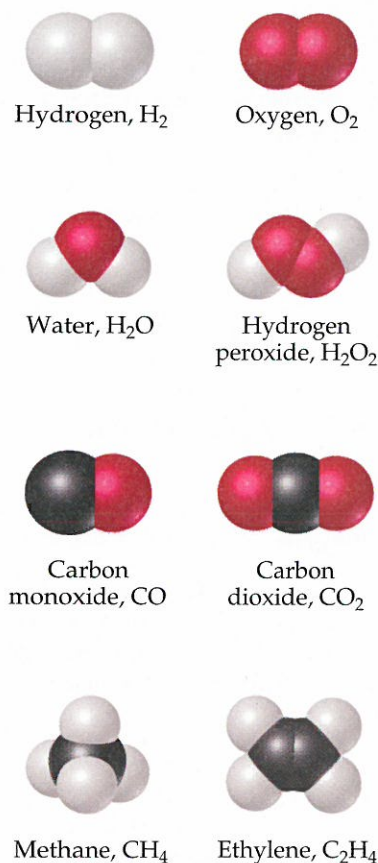


oxygen by the **chemical formula** O_2 (read “oh two”). The subscript tells us that two oxygen atoms are present in each molecule. A molecule made up of two atoms is called a **diatomic molecule**.

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

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

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



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

Molecular and Empirical Formulas

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

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

SAMPLE EXERCISE 2.6 Relating Empirical and Molecular Formulas

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

SOLUTION

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

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

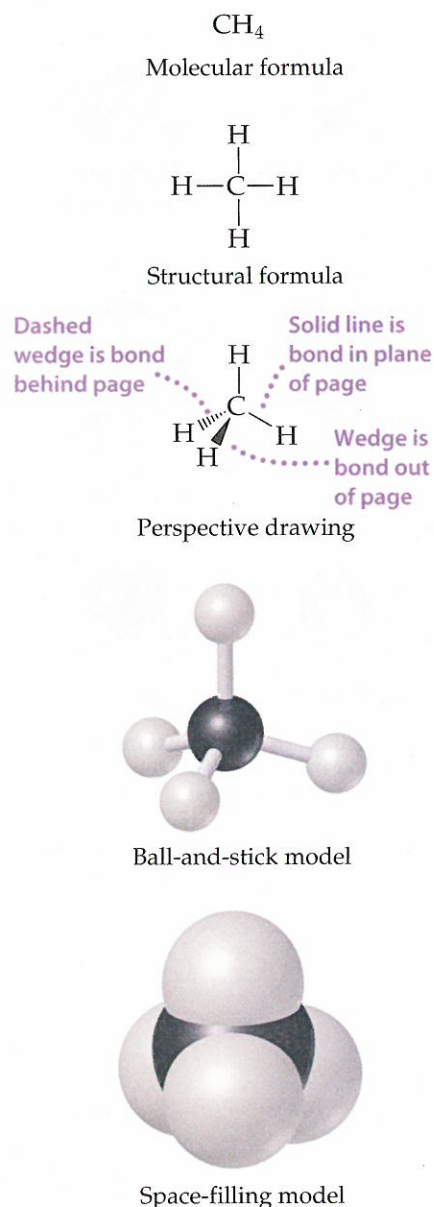
PRACTICE EXERCISE

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

Answer: BH_3

GO FIGURE

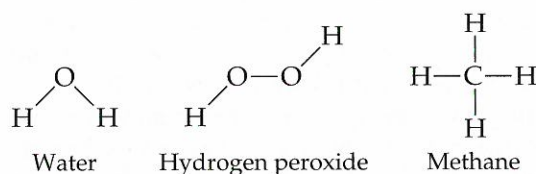
What advantage does a ball-and-stick model have over a space-filling model?



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

Picturing Molecules

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



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

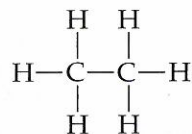
A structural formula usually does not depict the actual geometry of the molecule, that is, the actual angles at which atoms are joined together. A structural formula can be written as a *perspective drawing* (◀ FIGURE 2.19), however, to give some sense of three-dimensional shape.

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

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

GIVE IT SOME THOUGHT

The structural formula for ethane is

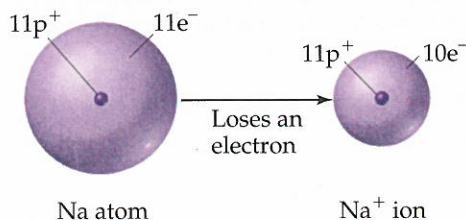


- What is the molecular formula for ethane?
- What is its empirical formula?
- Which kind of molecular model would most clearly show the angles between atoms?

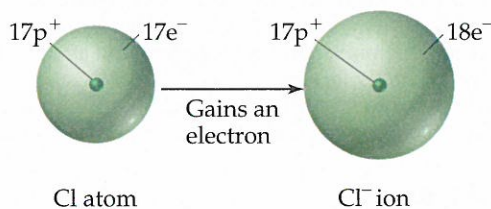
2.7 IONS AND IONIC COMPOUNDS

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

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



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



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

SAMPLE EXERCISE 2.7 Writing Chemical Symbols for Ions

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

SOLUTION

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

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

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 ${}^{79}\text{Se}^{2-}$ ion possess?

Answer: 34 protons, 45 neutrons, and 36 electrons

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

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

Predicting Ionic Charges

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

SAMPLE EXERCISE 2.8 Predicting Ionic Charge

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

SOLUTION

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

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

PRACTICE EXERCISE

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

Answer: (a) $3+$, (b) $1-$

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

Ionic Compounds

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

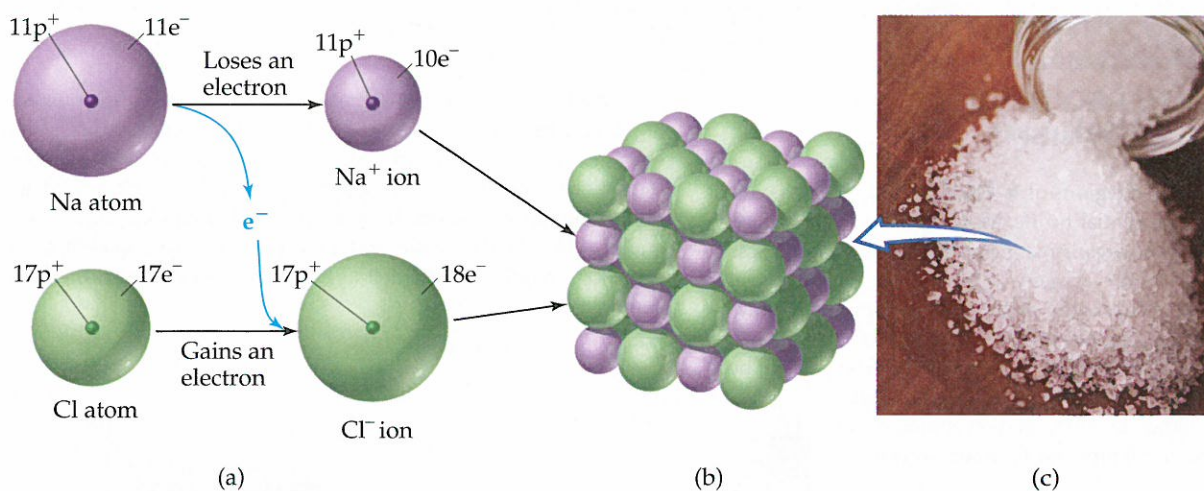
We can often tell whether a compound is ionic (consisting of ions) or molecular (consisting of molecules) from its composition. In general, cations are metal ions and anions are nonmetal ions. Consequently, *ionic compounds are generally combinations of metals and nonmetals*, as in NaCl . In contrast,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

GO FIGURE

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

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

1A	2A	Transition metals										3A	4A	5A	6A	7A	8A			
H^+																	H^-	Noble Gas		
Li^+																	N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}																Al^{3+}		S^{2-}	Cl^-
K^+	Ca^{2+}																		Se^{2-}	Br^-
Rb^+	Sr^{2+}																		Te^{2-}	I^-
Cs^+	Ba^{2+}																			



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

SAMPLE EXERCISE 2.9 Identifying Ionic and Molecular Compounds

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

SOLUTION

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

PRACTICE EXERCISE

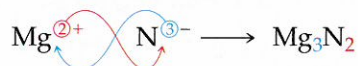
Which of these compounds are molecular: CBr_4 , FeS , P_4O_6 , PbF_2 ?

Answer: CBr_4 and P_4O_6

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

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

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



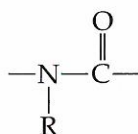
▲ GIVE IT SOME THOUGHT

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

CHEMISTRY AND LIFE

ELEMENTS REQUIRED BY LIVING ORGANISMS

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



which occurs repeatedly in the molecules. (R is either an H atom or a combination of atoms, such as CH₃.)

1A																	8A			
H																	He			
2A	Li	Be											3A	4A	5A	6A	7A	8A		
	Na	Mg	3B	4B	5B	6B	7B	8B			9	10	11B	12B	B	C	N	O	F	Ne
	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr		
	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe		

■ 6 most abundant essential elements

■ 5 next most abundant essential elements

■ elements needed only in trace quantities

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

In addition, 23 more elements have been found in various living organisms. Five are ions required by all organisms: Ca²⁺, Cl⁻, Mg²⁺, K⁺, and Na⁺. Calcium ions, for example, are necessary for the formation of bone and transmission of nervous system signals. Many other elements are needed in only very small quantities and consequently are called *trace* elements. For example, trace quantities of copper are required in the diet of humans to aid in the synthesis of hemoglobin.

RELATED EXERCISE: 2.96

STRATEGIES IN CHEMISTRY

PATTERN RECOGNITION

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

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

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

You may surprise yourself by observing patterns that are not explicitly spelled out yet. Perhaps you have noticed certain trends in chemical formulas, for instance. Moving across the periodic table from element 11 (Na), we find that the elements form compounds with F having the following compositions: NaF, MgF₂, and AlF₃. Does this trend continue? Do SiF₄, PF₅, and SF₆ exist? Indeed they do. If you have noticed trends like this from the scraps of information you have seen so far, then you are ahead of the game and have prepared yourself for some topics we will address in later chapters.

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

Write the empirical formula of the compound formed by (a) Al³⁺ and Cl⁻ ions, (b) Al³⁺ and O²⁻ ions, (c) Mg²⁺ and NO₃⁻ ions.

SOLUTION

(a) Three Cl⁻ ions are required to balance the charge of one Al³⁺ ion, making the formula AlCl₃.

(b) Two Al³⁺ ions are required to balance the charge of three O²⁻ ions. That is, the total positive charge is 6+, and the total negative charge is 6-. The formula is Al₂O₃.

(c) Two NO_3^- ions are needed to balance the charge of one Mg^{2+} , yielding $\text{Mg}(\text{NO}_3)_2$. Note that the formula for the polyatomic ion, NO_3^- , must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion.

PRACTICE EXERCISE

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

Answers: (a) Na_3PO_4 , (b) ZnSO_4 , (c) $\text{Fe}_2(\text{CO}_3)_3$

2.8 NAMING INORGANIC COMPOUNDS

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

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

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

Names and Formulas of Ionic Compounds

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

1. Cations

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

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

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

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

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

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



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

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

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

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

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

NH_4^{+}	ammonium ion	H_3O^{+}	hydronium ion
-------------------	--------------	--------------------------	---------------

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

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

▲ GIVE IT SOME THOUGHT

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

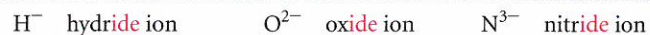
TABLE 2.4 • Common Cations*

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

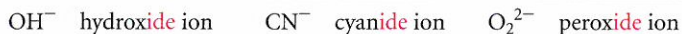
*The ions we use most often in this course are in boldface. Learn them first.

2. Anions

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



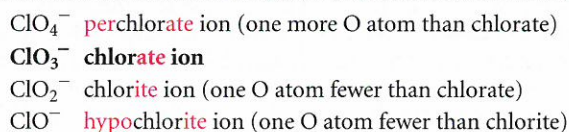
A few polyatomic anions also have names ending in *-ide*:



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



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



These rules are summarized in ▼ FIGURE 2.24.

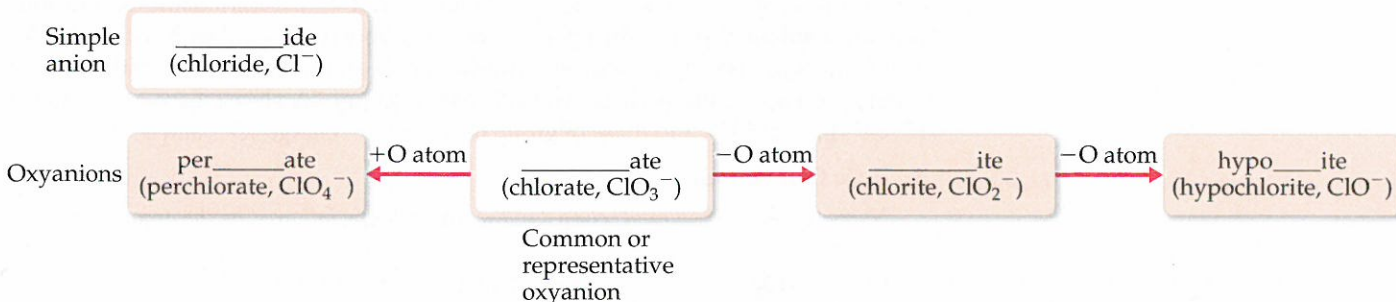
GIVE IT SOME THOUGHT

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

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

GO FIGURE

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



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

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

Maximum of 3 O atoms in period 2.

Maximum of 4 O atoms in period 3.

Charges increase right to left.

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

▲ GIVE IT SOME THOUGHT

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

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

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

SOLUTION

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

PRACTICE EXERCISE

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

Answer: BrO^- and BrO_2^-

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

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

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

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

3. Ionic Compounds

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

CaCl_2	calcium chloride
$\text{Al}(\text{NO}_3)_3$	aluminum nitrate
$\text{Cu}(\text{ClO}_4)_2$	copper(II) perchlorate (or cupric perchlorate)

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

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

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

SOLUTION

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

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

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

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

PRACTICE EXERCISE

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

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

TABLE 2.5 • Common Anions*

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

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

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

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

SOLUTION

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

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

(b) The calcium ion is Ca^{2+} . The carbonate ion is CO_3^{2-} , so the hydrogen carbonate ion is HCO_3^- . Two HCO_3^- ions are needed to balance the positive charge of Ca^{2+} , giving $\text{Ca}(\text{HCO}_3)_2$.

(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 $\text{Ni}(\text{ClO}_4)_2$.

PRACTICE EXERCISE

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

Answers: (a) MgSO_4 , (b) Ag_2S , (c) $\text{Pb}(\text{NO}_3)_2$

Names and Formulas of Acids

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

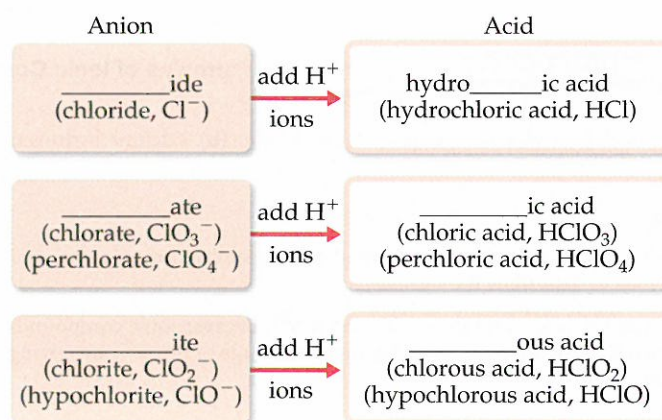
An acid is composed of an anion connected to enough H^+ ions to neutralize, or balance, the anion's charge. Thus, the SO_4^{2-} ion requires two H^+ ions, forming H_2SO_4 . The name of an acid is related to the name of its anion, as summarized in ▼ FIGURE 2.26.

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

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

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

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



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

GIVE IT SOME THOUGHT

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

SAMPLE EXERCISE 2.14 Relating the Names and Formulas of Acids

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

SOLUTION

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

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

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

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

PRACTICE EXERCISE

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

Answers: (a) HBr , (b) H_2CO_3

Names and Formulas of Binary Molecular Compounds

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

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

The following examples illustrate these rules:

Cl_2O	dichlorine monoxide	NF_3	nitrogen trifluoride
N_2O_4	dinitrogen tetroxide	P_4S_{10}	tetraphosphorus decaulfide

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

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

Prefix	Meaning
<i>Mono-</i>	1
<i>Di-</i>	2
<i>Tri-</i>	3
<i>Tetra-</i>	4
<i>Penta-</i>	5
<i>Hexa-</i>	6
<i>Hepta-</i>	7
<i>Octa-</i>	8
<i>Nona-</i>	9
<i>Deca-</i>	10

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

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

SOLUTION

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

PRACTICE EXERCISE

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

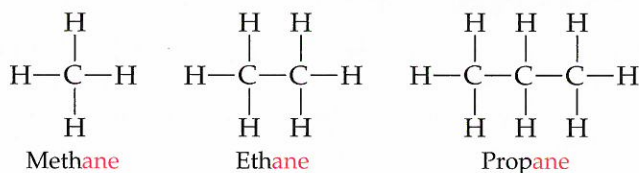
Answers: (a) SiBr_4 , (b) S_2Cl_2

2.9 SOME SIMPLE ORGANIC COMPOUNDS

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

Alkanes

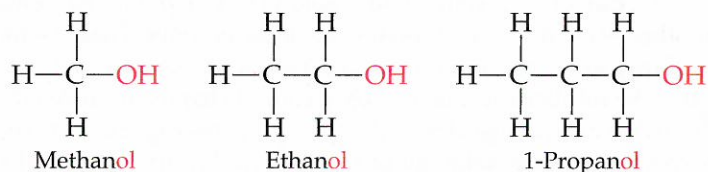
Compounds that contain only carbon and hydrogen are called **hydrocarbons**. In the simplest class of hydrocarbons, **alkanes**, each carbon is bonded to four other atoms. The three smallest alkanes are methane (CH_4), ethane (C_2H_6), and propane (C_3H_8). The structural formulas of these three alkanes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



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

Some Derivatives of Alkanes

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



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

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

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

▲ GIVE IT SOME THOUGHT

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

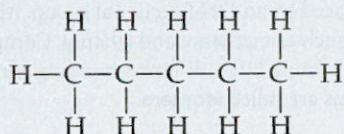
Much of the richness of organic chemistry is possible because organic compounds can form long chains of carbon–carbon bonds. The series of alkanes that begins with methane, ethane, and propane and the series of alcohols that begins with methanol, ethanol, and propanol can both be extended for as long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

SAMPLE EXERCISE 2.16 Writing Structural and Molecular Formulas for Hydrocarbons

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

SOLUTION

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



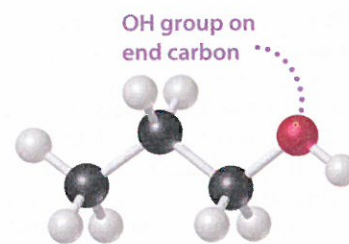
This form of pentane is often called *n*-pentane, where the *n*- stands for “normal” because all five carbon atoms are in one line in the structural formula.

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

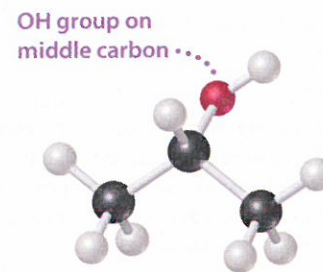
PRACTICE EXERCISE

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

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



1-Propanol



2-Propanol

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

CHAPTER SUMMARY AND KEY TERMS

SECTIONS 2.1 AND 2.2 Atoms are the basic building blocks of matter. They are the smallest units of an element that can combine with other elements. Atoms are composed of even smaller particles, called **subatomic particles**. Some of these subatomic particles are charged and follow the usual behavior of charged particles: Particles with the same charge repel one another, whereas particles with unlike charges are attracted to one another. We considered some of the important experiments that led to the discovery and characterization of subatomic particles. Thomson’s experiments on the behavior of

cathode rays in magnetic and electric fields led to the discovery of the electron and allowed its charge-to-mass ratio to be measured. Millikan’s oil-drop experiment determined the charge of the electron. Becquerel’s discovery of **radioactivity**, the spontaneous emission of radiation by atoms, gave further evidence that the atom has a substructure. Rutherford’s studies of how thin metal foils scatter α particles led to the **nuclear model** of the atom, showing that the atom has a dense, positively charged **nucleus**.

SECTION 2.3 Atoms have a nucleus that contains **protons** and **neutrons**; **electrons** move in the space around the nucleus. The magnitude of the charge of the electron, $1.602 \times 10^{-19} \text{ C}$, 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** ($1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g}$). The dimensions of atoms are often expressed in units of **angstroms** ($1 \text{ \AA} = 10^{-10} \text{ m}$).

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

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

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

SECTION 2.6 Atoms can combine to form **molecules**. Compounds composed of molecules (**molecular compounds**) usually contain only nonmetallic elements. A molecule that contains two atoms is called a **diatomic molecule**. The composition of a substance is given by its **chemical formula**. A molecular substance can be represented by its **empirical formula**, which gives the relative numbers of atoms of each kind. It is usually represented by its **molecular formula**, however, which gives the actual numbers of each type of atom in a molecule.

Structural formulas show the order in which the atoms in a molecule are connected. Ball-and-stick models and space-filling models are often used to represent molecules.

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

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

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

KEY SKILLS

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