs in carbonated beverages. When allowed to react with lithium hydroxide it produces lithium carbonate. Lithium carbonate is used to treat depression and bipolar disorder. Write chemical formulas for carbonic acid, lithium hydroxide, and lithium carbonate.

2.102 Give the chemical names of each of the following familiar compounds: **(a)** NaCl (table salt) , **(b)** $NaHCO_3$ (baking soda), **(c)** NaOCl (in many bleaches), **(d)** NaOH (caustic soda), **(e)** $(NH_4)_2CO_3$ (smelling salts), **(f)** $CaSO_4$ (plaster of Paris).

2.103 Many familiar substances have common, unsystematic names. For each of the following, give the correct systematic name: **(a)** saltpeter, KNO_3; **(b)** soda ash, Na_2CO_3; **(c)** lime, CaO; **(d)** muriatic acid, HCl; **(e)** Epsom salts, $MgSO_4$; **(f)** milk of magnesia, $Mg(OH)_2$.

2.104 Many ions and compounds have very similar names, and there is great potential for confusing them. Write the correct chemical formulas to distinguish between **(a)** calcium sulfide and calcium hydrogen sulfide, **(b)** hydrobromic acid and bromic acid, **(c)** aluminum nitride and aluminum nitrite, **(d)** iron(II) oxide and iron(III) oxide, **(e)** ammonia and ammonium ion, **(f)** potassium sulfite and potassium bisulfite, **(g)** mercurous chloride and mercuric chloride, **(h)** chloric acid and perchloric acid.

2.105 The compound *cyclohexane* is an alkane in which six carbon atoms form a ring. The partial structural formula of the compound is as follows:



(a) Complete the structural formula for cyclohexane. **(b)** Is the molecular formula for cyclohexane the same as that for *n*-hexane, in which the carbon atoms are in a straight line? If possible, comment on the source of any differences. **(c)** Propose a structural formula for *cyclohexanol*, the alcohol derived from cyclohexane.

2.106 The periodic table helps organize the chemical behaviors of the elements. As a class discussion or as a short essay, describe how the table is organized, and mention as many ways as you can think of in which the position of an element in the table relates to the chemical and physical properties of the element.