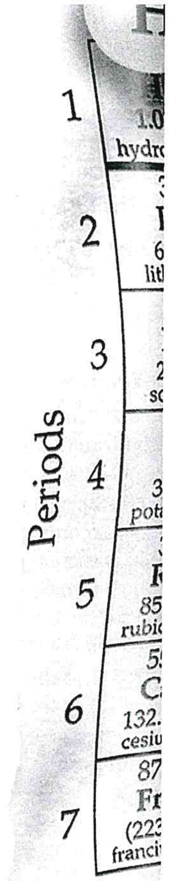
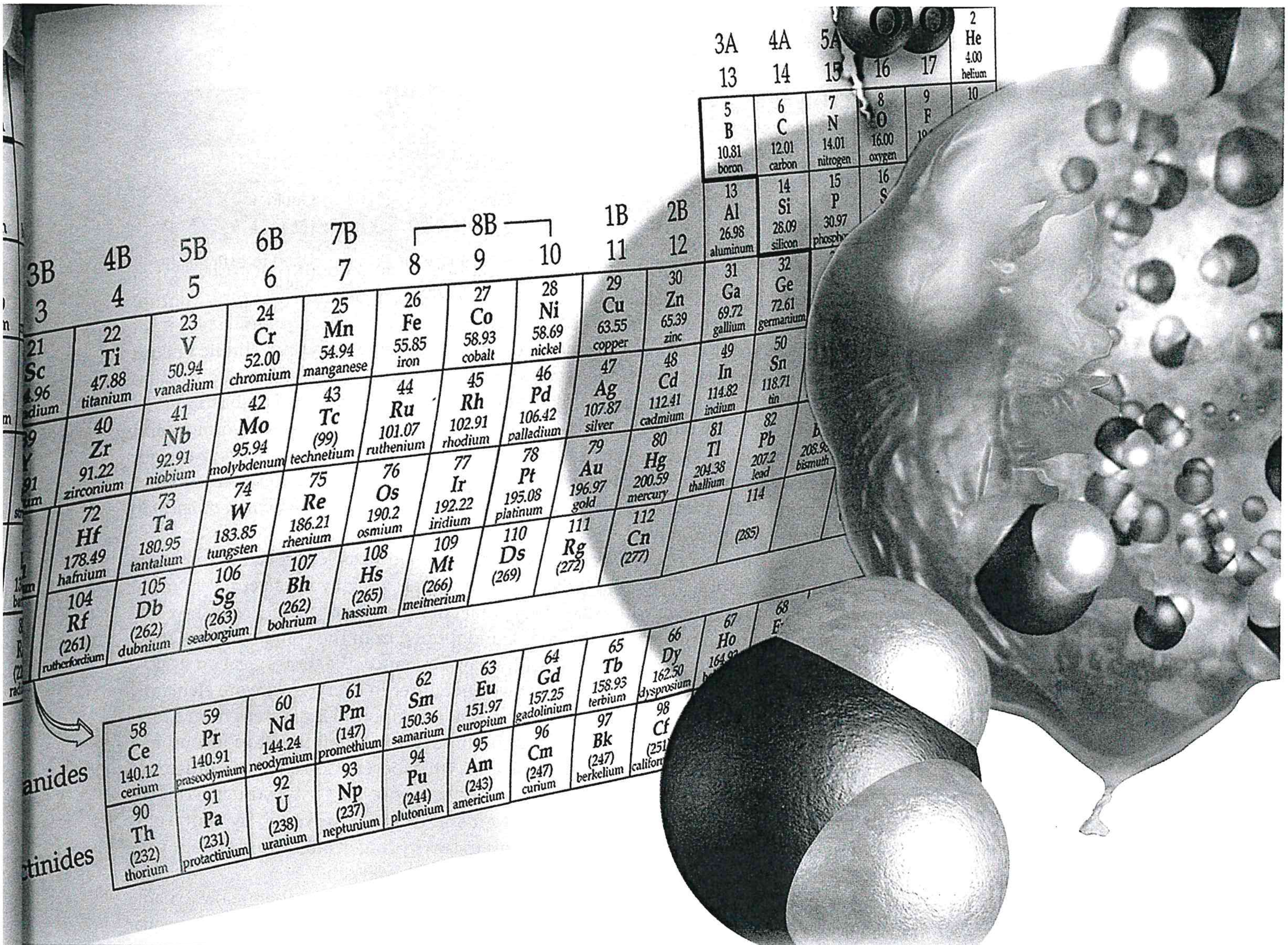
CHAPTER 3

Molecules, Compounds, and Chemical Equations



> **Almost all aspects of life are engineered at the molecular level, and without understanding molecules we can only have a very sketchy understanding of life itself.**
>
> —FRANCIS HARRY COMPTON CRICK (1916–2004)



▲ *When a balloon filled with H_2 and O_2 is ignited, the two elements react violently to form H_2O.*

How many different substances exist? Recall from Chapter 2 that about 91 different elements exist in nature, so there are at least 91 different substances. However, the world would be dull—not to mention lifeless—with only 91 different substances. Fortunately, elements combine with each other to form *compounds*. Just as combinations of only 26 letters in our English alphabet allow for an almost limitless number of words, each with its own specific meaning, combinations of the 91 naturally occurring elements allow for an almost limitless number of compounds, each with its own specific properties. The great diversity of substances that we find in nature is a direct result of the ability of elements to form compounds. Life, for example, could not exist with just 91 different elements. It takes compounds, in all of their diversity, to make life possible.

3.1 Hydrogen, Oxygen, and Water

Hydrogen (H_2) is an explosive gas used as a fuel in rocket engines. Oxygen (O_2), also a gas, is a natural component of the air on Earth. Oxygen itself is not flammable but must be present for combustion (burning) to occur. Hydrogen and oxygen both have extremely low boiling points, as you can see from the table on the following page. When hydrogen and oxygen combine to form the compound water (H_2O), however, a dramatically different substance results.

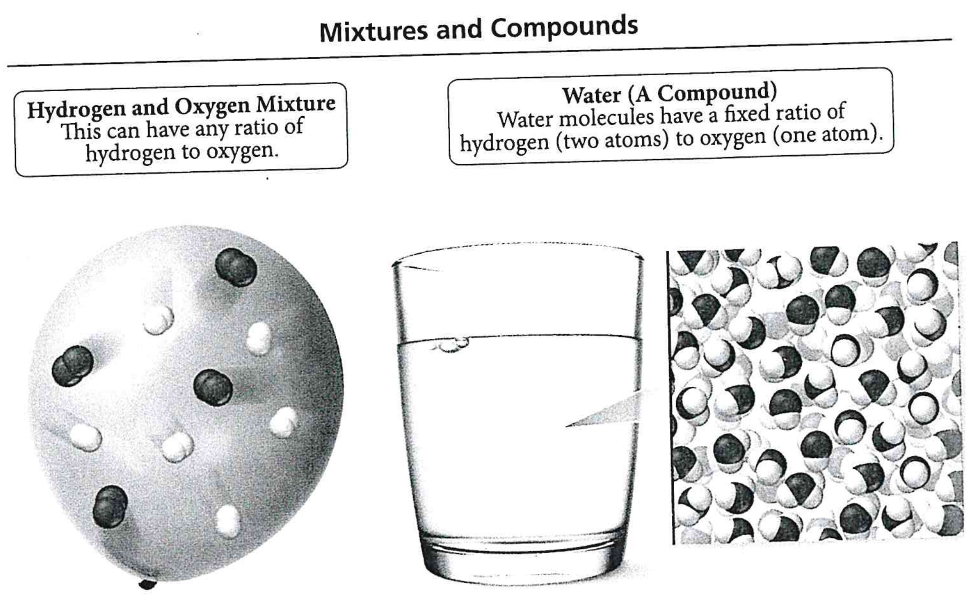
Selected Properties	Hydrogen	Oxygen	Water
Boiling Point	−253 °C	−183 °C	100 °C
State at Room Temperature	Gas	Gas	Liquid
Flammability	Explosive	Necessary for combustion	Used to extinguish flame

First of all, water is a liquid rather than a gas at room temperature, and its boiling point is hundreds of degrees higher than the boiling points of hydrogen and oxygen. Second, instead of being flammable (like hydrogen gas) or supporting combustion (like oxygen gas), water actually smothers flames. Water is nothing like the hydrogen and oxygen from which it forms. The dramatic difference between the elements hydrogen and oxygen and the compound water is typical of the differences between elements and the compounds that they form. *When two or more elements combine to form a compound, an entirely new substance results.*

Consider as another example common table salt, a highly stable compound composed of sodium and chlorine. Elemental sodium, by contrast, is a highly reactive, silvery metal that can explode on contact with water. Elemental chlorine is a corrosive, greenish-yellow gas that can be fatal if inhaled. Yet the compound formed from the combination of these two elements is sodium chloride (or table salt), a flavor enhancer that tastes great on steak.

Although some of the substances that we encounter in everyday life are elements, most are compounds. As we discussed in Chapter 1, a compound is different from a mixture of elements. In a compound, elements combine in fixed, definite proportions; in a mixture, elements can mix in any proportions whatsoever. Consider the difference between a hydrogen–oxygen mixture and water in Figure 3.1 ▼. A hydrogen–oxygen mixture can have any proportions of hydrogen and oxygen gas. Water, by contrast, is composed of water molecules that always contain two hydrogen atoms to every one oxygen atom. Water has a definite proportion of hydrogen to oxygen.

In this chapter you will learn about compounds: how to represent them, how to name them, how to distinguish between their different types, and how to write chemical equations showing how they form and change. You will also learn how to quantify the elemental composition of a compound. This is important in determining how much of a particular element is contained within a particular compound. For example, patients with high blood pressure (hypertension) often have to reduce their sodium ion intake. The sodium ion is normally consumed in the form of sodium chloride, so a hypertension patient needs to know



▶ FIGURE 3.1 **Mixtures and Compounds** The balloon in this illustration is filled with a mixture of hydrogen gas and oxygen gas. The proportions of hydrogen and oxygen are variable. The glass is filled with water, a compound of hydrogen and oxygen. The ratio of hydrogen to oxygen in water is fixed: water molecules always have two hydrogen atoms for each oxygen atom.

how much sodium is present in a given amount of sodium chloride. Similarly, an iron-mining company needs to know how much iron it can recover from a given amount of iron ore. This chapter provides the tools to understand and answer these kinds of questions.

3.2 Chemical Bonds

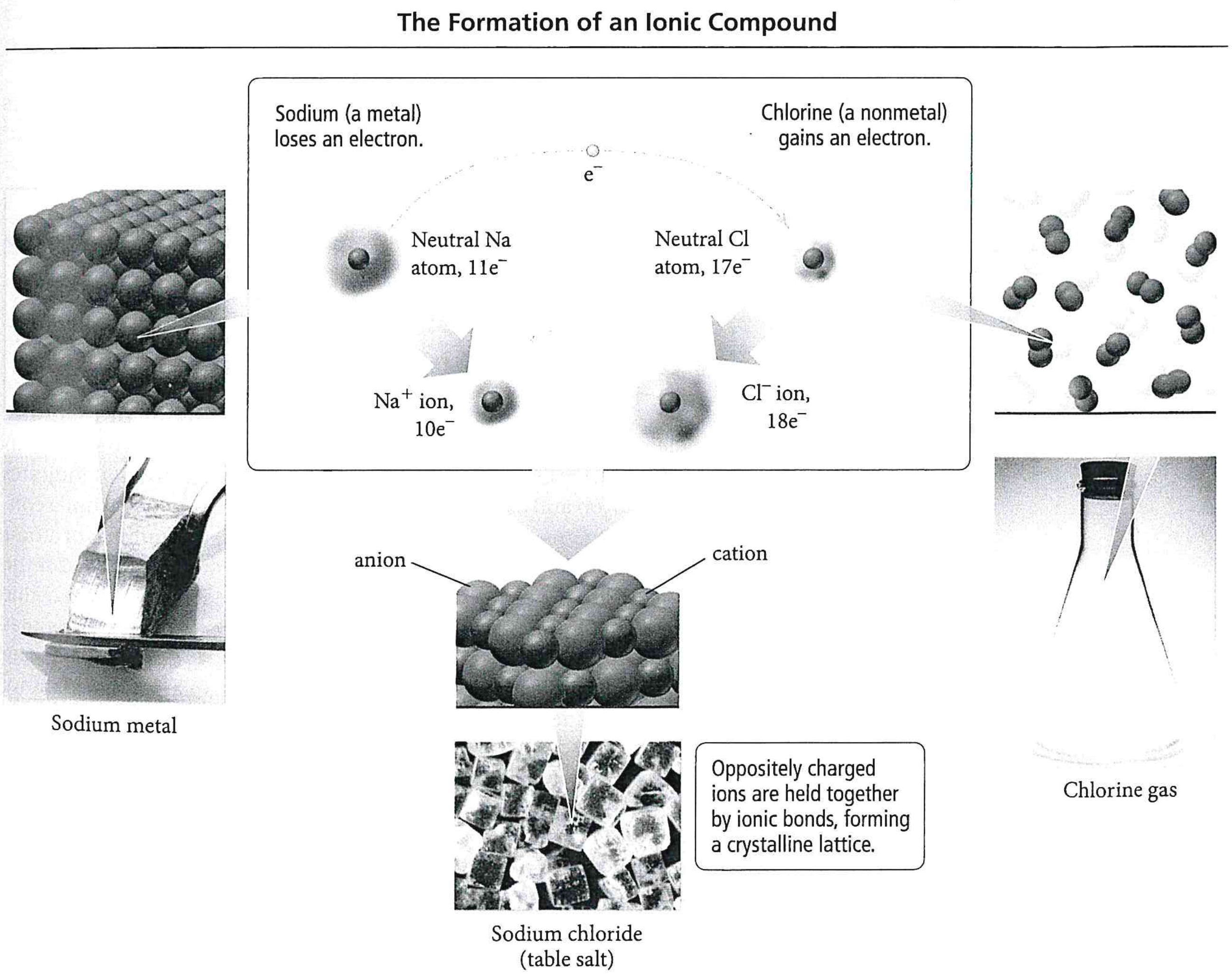
Compounds are composed of atoms held together by *chemical bonds.* Chemical bonds form because of the attractions between the charged particles (the electrons and protons) that compose atoms. We discuss these interactions in more detail in Chapter 9 (see Section 9.2). For now, remember that, as we discussed in Section 2.4, charged particles exert electrostatic forces on one another: like charges repel and opposite charges attract. These forces are responsible for chemical bonding.

We can broadly classify most chemical bonds into two types: ionic and covalent. *Ionic bonds*—which occur between metals and nonmetals—involve the *transfer* of electrons from one atom to another. *Covalent bonds*—which occur between two or more nonmetals—involve the *sharing* of electrons between two atoms.

Ionic Bonds

Recall from Chapter 2 that metals have a tendency to lose electrons and that nonmetals have a tendency to gain them. Therefore, when a metal interacts with a nonmetal, it can transfer one or more of its electrons to the nonmetal. The metal atom then becomes a *cation* (a positively charged ion), and the nonmetal atom becomes an *anion* (a negatively charged ion) as shown in Figure 3.2 ▼. These oppositely charged ions attract one another

▼ FIGURE 3.2 **The Formation of an Ionic Compound** An atom of sodium (a metal) loses an electron to an atom of chlorine (a nonmetal), creating a pair of oppositely charged ions. The sodium cation is attracted to the chloride anion, and the two are held together as part of a crystalline lattice.

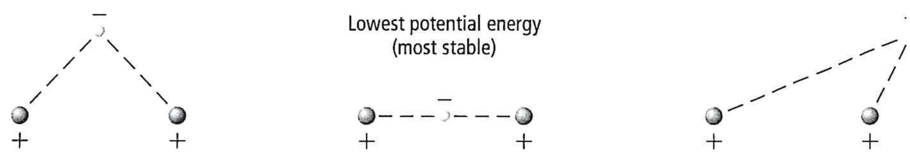


by electrostatic forces and form an **ionic bond**. The result is an **ionic compou** which in the solid phase is composed of a lattice—a regular three-dimensional array– alternating cations and anions.

Covalent Bonds

When a nonmetal bonds with another nonmetal, neither atom transfers its electron the other. Instead the bonding atoms *share* some of their electrons. The shared electr have lower potential energy than they do in the isolated atoms because they inter with the nuclei of both atoms. The bond is a **covalent bond** and the covalently bou atoms compose a *molecule*. Each molecule is independent of the others—the molecu are themselves not covalently bound to one another. Therefore, we call covalen bonded compounds **molecular compounds**.

We can begin to understand the stability of a covalent bond by considering the m stable (or lowest potential energy) configuration of a negative charge interacting w two positive charges (which are separated by some small distance). Figure 3.3 ▼ sho that the lowest potential energy occurs when the negative charge lies *between* the t positive charges because in this arrangement the negative charge can interact with *b positive charges*. Similarly, shared electrons in a covalent chemical bond hold the bond atoms together by attracting the positively charged nuclei of both bonding atoms.



▶ FIGURE 3.3 **The Stability of a Covalent Bond** The potential energy of a negative charge interacting with two positive charges is lowest when the negative charge is between the two positive charges.

CONCEPTUAL CONNECTION **3.1**

Types of Chemical Bonds What type of bond—ionic or covalent—forms between nitrogen and oxygen?

3.3 Representing Compounds: Chemical Formulas and Molecular Models

The quickest and easiest way to represent a compound is with its **chemical formu** which indicates the elements present in the compound and the relative number of ato or ions of each element. For example, H_2O is the chemical formula for water—it indica that water consists of hydrogen and oxygen atoms in a two-to-one ratio. The formula co tains the symbol for each element and a subscript indicating the relative number of ato of the element. A subscript of 1 is typically omitted. Chemical formulas generally list more metallic (or more positively charged) elements first, followed by the less meta (or more negatively charged) elements. Other examples of common chemical formu include NaCl for sodium chloride, indicating sodium and chloride ions in a one-to-c ratio; CO_2 for carbon dioxide, indicating carbon and oxygen atoms in a one-to-two ra and CCl_4 for carbon tetrachloride, indicating carbon and chlorine in a one-to-four rati

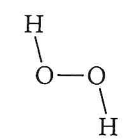
Types of Chemical Formulas

We can categorize chemical formulas into three different types: empirical, molecu and structural. An **empirical formula** gives the *relative* number of atoms of each ment in a compound. A **molecular formula** gives the *actual* number of atoms of ea element in a molecule of a compound. For example, the empirical formula for hydro peroxide is HO, but its molecular formula is H_2O_2. The molecular formula is alway whole-number multiple of the empirical formula. For some compounds, the empirical formula and the molecular formula are identical. For example, the empirical and molecular formula for water is H_2O because water molecules contain two hydrogen atoms and one oxygen atom, and no simpler whole-number ratio can express the relative number of hydrogen atoms to oxygen atoms.

A **structural formula** uses lines to represent covalent bonds and shows how atoms in a molecule connect or bond to each other. The structural formula for H_2O_2 is:

$$H—O—O—H$$

We can also write structural formulas to convey a sense of the molecule's geometry. For example, we can write the structural formula for hydrogen peroxide as:



This version of the formula represents the approximate angles between bonds, giving a sense of the molecule's shape.

Structural formulas can also depict the different types of bonds that occur within molecules. For example, consider the structural formula for carbon dioxide:

$$O{=}C{=}O$$

The two lines between each carbon and oxygen atom represent a double bond, which is generally stronger and shorter than a single bond (represented by a single line). A single bond corresponds to one shared electron pair, while a double bond corresponds to two shared electron pairs. We will learn more about single, double, and even triple bonds in Chapter 9.

The type of formula we use depends on how much we know about the compound and how much we want to communicate. A structural formula communicates the most information, while an empirical formula communicates the least.

Structural Formulas Write the structural formula for water.

CONCEPTUAL **3.2**

Example 3.1 Molecular and Empirical Formulas

Write empirical formulas for the compounds represented by the molecular formulas.

(a) C_4H_8 **(b)** B_2H_6 **(c)** CCl_4

SOLUTION

To determine the empirical formula from a molecular formula, divide the subscripts by the greatest common factor (the largest number that divides exactly into all of the subscripts).

(a) For C_4H_8, the greatest common factor is 4. The empirical formula is therefore CH_2.

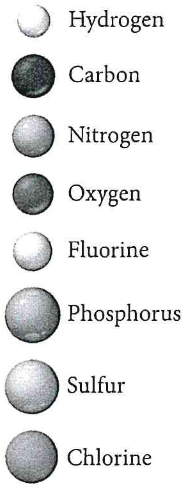
(b) For B_2H_6, the greatest common factor is 2. The empirical formula is therefore BH_3.

(c) For CCl_4, the only common factor is 1, so the empirical formula and the molecular formula are identical.

FOR PRACTICE 3.1 Write the empirical formula for the compounds represented by each molecular formula.

(a) C_5H_{12} **(b)** Hg_2Cl_2 **(c)** $C_2H_4O_2$

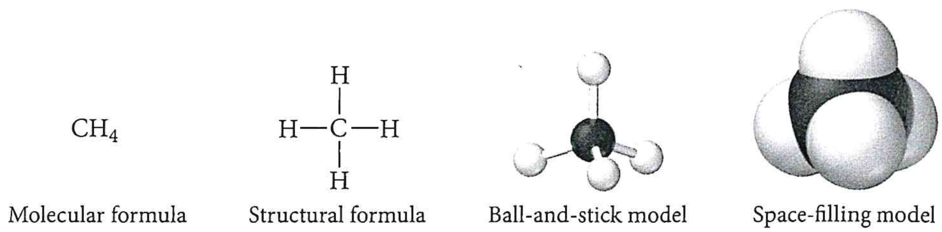
Answers to For Practice and For More Practice problems can be found in Appendix IV.

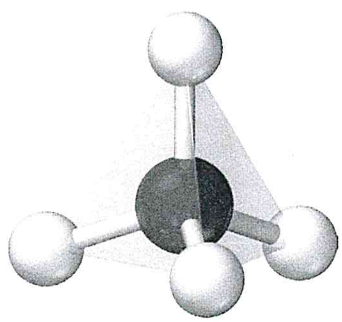


Molecular Models

A *molecular model* is a more accurate and complete way to specify a compound. A **ball-and-stick molecular model** represents atoms as balls and chemical bonds as sticks; how the two connect reflects a molecule's shape. The balls are typically color-coded to specific elements. For example, carbon is customarily black, hydrogen is white, nitrogen is blue, and oxygen is red. (For a complete list of colors of elements in the molecular models used in this book, see Appendix IIA.)

In a **space-filling molecular model**, atoms fill the space between each other to more closely represent our best estimates for how a molecule might appear if scaled to visible size. Consider the following ways to represent a molecule of methane, the main component of natural gas:





▲ A tetrahedron is a three-dimensional geometrical shape characterized by four equivalent triangular faces.

The molecular formula of methane indicates the number and type of each atom in the molecule: one carbon atom and four hydrogen atoms. The structural formula indicates how the atoms connect: the carbon atom bonds to the four hydrogen atoms. The ball-and-stick model clearly portrays the geometry of the molecule: the carbon atom sits in the center of a *tetrahedron* formed by the four hydrogen atoms. And finally, the space-filling model gives the best sense of the relative sizes of the atoms and how they merge together in bonding.

Throughout this book, you will see molecules represented in all of these ways. As you look at these representations, keep in mind what you learned in Chapter 1: the details about a molecule—the atoms that compose it, the lengths of the bonds between atoms, the angles of the bonds between atoms, and its overall shape—determine the properties of the substance that the molecule composes. Change any of these details and those properties change. Table 3.1 shows various compounds represented in the different ways we have just discussed.

CONCEPTUAL CONNECTION **3.3**

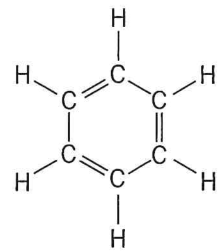
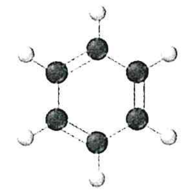
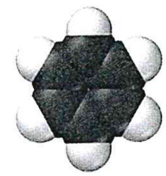
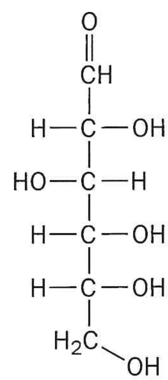
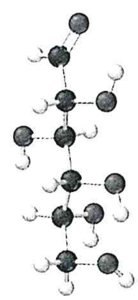
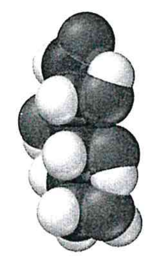
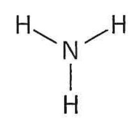
Representing Molecules Based on what you learned in Chapter 2 about atoms, what part of the atom do you think the spheres in the molecular space-filling models shown in Table 3.1 represent? If you were to superimpose a nucleus on one of these spheres, how big would you draw it?

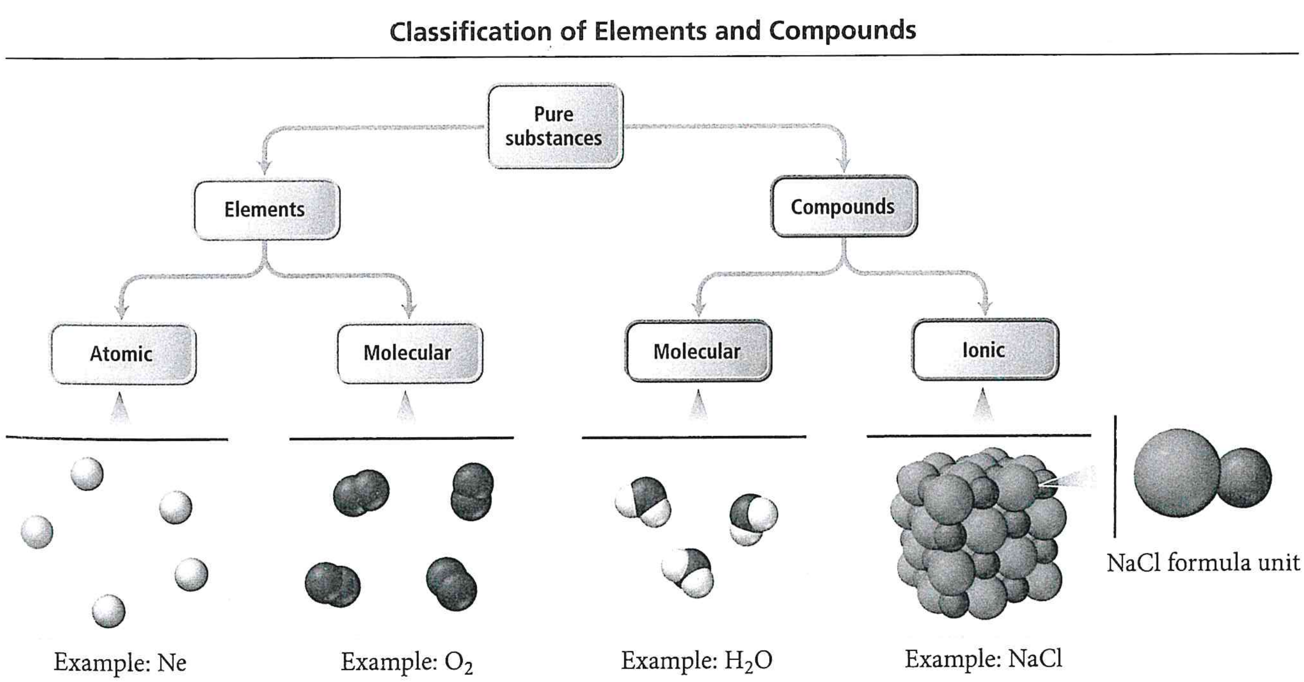
3.4 An Atomic-Level View of Elements and Compounds

Recall from Chapter 1 that we categorize pure substances as either elements or compounds. We can subcategorize elements and compounds according to the basic units that compose them, as shown in Figure 3.4►. Elements may be either atomic or molecular. Compounds may be either molecular or ionic.

Atomic elements exist in nature with single atoms as their basic units. Most elements fall into this category. For example, helium is composed of helium atoms, aluminum is composed of aluminum atoms, and iron is composed of iron atoms. **Molecular elements** do not normally exist in nature with single atoms as their basic units.

TABLE 3.1 Benzene, Acetylene, Glucose, and Ammonia

Name of Compound	Empirical Formula	Molecular Formula	Structural Formula	Ball-and-Stick Model	Space-Filling Model
Benzene	CH	C_6H_6			
Acetylene	CH	C_2H_2	H—C≡C—H		
Glucose	CH_2O	$C_6H_{12}O_6$			
Ammonia	NH_3	NH_3			



▲ FIGURE 3.4 **A Molecular View of Elements and Compounds**

Molecular Elements

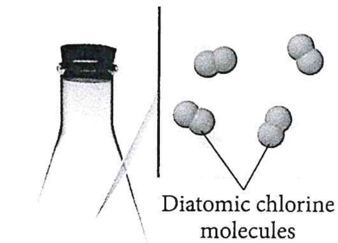
Periods	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 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	57 La	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	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114 Fl	115	116 Lv	117	118

☐ Elements that exist as diatomic molecules

☐ Elements that exist as polyatomic molecules

Lanthanides	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	71 Lu
Actinides	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	103 Lr

▶ FIGURE 3.5 **Molecular Elements** The highlighted elements exist primarily as diatomic molecules (yellow) or polyatomic molecules (red).



▲ The basic units that compose chlorine gas are diatomic chlorine molecules.

Some ionic compounds, such as K_2NaPO_4, contain more than one type of metal ion.

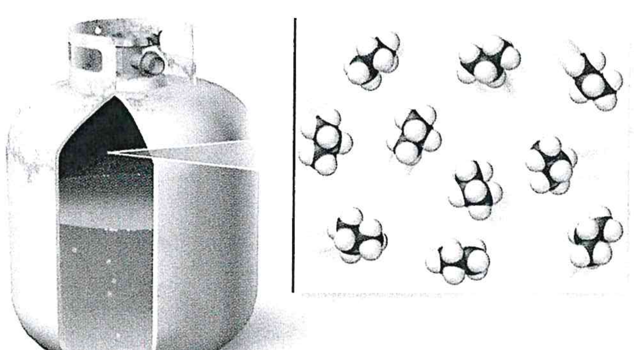
People occasionally refer to formula units as molecules. This is incorrect because ionic compounds do not contain distinct molecules.

Instead, they exist as molecules—two or more atoms of the element bonded together. Most molecular elements exist as *diatomic* molecules. For example, hydrogen is composed of H_2 molecules, nitrogen is composed of N_2 molecules, and chlorine is composed of Cl_2 molecules. A few molecular elements exist as *polyatomic molecules*. Phosphorus exists as P_4, and sulfur exists as S_8. Figure 3.5 ▲ shows the elements that exist primarily as diatomic or polyatomic molecules.

Molecular compounds are usually composed of two or more covalently bonded nonmetals. The basic units of molecular compounds are molecules composed of the constituent atoms. For example, water is composed of H_2O molecules, dry ice is composed of CO_2 molecules, and propane (often used as a fuel for grills) is composed of C_3H_8 molecules as illustrated in Figure 3.6(a) ▼.

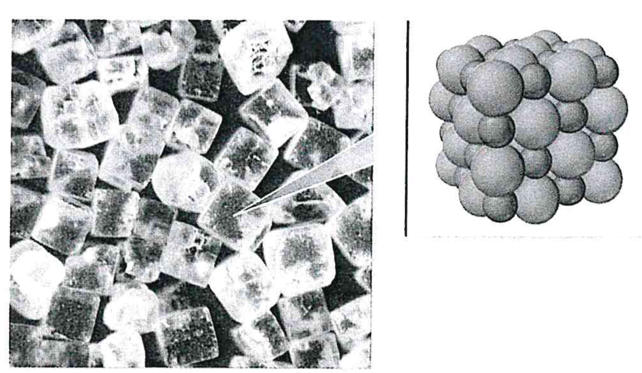
Ionic compounds are usually composed of a metal ionically bonded to one or more nonmetals. The basic unit of an ionic compound is the **formula unit**, the smallest, electrically neutral collection of ions. A formula unit is not a molecule—it does not usually exist as a discrete entity but rather as part of a larger lattice. For example, the ionic

A Molecular Compound



(a)

An Ionic Compound

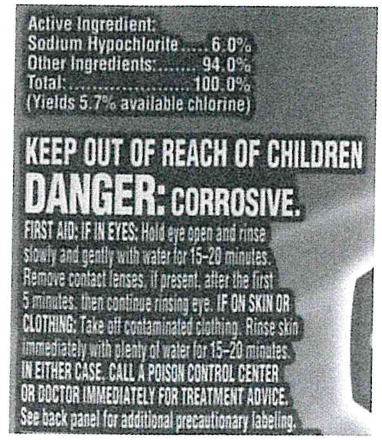


(b)

▲ FIGURE 3.6 **Molecular and Ionic Compounds** **(a)** Propane is a molecular compound. The basic units that compose propane gas are propane (C_3H_8) molecules. **(b)** Table salt (NaCl) is an ionic compound. Its formula unit is the simplest charge-neutral collection of ions: one Na^+ ion and one Cl^- ion.

compound table salt, with the formula unit NaCl, is composed of Na^+ and Cl^- ions in a one-to-one ratio. In table salt, Na^+ and Cl^- ions exist in a three-dimensional alternating array. Because ionic bonds are not directional, no one Na^+ ion pairs with a specific Cl^- ion. Rather, as illustrated in Figure 3.6(b), any one Na^+ cation is surrounded by Cl^- anions and vice versa.

Many common ionic compounds contain ions that are themselves composed of a group of covalently bonded atoms with an overall charge. For example, the active ingredient in household bleach is sodium hypochlorite, which acts to chemically alter color-causing molecules in clothes (bleaching action) and to kill bacteria (disinfection). Hypochlorite is a **polyatomic ion**—an ion composed of two or more atoms—with the formula ClO^-. (Note that the charge on the hypochlorite ion is a property of the whole ion, not just the oxygen atom; this is true for all polyatomic ions.) The hypochlorite ion is often found as a unit in other compounds as well (such as KClO and $Mg(ClO)_2$). Other common compounds that contain polyatomic ions include sodium bicarbonate ($NaHCO_3$), also known as baking soda; sodium nitrite ($NaNO_2$), an inhibitor of bacterial growth in packaged meats; and calcium carbonate ($CaCO_3$), the active ingredient in antacids such as TUMS®.



▲ Polyatomic ions are common in household products such as bleach, which contains sodium hypochlorite (NaClO).

A Molecular View of Elements and Compounds Suppose that the two elements A (represented by triangles) and B (represented by squares) form a molecular compound with the molecular formula A_2B and that two other elements, C (represented by circles) and D (represented by diamonds), form an ionic compound with the formula CD. Draw a molecular-level view of each compound.

CONCEPTUAL CONNECTION 3.4

Example 3.2 Classifying Substances as Atomic Elements, Molecular Elements, Molecular Compounds, or Ionic Compounds

Classify each of the substances as an atomic element, molecular element, molecular compound, or ionic compound.

(a) xenon **(b)** $NiCl_2$ **(c)** bromine **(d)** NO_2 **(e)** $NaNO_3$

SOLUTION

(a) Xenon is an element. It is not a molecular element (see Figure 3.5); therefore, it is an atomic element.

(b) $NiCl_2$ is a compound composed of a metal (nickel is on the left side of the periodic table) and nonmetal (chlorine is on the right side of the periodic table); therefore, it is an ionic compound.

(c) Bromine is one of the elements that exists as a diatomic molecule (see Figure 3.5); therefore, it is a molecular element.

(d) NO_2 is a compound composed of a nonmetal and a nonmetal; therefore, it is a molecular compound.

(e) $NaNO_3$ is a compound composed of a metal and a polyatomic ion; therefore, it is an ionic compound.

FOR PRACTICE 3.2 Classify each of the substances as an atomic element, molecular element, molecular compound, or ionic compound.

(a) fluorine **(b)** N_2O **(c)** silver **(d)** K_2O **(e)** Fe_2O_3

CONCEPTUAL CONNECTION 3.5

Ionic and Molecular Compounds Which statement best summarizes the difference between ionic and molecular compounds?

(a) Molecular compounds contain highly directional covalent bonds, which result in the formation of molecules—discrete particles that do not covalently bond to each other. Ionic compounds contain nondirectional ionic bonds, which result (in the solid phase) in the formation of ionic lattices—extended networks of alternating cations and anions.

(b) Molecular compounds contain covalent bonds in which one of the atoms shares an electron with the other one, resulting in a new force that holds the atoms together in a covalent molecule. Ionic compounds contain ionic bonds in which one atom donates an electron to the other, resulting in a new force that holds the ions together in pairs (in the solid phase).

(c) The key difference between ionic and covalent compounds is the types of elements that compose them, not the way that the atoms bond together.

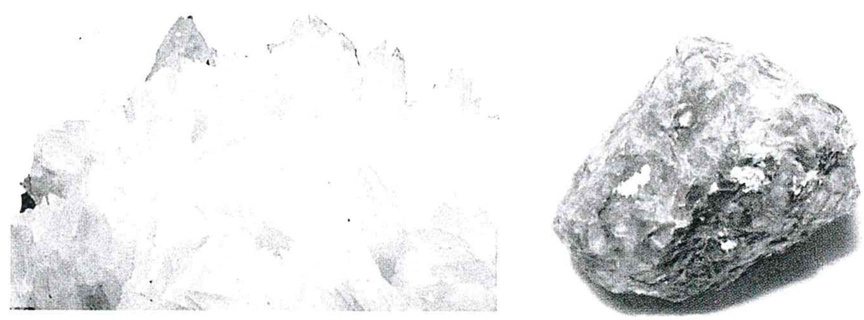
(d) A molecular compound is composed of covalently bonded molecules. An ionic compound is composed of ionically bonded molecules (in the solid phase).

KCV **Naming Ionic Compounds**

▲ Ionic compounds are common in food and consumer products such as reduced-sodium salt (a mixture of NaCl and KCl) and TUMS®($CaCO_3$).

3.5 Ionic Compounds: Formulas and Names

Ionic compounds occur throughout Earth's crust as minerals. Examples include limestone ($CaCO_3$), a type of sedimentary rock; gibbsite [$Al(OH)_3$], a mineral; and soda ash (Na_2CO_3), a natural deposit. We can also find ionic compounds in the foods that we eat. Examples include sodium chloride (NaCl), which is table salt; calcium carbonate ($CaCO_3$), a source of calcium necessary for bone health; and potassium chloride (KCl), a source of potassium necessary for fluid balance and muscle function. Ionic compounds are generally very stable because the attractions between cations and anions within ionic compounds are strong and because each ion interacts with several oppositely charged ions in the crystalline lattice.



▶ Calcite (left) is the main component of limestone, marble, and other forms of calcium carbonate ($CaCO_3$) commonly found in Earth's crust. Trona (right) is a crystalline form of hydrated sodium carbonate ($Na_3H(CO_3)_2 \cdot 2\ H_2O$).

Writing Formulas for Ionic Compounds

See Figure 2.13 to review the elements that form ions with a predictable charge.

Since ionic compounds are charge-neutral, and since many elements form only one type of ion with a predictable charge, we can deduce the formulas for many ionic compounds from their constituent elements. For example, the formula for the ionic compound composed of sodium and chlorine must be NaCl because in compounds Na always forms 1+ cations and Cl always forms 1− anions. In order for the compound to be charge-neutral, it must contain one Na^+ cation for every one Cl^- anion. The formula for the ionic compound composed of *calcium* and chlorine, however, is $CaCl_2$ because Ca always forms 2+ cations and Cl always forms 1− anions. In order for this compound to be charge-neutral, it must contain one Ca^{2+} cation for every two Cl^- anions.

Summarizing Ionic Compound Formulas:

- Ionic compounds always contain positive and negative ions.
- In a chemical formula, the sum of the charges of the positive ions (cations) must equal the sum of the charges of the negative ions (anions).
- The formula of an ionic compound reflects the smallest whole-number ratio of ions.

To write the formula for an ionic compound, follow the procedure in the left column. Two examples of how to apply the procedure are provided in the center and right columns.

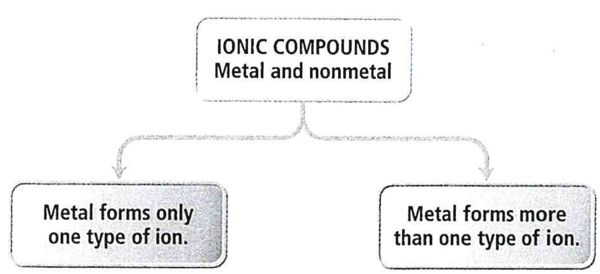
Procedure for...

Writing Formulas for Ionic Compounds	Example 3.3 Writing Formulas for Ionic Compounds Write the formula for the ionic compound that forms between aluminum and oxygen.	Example 3.4 Writing Formulas for Ionic Compounds Write the formula for the ionic compound that forms between calcium and oxygen.
1. Write the symbol for the metal cation and its charge followed by the symbol for the nonmetal anion and its charge. Determine charges from the element's group number in the periodic table (refer to Figure 2.13).	Al^{3+} O^{2-}	Ca^{2+} O^{2-}
2. Adjust the subscript on each cation and anion to balance the overall charge.	Al^{3+} O^{2-} ↓ Al_2O_3	Ca^{2+} O^{2-} ↓ CaO
3. Check that the sum of the charges of the cations equals the sum of the charges of the anions.	cations: 2(3+) = 6+ anions: 3(2−) = 6− The charges cancel.	cations: 2+ anions: 2− The charges cancel.
	FOR PRACTICE 3.3 Write the formula for the compound formed between potassium and sulfur.	FOR PRACTICE 3.4 Write the formula for the compound formed between aluminum and nitrogen.

Naming Ionic Compounds

Some ionic compounds—such as NaCl (table salt) and $NaHCO_3$ (baking soda)—have **common names**, which are nicknames of sorts learned by familiarity. However, chemists have developed **systematic names** for different types of compounds including ionic ones. Even if you are not familiar with a compound, you can determine its systematic name from its chemical formula. Conversely, you can deduce the formula of a compound from its systematic name.

The first step in naming an ionic compound is to identify it as one. Remember, *ionic compounds are usually composed of metals and nonmetals*; any time you see a metal and one or more nonmetals together in a chemical formula, assume that you have an ionic compound. Ionic compounds can be categorized into two types, depending on the metal in the compound. The first type contains a metal whose charge is invariant from one compound to another. Whenever the metal in this first type of compound forms an ion, the ion always has the same charge.

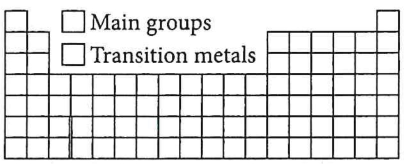


Since the charge of the metal in this first type of ionic compound is always the same, it need not be specified in the name of the compound. Sodium, for instance, has a 1+ charge in all of its

Metals Whose Charge Is Invariant from One Compound to Another

Periods	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 1+	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na 1+	12 Mg 2+											13 Al 3+	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K 1+	20 Ca 2+	21 Sc 3+	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn 2+	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb 1+	38 Sr 2+	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag 1+	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs 1+	56 Ba 2+	57 La	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	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114 Fl	115	116 Lv	117	118

▲ FIGURE 3.7 **Metals with Invariant Charges** The metals highlighted in this table form cations with the same charges in all of their compounds. (Note that silver sometimes forms compounds with other charges, but these are rare.)



▲ FIGURE 3.8 **Transition Metals** Metals that can have different charges in different compounds are usually (but not always) transition metals.

compounds. Figure 3.7▲ lists some examples of these types of metals; the charges of these metals can be inferred from their group number in the periodic table.

The second type of ionic compound contains a metal with a charge that can differ in different compounds. In other words, the metal in this second type of ionic compound can form more than one kind of cation (depending on the compound), and its charge must therefore be specified for a given compound. Iron, for instance, forms a 2+ cation in some of its compounds and a 3+ cation in others. Metals of this type are often *transition metals* (Figure 3.8◄). However, some transition metals, such as Zn and Ag, form cations with the same charge in all of their compounds, and some main-group metals, such as Pb and Sn, form more than one type of cation.

Naming Binary Ionic Compounds Containing a Metal That Forms Only One Type of Cation

Binary compounds contain only two different elements. The names of binary ionic compounds take the form:

name of cation (metal) **base name of anion (nonmetal) + *-ide***

For example, the name for KCl consists of the name of the cation, *potassium*, followed by the base name of the anion, *chlor*, with the ending *-ide*. Its full name is *potassium chloride*.

KCl potassium chloride

The name for CaO consists of the name of the cation, *calcium*, followed by the base name of the anion, *ox*, with the ending *-ide*. Its full name is *calcium oxide*.

CaO calcium oxide

The base names for various nonmetals, and their most common charges in ionic compounds, are shown in Table 3.2.

TABLE 3.2 Some Common Monoatomic Anions

Nonmetal	Symbol for Ion	Base Name	Anion Name
Fluorine	F^-	fluor	Fluoride
Chlorine	Cl^-	chlor	Chloride
Bromine	Br^-	brom	Bromide
Iodine	I^-	iod	Iodide
Oxygen	O^{2-}	ox	Oxide
Sulfur	S^{2-}	sulf	Sulfide
Nitrogen	N^{3-}	nitr	Nitride
Phosphorus	P^{3-}	phosph	Phosphide

Example 3.5 **Naming Ionic Compounds Containing a Metal That Forms Only One Type of Cation**

Name the compound $CaBr_2$.

SOLUTION

The cation is *calcium*. The anion is from bromine, which becomes *bromide*. The correct name is *calcium bromide*.

FOR PRACTICE 3.5 Name the compound Ag_3N.

FOR MORE PRACTICE 3.5 Write the formula for rubidium sulfide.

Naming Binary Ionic Compounds Containing a Metal That Forms More Than One Kind of Cation

For these types of metals, the name of the cation is followed by a Roman numeral (in parentheses) that indicates the charge of the metal in that particular compound. For example, we distinguish between Fe^{2+} and Fe^{3+} as follows:

Note that there is no space between the name of the cation and the parenthetical number indicating its charge.

Fe^{2+} iron(II)

Fe^{3+} iron(III)

The full names for compounds containing metals that form more than one kind of cation have the form:

name of cation (metal) **(charge of cation (metal) in Roman numerals in parentheses)** **base name of anion (nonmetal) + *-ide***

You can determine the charge of the metal cation by inference from the sum of the charges of the nonmetal anions—remember that the sum of all the charges in the compound must be zero. Table 3.3 shows some of the metals that form more than one cation and their most common charges. For example, in $CrBr_3$, the charge of chromium must be 3+ in order for the compound to be charge-neutral with three Br^- anions. The cation is named:

Cr^{3+} chromium(III)

The full name of the compound is:

$CrBr_3$ chromium(III) bromide

Similarly, in CuO the charge of copper must be 2+ in order for the compound to be charge-neutral with one O^{2-} anion. The cation is therefore named:

Cu^{2+} copper(II)

The full name of the compound is:

CuO copper(II) oxide

TABLE 3.3 Some Metals That Form Cations with Different Charges

Metal	Ion	Name	Older Name*
Chromium	Cr^{2+}	Chromium(II)	Chromous
	Cr^{3+}	Chromium(III)	Chromic
Iron	Fe^{2+}	Iron(II)	Ferrous
	Fe^{3+}	Iron(III)	Ferric
Cobalt	Co^{2+}	Cobalt(II)	Cobaltous
	Co^{3+}	Cobalt(III)	Cobaltic
Copper	Cu^{+}	Copper(I)	Cuprous
	Cu^{2+}	Copper(II)	Cupric
Tin	Sn^{2+}	Tin(II)	Stannous
	Sn^{4+}	Tin(IV)	Stannic
Mercury	Hg_2^{2+}	Mercury(I)	Mercurous
	Hg^{2+}	Mercury(II)	Mercuric
Lead	Pb^{2+}	Lead(II)	Plumbous
	Pb^{4+}	Lead(IV)	Plumbic

*An older naming system substitutes the names found in this column for the name of the metal and its charge. Under this system, chromium(II) oxide is named chromous oxide. Additionally, the suffix *-ous* indicates the ion with the lesser charge, and *-ic* indicates the ion with the greater charge. We will *not* use the older system in this text.

Example 3.6 Naming Ionic Compounds Containing a Metal That Forms More Than One Kind of Cation

Name the compound $PbCl_4$.

SOLUTION

The charge on Pb must be 4+ for the compound to be charge-neutral with four Cl^- anions. The name for $PbCl_4$ is the name of the cation, *lead*, followed by the charge of the cation in parentheses *(IV)* and the base name of the anion, *chlor*, with the ending *-ide*. The full name is *lead(IV) chloride*.

$PbCl_4$ lead(IV) chloride

FOR PRACTICE 3.6 Name the compound FeS.

FOR MORE PRACTICE 3.6 Write the formula for ruthenium(IV) oxide.

Naming Ionic Compounds Containing Polyatomic Ions

We name ionic compounds that contain a polyatomic ion in the same way as other ionic compounds, except that we use the name of the polyatomic ion whenever it occurs. Table 3.4 lists common polyatomic ions and their formulas. For example, $NaNO_2$ is named according to its cation, Na^+ (*sodium*), and its polyatomic anion, NO_2^- (*nitrite*). Its full name is *sodium nitrite*.

$NaNO_2$ sodium nitrite

$FeSO_4$ is named according to its cation *iron*, its charge *(II)*, and its polyatomic ion *sulfate*. Its full name is *iron(II) sulfate*.

$FeSO_4$ iron(II) sulfate

If the compound contains both a polyatomic cation and a polyatomic anion, use the names of both polyatomic ions. For example, NH_4NO_3 is *ammonium nitrate*.

NH_4NO_3 ammonium nitrate

You should be able to recognize polyatomic ions in a chemical formula, so become familiar with the ions listed in Table 3.4. Most polyatomic ions are **oxyanions**, anions containing oxygen and another element. Notice that when a series of oxyanions contains different numbers of oxygen atoms, they are named systematically according to

TABLE 3.4 Some Common Polyatomic Ions

Name	Formula	Name	Formula
Acetate	$C_2H_3O_2^-$	Hypochlorite	ClO^-
Carbonate	CO_3^{2-}	Chlorite	ClO_2^-
Hydrogen carbonate (or bicarbonate)	HCO_3^-	Chlorate	ClO_3^-
Hydroxide	OH^-	Perchlorate	ClO_4^-
Nitrite	NO_2^-	Permanganate	MnO_4^-
Nitrate	NO_3^-	Sulfite	SO_3^{2-}
Chromate	CrO_4^{2-}	Hydrogen sulfite (or bisulfite)	HSO_3^-
Dichromate	$Cr_2O_7^{2-}$	Sulfate	SO_4^{2-}
Phosphate	PO_4^{3-}	Hydrogen sulfate (or bisulfate)	HSO_4^-
Hydrogen phosphate	HPO_4^{2-}	Cyanide	CN^-
Dihydrogen phosphate	$H_2PO_4^-$	Peroxide	O_2^{2-}
Ammonium	NH_4^+		

the number of oxygen atoms in the ion. If there are only two ions in the series, the one with more oxygen atoms has the ending *-ate* and the one with fewer has the ending *-ite*. For example, NO_3^- is *nitrate* and NO_2^- is *nitrite*.

NO_3^-	nitr*ate*
NO_2^-	nitr*ite*

If there are more than two ions in the series, then the prefixes *hypo-*, meaning *less than*, and *per-*, meaning *more than*, are used. So ClO^- is hypochlorite (less oxygen than chlorite), and ClO_4^- is perchlorate (more oxygen than chlorate).

ClO^-	*hypo*chlor*ite*
ClO_2^-	chlor*ite*
ClO_3^-	chlor*ate*
ClO_4^-	*per*chlor*ate*

Other halides (halogen ions) form similar series with similar names. Thus, IO_3^- is iodate and BrO_3^- is bromate.

Example 3.7 Naming Ionic Compounds That Contain a Polyatomic Ion

Name the compound $Li_2Cr_2O_7$.

SOLUTION

The name for $Li_2Cr_2O_7$ is the name of the cation, *lithium*, followed by the name of the polyatomic ion, *dichromate*. Its full name is *lithium dichromate*.

$Li_2Cr_2O_7$ lithium dichromate

FOR PRACTICE 3.7 Name the compound $Sn(ClO_3)_2$.

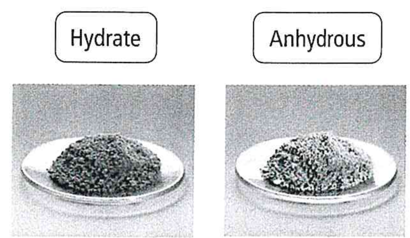
FOR MORE PRACTICE 3.7 Write the formula for cobalt(II) phosphate.

Hydrated Ionic Compounds

The ionic compounds called **hydrates** contain a specific number of water molecules associated with each formula unit. For example, the formula for epsom salts is $MgSO_4 \cdot 7\,H_2O$, and its systematic name is magnesium sulfate heptahydrate. The seven H_2O molecules associated with the formula unit are *waters of hydration.* Waters of hydration can usually be removed by heating the compound. Figure 3.9 ▸ shows a sample of cobalt(II) chloride hexahydrate ($CoCl_2 \cdot 6\,H_2O$) before and after heating. The hydrate is pink and the anhydrous salt (the salt without any associated water molecules) is blue. Hydrates are named just as other ionic compounds, but they are given the additional name "*prefix*hydrate," where the *prefix* indicates the number of water molecules associated with each formula unit.

Common hydrated ionic compounds and their names are as follows:

$CaSO_4 \cdot \frac{1}{2}\,H_2O$	calcium sulfate hemihydrate
$BaCl_2 \cdot 6\,H_2O$	barium chloride hexahydrate
$CuSO_4 \cdot 5\,H_2O$	copper(II) sulfate pentahydrate



$CoCl_2 \cdot 6\,H_2O$ $CoCl_2$

▲ FIGURE 3.9 **Hydrates** Cobalt(II) chloride hexahydrate is pink. Heating the compound removes the waters of hydration, leaving the blue anhydrous cobalt(II) chloride.

Common hydrate prefixes
hemi = 1/2
mono = 1
di = 2
tri = 3
tetra = 4
penta = 5
hexa = 6
hepta = 7
octa = 8

3.6 Molecular Compounds: Formulas and Names

In contrast to ionic compounds, the formula for a molecular compound *cannot* readily be determined from its constituent elements because the same combination of elements may form many different molecular compounds, each with a different formula. For example, carbon and oxygen form both CO and CO_2, and hydrogen and oxygen form both H_2O and H_2O_2. Nitrogen and oxygen form all of the following molecular compounds: NO, NO_2, N_2O, N_2O_3, N_2O_4, and N_2O_5. In Chapter 9, we will discuss the stability of these

KEY **Naming Molecular Compounds**

For example, $HCl(aq)$ is hydro*chlor*ic acid and $HBr(aq)$ is hydro*brom*ic acid.

$HCl(aq)$ hydrochloric acid $HBr(aq)$ hydrobromic acid

Example 3.9 Naming Binary Acids

Name the acid $HI(aq)$.

SOLUTION

The base name of I is *iod*, so $HI(aq)$ is hydroiodic acid.

$HI(aq)$ hydroiodic acid

FOR PRACTICE 3.9 Name the acid $HF(aq)$.

Naming Oxyacids

Oxyacids contain hydrogen and an oxyanion (an anion containing a nonmetal an oxygen). The common oxyanions are listed in the table of polyatomic ions (Table 3.4) For example, $HNO_3(aq)$ contains the nitrate (NO_3^-) ion, $H_2SO_3(aq)$ contains the sulfit (SO_3^{2-}) ion, and $H_2SO_4(aq)$ contains the sulfate (SO_4^{2-}) ion. Oxyacids are a combina tion of one or more H^+ ions with an oxyanion. The number of H^+ ions depends on th

Chemistry IN THE Environment Acid Rain

Certain pollutants—such as NO, NO_2, and SO_2—form acids when mixed with water. NO and NO_2, primarily emitted in vehicular exhaust, combine with atmospheric oxygen and water to form nitric acid, $HNO_3(aq)$. SO_2, emitted primarily from coal-powered electricity generation, combines with atmospheric oxygen and water to form sulfuric acid, $H_2SO_4(aq)$. Both $HNO_3(aq)$ and $H_2SO_4(aq)$ result in acidic rainwater. The problem is greatest in the northeastern United States where pollutants from midwestern electrical power plants combine with rainwater to produce rain that is up to ten times more acidic than normal.

Acid rain can fall or flow into lakes and streams, making these bodies of water more acidic. Some species of aquatic animals—such as trout, bass, snails, salamanders, and clams—cannot tolerate the increased acidity and die. This in turn disturbs the ecosystem of the lake, resulting in imbalances that may lead to the death of other aquatic species. Acid rain also weakens trees by dissolving and washing away nutrients in the soil and by damaging leaves. Appalachian red spruce trees have been the hardest hit, with many forests showing significant acid rain damage.

In addition, acid rain degrades building materials because acids dissolve iron, the main component of steel, and $CaCO_3$ (limestone), a main component of marble and concrete. Consequently, acid rain has damaged many statues, buildings, and bridges in the northeastern United States.

Acid rain has been a problem for many years, but legislation passed toward the end of the last century has begun to address this issue. In 1990, Congress passed several amendments to the Clean Air Act that included provisions requiring electrical utilities to lower SO_2 emissions. Since then, SO_2 emissions have decreased and rain in the northeastern United States has become less acidic. With time, and with continued enforcement of the acid rain regulation, lakes, streams, and forests damaged by acid rain should recover.

QUESTION

Name each compound: NO, NO_2, SO_2, H_2SO_4, HNO_3, $CaCO_3$.

A forest damaged by acid rain.

Acid rain damages building materials including the limestone that composes many statues.

charge of the oxyanion; the formula is always charge-neutral. The names of oxyacids depend on the ending of the oxyanion and take the following forms:

oxyanions ending with *-ate*: **base name of oxyanion + *-ic*** | **acid**

oxyanions ending with *-ite*: **base name of oxyanion + *-ous*** | **acid**

For example, $HNO_3(aq)$ is nitric acid (oxyanion is nitrate), and $H_2SO_3(aq)$ is sulfurous acid (oxyanion is sulfite).

$HNO_3(aq)$ nitric acid $H_2SO_3(aq)$ sulfurous acid

Example 3.10 Naming Oxyacids

Name the acid $HC_2H_3O_2(aq)$.

SOLUTION

The oxyanion is acetate, which ends in *-ate*; therefore, the name of the acid is *acetic acid.*

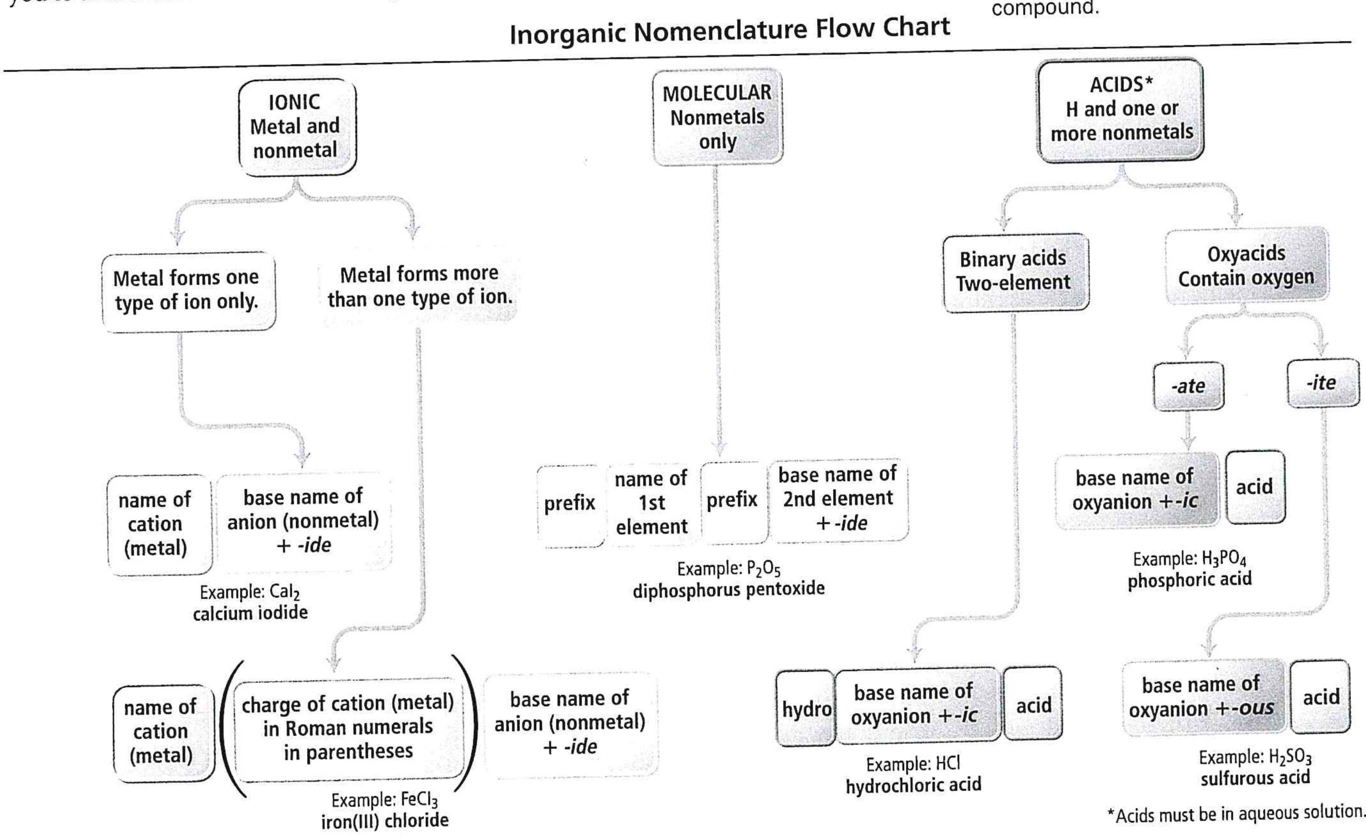
$HC_2H_3O_2(aq)$ acetic acid

FOR PRACTICE 3.10 Name the acid $HNO_2(aq)$.

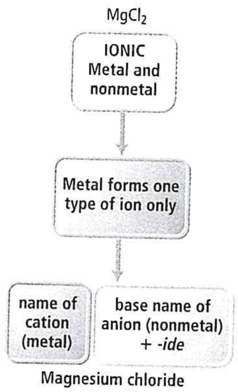
FOR MORE PRACTICE 3.10 Write the formula for perchloric acid.

3.7 Summary of Inorganic Nomenclature

In Section 3.5 and Section 3.6, we discussed naming inorganic compounds, specifically ionic compounds, molecular compounds, and acids. However, we often have to name a compound without initially knowing the category into which it falls. In other words, real-life nomenclature is a bit messier than the categorized nomenclature we just worked through. Figure 3.11 summarizes inorganic nomenclature in a flowchart that will help you to tackle nomenclature from beginning to end.



▼ FIGURE 3.11 **Inorganic Nomenclature Flowchart** The chart summarizes how to name inorganic compounds. Begin by determining if the compound is ionic, molecular, or an acid. Then follow the flowchart for that category from top to bottom until you arrive at a name for the compound.



▲ **FIGURE 3.12 Flowchart Path for $MgCl_2$**

To use the flowchart, begin by determining what type of compound you are ing to name. For example, to name the compound $MgCl_2$, you need to decide if compound is ionic, molecular, or an acid. In this case, since $MgCl_2$ is composed metal and nonmetal, it is ionic. Therefore, you begin at the box labeled "IONIC the far left side of the flowchart.

Next, decide whether the metal in the compound forms only one type of ior more than one type. You can determine this by looking for the metal (in this case m nesium) in Figure 3.7. Since magnesium is listed in the figure, it forms only one typ ion; therefore, you take the left branch in the flowchart as shown in Figure 3.12◄.

Finally, name the compound according to the blocks at the end of the path in flowchart. In this case, write the name of the cation (the metal) followed by the b name of the anion (the nonmetal) appended with the ending *-ide*. Its full name is m nesium chloride.

Example 3.11 Using the Nomenclature Flowchart to Name Compounds

Use the flowchart in Figure 3.11 to name each compound.

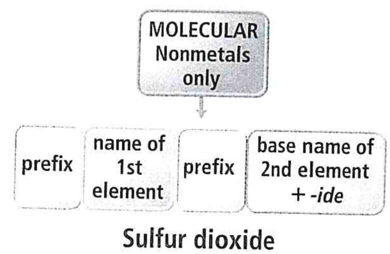
(a) SO_2 **(b)** $HClO_4(aq)$ **(c)** CoF_2

SOLUTION

(a) SO_2

Begin by determining whether the compound is ionic, molecular, or an acid. SO_2 contains only nonmetals; therefore it is molecular.

Name the compound as the name of the first element, *sulfur* (no prefix since the prefix is dropped for mono), followed by the base name of the second element, *ox*, prefixed by *di-* to indicate two, and given the suffix *-ide*.



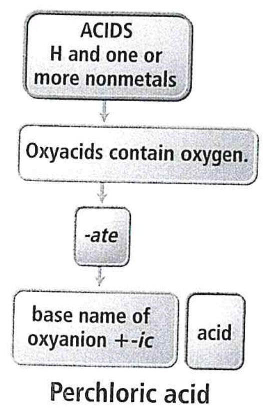
(b) $HClO_4(aq)$

Begin by determining whether the compound is ionic, molecular, or an acid. Since $HClO_4(aq)$ contains H and one more nonmetal and is designated as aqueous, it is an acid.

Next determine whether the acid contains oxygen. Since $HClO_4$ contains oxygen, it is an oxyacid.

Then determine whether the name of the oxyanion ends in *-ate* or *-ite*. Since the oxyanion is perchlor*ate*, it ends in *-ate*.

Finally, name the acid as the base name of the oxyanion, *perchlor*, with the ending *-ic*, followed by the word *acid*.

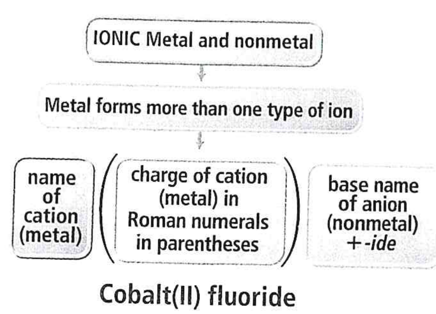


(c) CoF_2

Begin by determining whether the compound is ionic, molecular, or an acid. Since CoF_2 contains a metal and a nonmetal, it is ionic.

Next refer to Figure 3.7 to determine whether the metal forms one type of ion or more than one type. Since Co is not listed in Figure 3.7, it must form more than one type of ion.

Name the compound as the name of the cation, *cobalt*, followed by the charge of the cation in parentheses *(II)*, and the base name of the anion, *fluor*, with the ending *-ide*.



FOR PRACTICE 3.11 Use the flowchart in Figure 3.11 to name $H_2SO_3(aq)$.

3.8 Formula Mass and the Mole Concept for Compounds

In Chapter 2, we defined the average mass of an atom of an element as its *atomic mass*. Similarly, we now define the average mass of a molecule (or a formula unit) of a compound as its **formula mass**. (The common terms *molecular mass* and *molecular weight* are synonymous with formula mass.) For any compound, the formula mass is the sum of the atomic masses of all the atoms in its chemical formula.

Remember, ionic compounds do not contain individual molecules. In casual language, the smallest electrically neutral collection of ions is sometimes called a molecule but is more correctly called a formula unit.

$$\text{Formula mass} = \left(\begin{array}{c}\text{Number of atoms}\\ \text{of 1st element in}\\ \text{chemical formula}\end{array} \times \begin{array}{c}\text{Atomic mass}\\ \text{of}\\ \text{1st element}\end{array}\right) + \left(\begin{array}{c}\text{Number of atoms}\\ \text{of 2nd element in}\\ \text{chemical formula}\end{array} \times \begin{array}{c}\text{Atomic mass}\\ \text{of}\\ \text{2nd element}\end{array}\right) + \ldots$$

For example, the formula mass of carbon dioxide, CO_2, is:

Multiply by 2 because formula has two oxygen atoms.

$$\begin{aligned}\text{Formula mass} &= 12.01\text{ amu} + 2(16.00\text{ amu})\\ &= 44.01\text{ amu}\end{aligned}$$

and that of sodium oxide, Na_2O, is:

Multiply by 2 because formula has two sodium atoms.

$$\begin{aligned}\text{Formula mass} &= 2(22.99\text{ amu}) + 16.00\text{ amu}\\ &= 61.98\text{ amu}\end{aligned}$$

Example 3.12 Calculating Formula Mass

Calculate the formula mass of glucose, $C_6H_{12}O_6$.

SOLUTION

To find the formula mass, add the atomic masses of each atom in the chemical formula.

$$\begin{aligned}\text{Formula mass} &= 6 \times (\text{atomic mass C}) + 12 \times (\text{atomic mass H}) + 6 \times (\text{atomic mass O})\\ &= 6(12.01\text{ amu}) + 12(1.008\text{ amu}) + 6(16.00\text{ amu})\\ &= 180.16\text{ amu}\end{aligned}$$

FOR PRACTICE 3.12 Calculate the formula mass of calcium nitrate.

Molar Mass of a Compound

In Chapter 2 (Section 2.9), we saw that an element's molar mass—the mass in grams of one mole of its atoms—is numerically equivalent to its atomic mass. We then used the molar mass in combination with Avogadro's number to determine the number of atoms in a given mass of the element. We can apply the same concept to compounds. The *molar mass of a compound*—the mass in grams of 1 mol of its molecules or formula units—is numerically equivalent to its formula mass. For example, we just calculated the formula mass of CO_2 to be 44.01 amu. The molar mass is, therefore:

$$CO_2\text{ molar mass} = 44.01\text{ g/mol}$$

Using Molar Mass to Count Molecules by Weighing

The molar mass of CO_2 is a conversion factor between mass (in grams) and amount (in moles) of CO_2. Suppose we want to find the number of CO_2 molecules in a sample of dry ice (solid CO_2) with a mass of 10.8 g. This calculation is analogous to Example 2.8, where we found the number of atoms in a sample of copper of a given mass. We begin with the mass of 10.8 g and use the molar mass to convert to the amount in moles. Then we us Avogadro's number to convert to number of molecules. The conceptual plan is as follow

Conceptual Plan

$$\text{g } CO_2 \xrightarrow{\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}} \text{mol } CO_2 \xrightarrow{\frac{6.022 \times 10^{23}\ CO_2 \text{ molecules}}{1 \text{ mol } CO_2}} CO_2 \text{ molecules}$$

To solve the problem, we follow the conceptual plan, beginning with 10.8 g CO, converting to moles, and then to molecules.

Solution

$$10.8 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{6.022 \times 10^{23}\ CO_2 \text{ molecules}}{1 \text{ mol } CO_2}$$
$$= 1.48 \times 10^{23}\ CO_2 \text{ molecules}$$

Example 3.13 The Mole Concept—Converting between Mass and Number of Molecules

An aspirin tablet contains 325 mg of acetylsalicylic acid ($C_9H_8O_4$). How many acetylsalicylic acid molecules does it contain?

SORT You are given the mass of acetylsalicylic acid and asked to find the number of molecules.

GIVEN: 325 mg $C_9H_8O_4$

FIND: number of $C_9H_8O_4$ molecules

STRATEGIZE First convert to moles (using the molar mass of the compound) and then to number of molecules (using Avogadro's number). You need both the molar mass of acetylsalicylic acid and Avogadro's number as conversion factors. You also need the conversion factor between g and mg.

CONCEPTUAL PLAN

$$\text{mg } C_9H_8O_4 \xrightarrow{\frac{10^{-3} \text{ g}}{1 \text{ mg}}} \text{g } C_9H_8O_4 \xrightarrow{\frac{1 \text{ mol } C_9H_8O_4}{180.15 \text{ g } C_9H_8O_4}}$$

$$\text{mol } C_9H_8O_4 \xrightarrow{\frac{6.022 \times 10^{23}\ C_9H_8O_4 \text{ molecules}}{1 \text{ mol } C_9H_8O_4}} \text{number of } C_9H_8O_4 \text{ molecules}$$

RELATIONSHIPS USED

$$C_9H_8O_4 \text{ molar mass} = 9(12.01) + 8(1.008) + 4(16.00) = 180.15 \text{ g/mol}$$

$6.022 \times 10^{23} = 1$ mol

$1 \text{ mg} = 10^{-3}$ g

SOLVE Follow the conceptual plan to solve the problem.

SOLUTION

$$325 \text{ mg } C_9H_8O_4 \times \frac{10^{-3} \text{ g}}{1 \text{ mg}} \times \frac{1 \text{ mol } C_9H_8O_4}{180.15 \text{ g } C_9H_8O_4} \times \frac{6.022 \times 10^{23}\ C_9H_8O_4 \text{ molecules}}{1 \text{ mol } C_9H_8O_4} = 1.09 \times 10^{21}\ C_9H_8O_4 \text{ molecules}$$

CHECK The units of the answer, $C_9H_8O_4$ molecules, are correct. The magnitude is smaller than Avogadro's number, as expected, since you have less than 1 molar mass of acetylsalicylic acid.

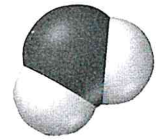
FOR PRACTICE 3.13 Find the number of ibuprofen molecules in a tablet containing 200.0 mg of ibuprofen ($C_{13}H_{18}O_2$).

FOR MORE PRACTICE 3.13 Determine the mass of a sample of water containing 3.55×10^{22} H_2O molecules.

CONCEPTUAL CONNECTION 3.7

Molecular Models and the Size of Molecules Throughout this book, you will find space-filling molecular models to represent molecules. Which number is the best estimate for the scaling factor used in these models? In other words, by approximately what number would you have to multiply the radius of an actual oxygen atom to get the radius of the sphere used to represent the oxygen atom in the water molecule shown to the right.

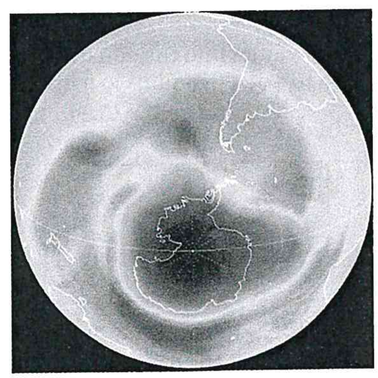
(a) 10 **(b)** 10^4 **(c)** 10^8 **(d)** 10^{16}



3.9 Composition of Compounds

A chemical formula, in combination with the molar masses of its constituent elements, indicates the relative quantities of each element in a compound, which is extremely useful information. For example, about 35 years ago, scientists began to suspect that synthetic compounds known as chlorofluorocarbons (or CFCs) were destroying ozone (O_3) in Earth's upper atmosphere. Upper atmospheric ozone is important because it acts as a shield, protecting life on Earth from the sun's harmful ultraviolet light.

CFCs are chemically inert compounds used primarily as refrigerants and industrial solvents. Over time, CFCs accumulated in the atmosphere. In the upper atmosphere, sunlight breaks bonds within CFCs, releasing chlorine atoms. The chlorine atoms react with ozone, converting it into O_2. So the harmful part of CFCs is the chlorine atoms that they carry. How can we determine the mass of chlorine in a given mass of a CFC?



▲ The chlorine in chlorofluorocarbons caused the ozone hole over Antarctica. The dark blue color indicates depressed ozone levels.

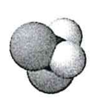
One way to express how much of an element is in a given compound is to use the element's mass percent composition for that compound. The **mass percent composition** or **mass percent** of an element is that element's percentage of the compound's total mass. We calculate the mass percent of element X in a compound from the chemical formula as follows:

$$\text{Mass percent of element X} = \frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100\%$$

Suppose, for example, that we want to calculate the mass percent composition of Cl in the chlorofluorocarbon CCl_2F_2. The mass percent Cl is given by:

CCl_2F_2

$$\text{Mass percent Cl} = \frac{2 \times \text{molar mass Cl}}{\text{molar mass } CCl_2F_2} \times 100\%$$



We multiply the molar mass of Cl by 2 because the chemical formula has a subscript of 2 for Cl, indicating that 1 mol of CCl_2F_2 contains 2 mol of Cl atoms. We calculate the molar mass of CCl_2F_2 as follows:

$$\begin{aligned}\text{Molar mass} &= 12.01\ \text{g/mol} + 2(35.45\ \text{g/mol}) + 2(19.00\ \text{g/mol}) \\ &= 120.91\ \text{g/mol}\end{aligned}$$

So the mass percent of Cl in CCl_2F_2 is:

$$\begin{aligned}\text{Mass percent Cl} &= \frac{2 \times \text{molar mass Cl}}{\text{molar mass } CCl_2F_2} \times 100\% \\ &= \frac{2 \times 35.45\ \text{g/}\cancel{\text{mol}}}{120.91\ \text{g/}\cancel{\text{mol}}} \times 100\% \\ &= 58.64\%\end{aligned}$$

Conversion Factors from Chemical Formulas

Mass percent composition is one way to understand how much chlorine is in a partic chlorofluorocarbon or, more generally, how much of a constituent element is preser a given mass of any compound. However, we can also approach this type of problei a different way. Chemical formulas contain within them inherent relationships betv atoms (or moles of atoms) and molecules (or moles of molecules). For example, the mula for CCl_2F_2 tells us that 1 mol of CCl_2F_2 contains 2 mol of Cl atoms. We write ratio as:

$$1 \text{ mol } CCl_2F_2 : 2 \text{ mol Cl}$$

With ratios such as these—which come from the chemical formula—we can dire determine the amounts of the constituent elements present in a given amount of a c pound without having to calculate mass percent composition. For example, we calcu the number of moles of Cl in 38.5 mol of CCl_2F_2 as follows:

Conceptual Plan

$$\text{mol } CCl_2F_2 \longrightarrow \text{mol Cl}$$

$$\frac{2 \text{ mol Cl}}{1 \text{ mol } CCl_2F_2}$$

Solution

$$38.5 \text{ mol } \cancel{CCl_2F_2} \times \frac{2 \text{ mol Cl}}{1 \text{ mol } \cancel{CCl_2F_2}} = 77.0 \text{ mol Cl}$$

As we have seen, however, we often want to know, not the *amount in moles* of an ment in a certain number of moles of compound, but the *mass in grams* (or other un of a constituent element in a given *mass* of the compound. Suppose we want to kn the mass (in grams) of Cl in 25.0 g CCl_2F_2. The relationship inherent in the chemi formula (2 mol Cl : 1 mol CCl_2F_2) applies to the amount in moles, not to mass. The fore, we first convert the mass of CCl_2F_2 to moles CCl_2F_2. *Then* we use the conversi factor from the chemical formula to convert to moles Cl. Finally, we use the molar m of Cl to convert to grams Cl.

Conceptual Plan

$$\text{g } CCl_2F_2 \longrightarrow \text{mol } CCl_2F_2 \longrightarrow \text{mol Cl} \longrightarrow \text{g Cl}$$

$$\frac{1 \text{ mol } CCl_2F_2}{120.91 \text{ g } CCl_2F_2} \qquad \frac{2 \text{ mol Cl}}{1 \text{ mol } CCl_2F_2} \qquad \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}}$$

Solution

$$25.0 \text{ g } \cancel{CCl_2F_2} \times \frac{1 \text{ mol } \cancel{CCl_2F_2}}{120.91 \text{ g } \cancel{CCl_2F_2}} \times \frac{2 \text{ mol } \cancel{Cl}}{1 \text{ mol } \cancel{CCl_2F_2}} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol } \cancel{Cl}} = 14.7 \text{ g Cl}$$

Notice that we must convert from g CCl_2F_2 to mol CCl_2F_2 *before* we can use t chemical formula as a conversion factor. *Always remember that the chemical formu indicates the relationship between the amounts (in moles) of substances, not between t masses (in grams) of them.*

The general form for solving problems in which we need to find the mass of an el ment present in a given mass of a compound is:

$$\text{Mass compound} \longrightarrow \text{moles compound} \longrightarrow \text{moles element} \longrightarrow \text{mass element}$$

We use the atomic or molar mass to convert between mass and moles, and we use rel tionships inherent in the chemical to convert between moles and moles.

Example 3.16 Chemical Formulas as Conversion Factors

Hydrogen may be used in the future to replace gasoline as a fuel. Most major automobile companies are developing vehicles that run on hydrogen. These cars have the potential to be less environmentally harmful than our current vehicles because their only emission is water vapor. One way to obtain hydrogen for fuel is to use an emission-free energy source such as wind power to form elemental hydrogen from water. What mass of hydrogen (in grams) is contained in 1.00 gallon of water? (The density of water is 1.00 g/mL.)

SORT You are given a volume of water and asked to find the mass of hydrogen it contains. You are also given the density of water.

GIVEN: 1.00 gal H_2O

$d_{H_2O} = 1.00\ g/mL$

FIND: g H

STRATEGIZE The first part of the conceptual plan shows how to convert the units of volume from gallons to liters and then to mL. It also shows how to use the density to convert mL to g.

CONCEPTUAL PLAN

$$\text{gal } H_2O \xrightarrow{\frac{3.785\ L}{1\ gal}} \text{L } H_2O \xrightarrow{\frac{1000\ mL}{1\ L}} \text{mL } H_2O \xrightarrow{\frac{1.00\ g}{1\ mL}} \text{g } H_2O$$

The second part of the conceptual plan is the basic sequence: mass → moles → moles → mass. Convert between moles and mass using the appropriate molar masses, and convert from mol H_2O to mol H using the conversion factor derived from the molecular formula.

$$\text{g } H_2O \xrightarrow{\frac{1\ mol\ H_2O}{18.02\ g\ H_2O}} \text{mol } H_2O \xrightarrow{\frac{2\ mol\ H}{1\ mol\ H_2O}} \text{mol H} \xrightarrow{\frac{1.008\ g\ H}{1\ mol\ H}} \text{g H}$$

RELATIONSHIPS USED

3.785 L = 1 gal (Table 1.3)

1000 mL = 1 L

1.00 g H_2O = 1 mL H_2O (density of H_2O)

Molar mass H_2O = 2(1.008) + 16.00 = 18.02 g/mol

2 mol H : 1 mol H_2O

1.008 g H = 1 mol H

SOLVE Follow the conceptual plan to solve the problem.

SOLUTION

$$1.00\ \text{gal } H_2O \times \frac{3.785\ L}{1\ gal} \times \frac{1000\ mL}{1\ L} \times \frac{1.00\ g}{mL} = 3.7\underline{8}5 \times 10^3\ g\ H_2O$$

$$3.7\underline{8}5 \times 10^3\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} \times \frac{1.008\ g\ H}{1\ mol\ H} = 4.23 \times 10^2\ g\ H$$

CHECK The units of the answer (g H) are correct. Since a gallon of water is about 3.8 L, its mass is about 3.8 kg. H is a light atom, so its mass should be significantly less than 3.8 kg, which it is in the answer.

FOR PRACTICE 3.16 Determine the mass of oxygen in a 7.2-g sample of $Al_2(SO_4)_3$.

FOR MORE PRACTICE 3.16 Butane (C_4H_{10}) is the liquid fuel in lighters. How many grams of carbon are present within a lighter containing 7.25 mL of butane? (The density of liquid butane is 0.601 g/mL.)

Chemical Formulas and Elemental Composition The molecular formula for water is H_2O. Which ratio can be correctly derived from this formula? Explain.

CONCEPTUAL 3.9

(a) 2 g H : 1 g H_2O **(b)** 2 mL H : 1 mL H_2O **(c)** 2 mol H : 1 mol H_2O

AND **Medicine** Methylmercury in Fish

In the last decade, the U.S. Environmental Protection Agency (EPA) has grown increasingly concerned about mercury levels in fish. Mercury—which is present in fish as methylmercury—affects the central nervous system of humans who eat the fish, especially children and developing fetuses. In a developing fetus, excessive mercury exposure can result in slowed mental development and even retardation. Some lakes now have warnings about eating too much fish caught in the lakes.

Recent regulations force fish vendors to alert customers about the dangers of eating too much of certain kinds of commercial fish, including shark, tuna, and mackerel. These fish tend to contain high levels of methylmercury and therefore should be eaten in moderation, especially by children and pregnant women. The U.S. Food and Drug Administration (FDA) *action level*—the level below which the FDA claims the food has no adverse health effects—for methylmercury in fish is 1.0 ppm or 1.0 g of methylmercury per million grams of fish. However, a number of environmental advocacy groups, including the EPA, have suggested that, while this level may be safe for adults, it is too high for children and pregnant women. Consequently, the FDA suggests that pregnant women limit their intake of fish to 12 ounces per week.

▲ Lakes containing mercury—either from natural sources or from pollution—often have posted limits fo the number of fish from the lake that can be eaten safely.

QUESTION

The levels of methylmercury in fish are normally tested by laboratory techniques that measure only the mercury (Hg). Suppose a lab analyzes a 14.5 g sample of fish and finds that it contains 1.03×10^{-5} g of mercury. How much methylmercury ($HgCH_3Cl$) is in the fish in parts per million (ppm)? Is this above the FDA action level?

3.10 Determining a Chemical Formula from Experimental Data

In Section 3.8, we calculated mass percent composition from a chemical formula. Ca also do the reverse? Can we calculate a chemical formula from mass percent composit This question is important because many laboratory analyses of compounds list the tive masses of each element present in the compound. For example, if we decompose w into hydrogen and oxygen in the laboratory, we can measure the masses of hydrogen oxygen produced. Can we determine a chemical formula from this data? The answe qualified yes. We can determine a chemical formula, but it is an empirical formula (r molecular formula). To get a molecular formula, we need additional information, suc the molar mass of the compound.

Suppose we decompose a sample of water in the laboratory and find that it prod 0.857 g of hydrogen and 6.86 g of oxygen. How do we determine an empirical formula f these data? We know that an empirical formula represents a ratio of atoms or a rati moles of atoms, *not a ratio of masses*. So the first thing we must do is convert our data f mass (in grams) to amount (in moles). How many moles of each element are present in sample? To convert to moles, we divide each mass by the molar mass of that element:

$$\text{Moles H} = 0.857\ \text{g H} \times \frac{1\ \text{mol H}}{1.01\ \text{g H}} = 0.849\ \text{mol H}$$

$$\text{Moles O} = 6.86\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 0.429\ \text{mol O}$$

From these data, we know there are 0.849 mol H for every 0.429 mol O. We can write a *pseudoformula* for water:

$$H_{0.849}O_{0.429}$$

To get the smallest whole-number subscripts in our formula, we divide all the subscr by the smallest one, in this case 0.429:

$$H_{\frac{0.849}{0.429}}O_{\frac{0.429}{0.429}} = H_{1.98}O = H_2O$$

Our empirical formula for water, which also happens to be the molecular formula, is H_2O. You can use the procedure shown here to obtain the empirical formula of any compound from experimental data giving the relative masses of the constituent elements. The left column outlines the procedure, and the center and right columns contain two examples of how to apply the procedure.

Procedure for...

Obtaining an Empirical Formula from Experimental Data	Example 3.17 Obtaining an Empirical Formula from Experimental Data	Example 3.18 Obtaining an Empirical Formula from Experimental Data
	A compound containing nitrogen and oxygen is decomposed in the laboratory. It produces 24.5 g nitrogen and 70.0 g oxygen. Calculate the empirical formula of the compound.	A laboratory analysis of aspirin determines the following mass percent composition: C 60.00%; H 4.48%; O 35.52% Find the empirical formula.
1. Write down (or calculate) as *given* the masses of each element present in a sample of the compound. If you are given mass percent composition, assume a 100-g sample and calculate the masses of each element from the given percentages.	GIVEN: 24.5 g N, 70.0 g O FIND: empirical formula	GIVEN: In a 100-g sample: 60.00 g C, 4.48 g H, 35.52 g O FIND: empirical formula
2. Convert each of the masses in step 1 to moles by using the appropriate molar mass for each element as a conversion factor.	$24.5\text{ g N} \times \frac{1\text{ mol N}}{14.01\text{ g N}} = 1.75\text{ mol N}$ $70.0\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 4.38\text{ mol O}$	$60.00\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 4.996\text{ mol C}$ $4.48\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 4.44\text{ mol H}$ $35.52\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 2.220\text{ mol O}$
3. Write down a pseudoformula for the compound using the number of moles of each element (from step 2) as subscripts.	$N_{1.75}O_{4.38}$	$C_{4.996}H_{4.44}O_{2.220}$
4. Divide all the subscripts in the formula by the smallest subscript.	$N_{\frac{1.75}{1.75}}O_{\frac{4.38}{1.75}} \longrightarrow N_1O_{2.5}$	$C_{\frac{4.996}{2.220}}H_{\frac{4.44}{2.220}}O_{\frac{2.220}{2.220}} \longrightarrow C_{2.25}H_2O_1$
5. If the subscripts are not whole numbers, multiply all the subscripts by a small whole number (see table) to get whole-number subscripts.	$N_1O_{2.5} \times 2 \longrightarrow N_2O_5$ The correct empirical formula is N_2O_5.	$C_{2.25}H_2O_1 \times 4 \longrightarrow C_9H_8O_4$ The correct empirical formula is $C_9H_8O_4$.
	FOR PRACTICE 3.17 A sample of a compound is decomposed in the laboratory and produces 165 g carbon, 27.8 g hydrogen, and 220.2 g oxygen. Calculate the empirical formula of the compound.	FOR PRACTICE 3.18 Ibuprofen has the following mass percent composition: C 75.69%, H 8.80%, O 15.51%. What is the empirical formula of ibuprofen?

Fractional Subscript	Multiply by This
0.20	5
0.25	4
0.33	3
0.40	5
0.50	2
0.66	3
0.75	4
0.80	5

Determining Molecular Formulas for Compounds

We can find the molecular formula of a compound from the empirical formula if w know the molar mass of the compound. Recall from Section 3.3 that the molecul mula is always a whole-number multiple of the empirical formula:

$$\text{Molecular formula} = \text{empirical formula} \times n, \text{ where } n = 1, 2, 3, \ldots$$

Suppose we want to find the molecular formula for fructose (a sugar found in from its empirical formula, CH_2O, and its molar mass, 180.2 g/mol. We know th molecular formula is a whole-number multiple of CH_2O:

$$\text{Molecular formula} = (CH_2O) \times n = C_nH_{2n}O_n$$

We also know that the molar mass is a whole-number multiple of the **empirical mula molar mass**, the sum of the masses of all the atoms in the empirical formul

$$\text{Molar mass} = \text{empirical formula molar mass} \times n$$

For a particular compound, the value of n in both cases is the same. Therefore, w find n by calculating the ratio of the molar mass to the empirical formula molar mas

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

For fructose, the empirical formula molar mass is:

$$\text{Empirical formula molar mass} = 12.01 \text{ g/mol} + 2(1.01 \text{ g/mol}) + 16.00 \text{ g/mol} = 30.03 \text{ g}$$

Therefore, n is:

$$n = \frac{180.2 \text{ g/mol}}{30.03 \text{ g/mol}} = 6$$

We can then use this value of n to find the molecular formula:

$$\text{Molecular formula} = (CH_2O) \times 6 = C_6H_{12}O_6$$

Example 3.19 Determining a Molecular Formula from an Empirical Formula and Molar Mas

Butanedione—the component responsible for the smell and taste of butter and cheese—contains the elements carbon, hydrogen, and oxygen. The empirical formula of butanedione is C_2H_3O, and its molar mass is 86.09 g/mol. Determine its molecular formula.

SORT You are given the empirical formula and molar mass of butanedione and asked to find the molecular formula.

GIVEN: Empirical formula = C_2H_3O
molar mass = 86.09 g/mol

FIND: Molecular formula

STRATEGIZE A molecular formula is always a whole-number multiple of the empirical formula. Divide the molar mass by the empirical formula molar mass to find the whole number.

$$\text{Molecular formula} = \text{empirical formula} \times n$$

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

SOLVE Calculate the empirical formula mass.

$$\text{Empirical formula molar mass} = 2(12.01 \text{ g/mol}) + 3(1.008 \text{ g/mol}) + 16.00 \text{ g/mol} = 43.04 \text{ g/m}$$

Divide the molar mass by the empirical formula mass to find n.

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}} = \frac{86.09 \text{ g/mol}}{43.04 \text{ g/mol}} = 2$$

Multiply the empirical formula by n to obtain the molecular formula.

$$\text{Molecular formula} = C_2H_3O \times 2 = C_4H_6O_2$$

CHECK Check the answer by calculating the molar mass of the formula as follows:

$4(12.01 \text{ g/mol}) + 6(1.008 \text{ g/mol}) + 2(16.00 \text{ g/mol}) = 86.09 \text{ g/mol}$

The calculated molar mass is in agreement with the given molar mass.

FOR PRACTICE 3.19 A compound has the empirical formula CH and a molar mass of 78.11 g/mol. What is its molecular formula?

FOR MORE PRACTICE 3.19 Determine the molecular formula for the compound with a molar mass of 60.10 g/mol and the following percent composition:

C, 39.97%

H, 13.41%

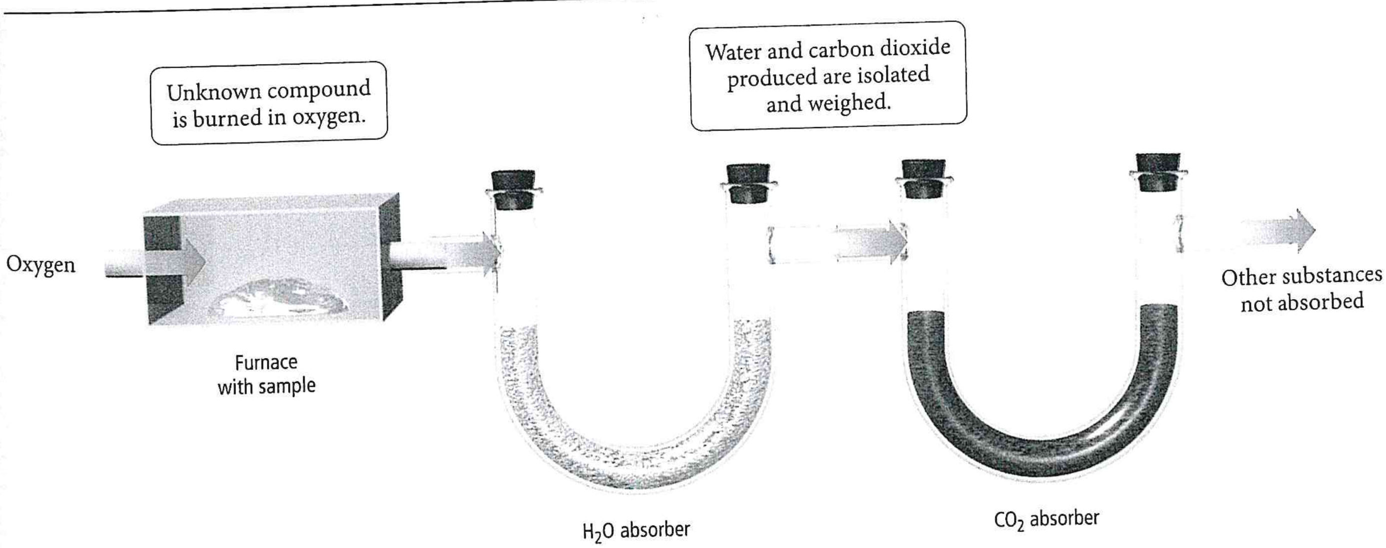
N, 46.62%

Combustion Analysis

In the previous section, we discussed how to determine the empirical formula of a compound from the relative masses of its constituent elements. Another common (and related) way to obtain empirical formulas for unknown compounds, especially those containing carbon and hydrogen, is **combustion analysis**. In combustion analysis, the unknown compound undergoes combustion (or burning) in the presence of pure oxygen, as shown in Figure 3.13 ▼. When the sample burns, all of the carbon converts to CO_2, and all of the hydrogen converts to H_2O. The CO_2 and H_2O are weighed. With these masses, we can use the numerical relationships between moles inherent in the formulas for CO_2 and H_2O (1 mol CO_2 : 1 mol C and 1 mol H_2O : 2 mol H) to determine the amounts of C and H in the original sample. We can determine the amounts of any other elemental constituents, such as O, Cl, or N, by subtracting the sum of the masses of C and H from the original mass of the sample. Examples 20 and 21 illustrate how to perform these calculations for a sample containing only C and H and for a sample containing C, H, and O.

Combustion is a type of *chemical reaction*. We discuss chemical reactions and their representation in Section 3.11.

Combustion Analysis



▲ FIGURE 3.13 **Combustion Analysis Apparatus** The sample is placed in a furnace and burned in oxygen. The water and carbon dioxide produced are absorbed into separate containers and weighed.

Procedure for...

Determining an Empirical Formula from Combustion Analysis	Example 3.20 Determining an Empirical Formula from Combustion Analysis	Example 3.21 Determining an Empirical Formula from Combustion Analysis
	Upon combustion, a compound containing only carbon and hydrogen produces 1.83 g CO_2 and 0.901 g H_2O. Find the empirical formula of the compound.	Upon combustion, a 0.8233-g sample of a compound containing only carbon, hydrogen, and oxygen produces 2.445 g CO_2 and 0.6003 g H_2O. Find the empirical formula of the compound.
1. Write down as *given* the masses of each combustion product and the mass of the sample (if given).	GIVEN: 1.83 g CO_2, 0.901 g H_2O FIND: empirical formula	GIVEN: 0.8233-g sample, 2.445 g CO_2, 0.6003 g H_2O FIND: empirical formula
2. Convert the masses of CO_2 and H_2O from step 1 to moles by using the appropriate molar mass for each compound as a conversion factor.	$1.83\ \text{g CO}_2 \times \frac{1\ \text{mol CO}_2}{44.01\ \text{g CO}_2} = 0.0416\ \text{mol CO}_2$ $0.901\ \text{g H}_2\text{O} \times \frac{1\ \text{mol H}_2\text{O}}{18.02\ \text{g H}_2\text{O}} = 0.0500\ \text{mol H}_2\text{O}$	$2.445\ \text{g CO}_2 \times \frac{1\ \text{mol CO}_2}{44.01\ \text{g CO}_2} = 0.05556\ \text{mol CO}_2$ $0.6003\ \text{g H}_2\text{O} \times \frac{1\ \text{mol H}_2\text{O}}{18.02\ \text{g H}_2\text{O}} = 0.03331\ \text{mol H}_2\text{O}$
3. Convert the moles of CO_2 and moles of H_2O from step 2 to moles of C and moles of H using the conversion factors inherent in the chemical formulas of CO_2 and H_2O.	$0.0416\ \text{mol CO}_2 \times \frac{1\ \text{mol C}}{1\ \text{mol CO}_2} = 0.0416\ \text{mol C}$ $0.0500\ \text{mol H}_2\text{O} \times \frac{2\ \text{mol H}}{1\ \text{mol H}_2\text{O}} = 0.100\ \text{mol H}$	$0.05556\ \text{mol CO}_2 \times \frac{1\ \text{mol C}}{1\ \text{mol CO}_2} = 0.05556\ \text{mol C}$ $0.03331\ \text{mol H}_2\text{O} \times \frac{2\ \text{mol H}}{1\ \text{mol H}_2\text{O}} = 0.06662\ \text{mol H}$
4. If the compound contains an element other than C and H, find the mass of the other element by subtracting the sum of the masses of C and H from the mass of the sample. Finally, convert the mass of the other element to moles.	The sample contains no elements other than C and H, so proceed to the next step.	$\text{Mass C} = 0.05556\ \text{mol C} \times \frac{12.01\ \text{g C}}{\text{mol C}} = 0.6673\ \text{g C}$ $\text{Mass H} = 0.06662\ \text{mol H} \times \frac{1.008\ \text{g H}}{\text{mol H}} = 0.06715\ \text{g H}$ $\text{Mass O} = 0.8233\ \text{g} - (0.6673\ \text{g} + 0.06715\ \text{g}) = 0.0889\ \text{g}$ $\text{Mol O} = 0.0889\ \text{g O} \times \frac{\text{mol O}}{16.00\ \text{g O}} = 0.00556\ \text{mol O}$
5. Write down a pseudoformula for the compound using the number of moles of each element (from steps 3 and 4) as subscripts.	$C_{0.0416}H_{0.100}$	$C_{0.05556}H_{0.06662}O_{0.00556}$
6. Divide all the subscripts in the formula by the smallest subscript. (Round all subscripts that are within 0.1 of a whole number.)	$C_{\frac{0.0416}{0.0416}}H_{\frac{0.100}{0.0416}} \longrightarrow C_1H_{2.4}$	$C_{\frac{0.05556}{0.00556}}H_{\frac{0.0662}{0.00556}}O_{\frac{0.00556}{0.00556}} \longrightarrow C_{10}H_{12}O_1$

7. If the subscripts are not whole numbers, multiply all the subscripts by a small whole number to get whole-number subscripts.	$C_1H_{2.4} \times 5 \longrightarrow C_5H_{12}$ The correct empirical formula is C_5H_{12}.	The subscripts are whole numbers; no additional multiplication is needed. The correct empirical formula is $C_{10}H_{12}O$.

FOR PRACTICE 3.20

Upon combustion, a compound containing only carbon and hydrogen produces 1.60 g CO_2 and 0.819 g H_2O. Find the empirical formula of the compound.

FOR PRACTICE 3.21

Upon combustion, a 0.8009-g sample of a compound containing only carbon, hydrogen, and oxygen produces 1.6004 g CO_2 and 0.6551 g H_2O. Find the empirical formula of the compound.

3.11 Writing and Balancing Chemical Equations

Writing and Balancing Chemical Equations

Combustion analysis (which we just examined in Section 3.10) employs a **chemical reaction**, a process in which one or more substances are converted into one or more different ones. Compounds form and change through chemical reactions. For example, water is formed by the reaction of hydrogen with oxygen. A **combustion reaction** is a particular type of chemical reaction in which a substance combines with oxygen to form one or more oxygen-containing compounds. Combustion reactions also emit heat. The heat produced in a number of combustion reactions is critical to supplying our society's energy needs. For example, the heat from the combustion of gasoline expands the gaseous combustion products in a car engine's cylinders, which push the pistons and propel the car. We use the heat released by the combustion of *natural gas* to cook food and to heat our homes.

We represent a chemical reaction with a **chemical equation**. For example, we represent the combustion of natural gas with the equation:

$$\underset{\text{reactants}}{CH_4 + O_2} \longrightarrow \underset{\text{products}}{CO_2 + H_2O}$$

The substances on the left side of the equation are the **reactants**, and the substances on the right side are the **products**. We often specify the states of each reactant or product in parentheses next to the formula as follows:

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

The (g) indicates that these substances are gases in the reaction. Table 3.5 summarizes the common states of reactants and products and their symbols used in chemical equations.

TABLE 3.5 States of Reactants and Products in Chemical Equations

Abbreviation	State
(g)	Gas
(l)	Liquid
(s)	Solid
(aq)	Aqueous (water solution)

The equation just presented for the combustion of natural gas is not complete, however. If we look closely, we can immediately spot a problem.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

$O_2(g)$: 2 O atoms; $CO_2(g)$ and $H_2O(g)$: 2 O atoms + 1 O atom = 3 O atoms

The left side of the equation has two oxygen atoms, while the right side has three. The reaction as written, therefore, violates the law of conservation of mass because an oxygen atom formed out of nothing. Notice also that the left side has four hydrogen atoms, while the right side has only two.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

$CH_4(g)$: 4 H atoms; $H_2O(g)$: 2 H atoms

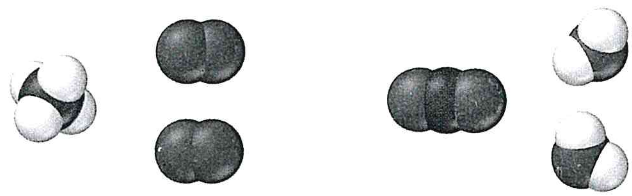
Two hydrogen atoms have vanished, again violating mass conservation. To correct these problems—that is, to write an equation that more closely represents *what actually happens*—we must **balance** the equation. We need to change the coefficients (the

We cannot change the subscripts when balancing a chemical equation because changing the subscripts changes the substance itself, while changing the coefficients changes the number of molecules of the substance. For example, 2 H_2O is simply two water molecules, but H_2O_2 is hydrogen peroxide, a drastically different compound.

numbers *in front of* the chemical formulas), not the subscripts (the numbers within t chemical formulas), to ensure that the number of each type of atom on the left side the equation is equal to the number on the right side. New atoms do not form durin reaction, nor do atoms vanish—matter must be conserved.

When we add coefficients to the reactants and products to balance an equation, change the number of molecules in the equation but not the *kind of* molecules. To b ance the equation for the combustion of methane, we put the coefficient 2 before O_2 the reactants, and the coefficient 2 before H_2O in the products.

$$CH_4(g) \quad + \quad 2\,O_2(g) \longrightarrow CO_2(g) \quad + \quad 2\,H_2O(g)$$



The equation is now balanced because the numbers of each type of atom on eit side of the equation are equal. The balanced equation tells us that one CH_4 molec reacts with two O_2 molecules to form one CO_2 molecule and two H_2O molecules. verify that the equation is balanced by summing the number of each type of atom each side of the equation.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Reactants	Products
1 C atom (1 × CH_4)	1 C atom (1 × CO_2)
4 H atoms (1 × CH_4)	4 H atoms (2 × H_2O)
4 O atoms (2 × O_2)	4 O atoms (1 × CO_2 + 2 × H_2O)

The number of each type of atom on both sides of the equation is now equal—the eq tion is balanced.

We can balance many chemical equations simply by trial and error. However, so guidelines are useful. For example, balancing the atoms in the most complex substan first and the atoms in the simplest substances (such as pure elements) last often ma the process shorter. The following examples of how to balance chemical equations presented in a two- or three-column format. The general guidelines are shown on left, with the application of the guidelines on the right. This procedure is meant only a flexible guide, not a rigid set of steps.

Procedure for...

Balancing Chemical Equations	**Example 3.22** Balancing Chemical Equations	**Example 3.23** Balancing Chemical Equations
	Write a balanced equation for the reaction between solid cobalt(III) oxide and solid carbon to produce solid cobalt and carbon dioxide gas.	Write a balanced equation for the combustion of gaseous butane (C_4H_{10}), a fuel used in portable stove and grills, in which it combines wit gaseous oxygen to form gaseous car-bon dioxide and gaseous water.
1. Write a skeletal equation by writing chemical formulas for each of the reactants and products. Review Sections 3.5 and 3.6 for nomenclature rules. (If a skeletal equation is provided, go to step 2.)	$Co_2O_3(s) + C(s) \longrightarrow Co(s) + CO_2(g)$	$C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O$

2. Balance atoms that occur in more complex substances first. Always balance atoms in compounds before atoms in pure elements.	**Begin with O:** $Co_2O_3(s) + C(s) \longrightarrow Co(s) + CO_2(g)$ 3 O atoms ⟶ 2 O atoms To balance O, put a 2 before $Co_2O_3(s)$ and a 3 before $CO_2(g)$. $2\ Co_2O_3(s) + C(s) \longrightarrow Co(s) + 3\ CO_2(g)$ 6 O atoms ⟶ 6 O atoms	**Begin with C:** $C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$ 4 C atoms ⟶ 1 C atom To balance C, put a 4 before $CO_2(g)$. $C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + H_2O(g)$ 4 C atoms ⟶ 4 C atoms **Balance H:** $C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + H_2O(g)$ 10 H atoms ⟶ 2 H atoms To balance H, put a 5 before $H_2O(g)$: $C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$ 10 H atoms ⟶ 10 H atoms
3. Balance atoms that occur as free elements on either side of the equation last. Balance free elements by adjusting their coefficients.	**Balance Co:** $2\ Co_2O_3(s) + C(s) \longrightarrow Co(s) + 3\ CO_2(g)$ 4 Co atoms ⟶ 1 Co atom To balance Co, put a 4 before $Co(s)$. $2\ Co_2O_3(s) + C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$ 4 Co atoms ⟶ 4 Co atoms **Balance C:** $2\ Co_2O_3(s) + C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$ 1 C atom ⟶ 3 C atoms To balance C, put a 3 before $C(s)$. $2\ Co_2O_3(s) + 3\ C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$	**Balance O:** $C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$ 2 O atoms ⟶ 8 O + 5 O = 13 O atom To balance O, put a 13/2 before $O_2(g)$. $C_4H_{10}(g) + 13/2\ O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$ 13 O atoms ⟶ 13 O atoms
4. If the balanced equation contains coefficient fractions, clear these by multiplying each of the coefficients in the entire equation by the denominator of the fraction.	This step is not necessary in this example. Proceed to step 5.	$[C_4H_{10}(g) + 13/2\ O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)] \times 2$ $2\ C_4H_{10}(g) + 13\ O_2(g) \longrightarrow 8\ CO_2(g) + 10\ H_2O(g)$
5. Check to make certain the equation is balanced by summing the total number of each type of atom on both sides of the equation.	$2\ Co_2O_3(s) + 3\ C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$ Left / Right: 4 Co atoms / 4 Co atoms 6 O atoms / 6 O atoms 3 C atoms / 3 C atoms The equation is balanced.	$2\ C_4H_{10}(g) + 13\ O_2(g) \longrightarrow 8\ CO_2(g) + 10\ H_2O(g)$ Left / Right: 8 C atoms / 8 C atoms 20 H atoms / 20 H atoms 26 O atoms / 26 O atoms The equation is balanced.
	FOR PRACTICE 3.22 Write a balanced equation for the reaction between solid silicon dioxide and solid carbon to produce solid silicon carbide and carbon monoxide gas.	FOR PRACTICE 3.23 Write a balanced equation for the combustion of gaseous ethane (C_2H_6), a minority component of natural gas, in which it combines with gaseous oxygen to form gaseous carbon dioxide and gaseous water.

CONCEPTUAL CONNECTION 3.10

Balanced Chemical Equations Which quantity or quantities must always be the same on both sides of a chemical equation?

(a) the number of atoms of each kind
(b) the number of molecules of each kind
(c) the number of moles of each kind of molecule
(d) the sum of masses of all substances involved

Example 3.24 Balancing Chemical Equations Containing Ionic Compounds with Polyatomic Ions

Write a balanced equation for the reaction between aqueous strontium chloride and aqueous lithium phosphate to form solid strontium phosphate and aqueous lithium chloride.

SOLUTION

1. Write a skeletal equation by writing chemical formulas for each of the reactants and products. Review Sections 3.5 and 3.6 for naming rules. (If a skeletal equation is provided, go to step 2.)	$SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$
2. Balance metal ions (cations) first. If a polyatomic cation exists on both sides of the equation, balance it as a unit.	**Begin with Sr^{2+}:** $SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ 1 Sr^{2+} ion $\longrightarrow$ 3 Sr^{2+} ions To balance Sr^{2+}, put a 3 before $SrCl_2(aq)$. $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ 3 Sr^{2+} ions $\longrightarrow$ 3 Sr^{2+} ions **Balance Li^+:** $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ 3 Li^+ ions $\longrightarrow$ 1 Li^+ ion To balance Li^+, put a 3 before $LiCl(aq)$. $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ 3 Li^+ ions $\longrightarrow$ 3 Li^+ ions
3. Balance nonmetal ions (anions) second. If a polyatomic anion exists on both sides of the equation, balance it as a unit.	**Balance PO_4^{3-}:** $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ 1 PO_4^{3-} ions $\longrightarrow$ 2 PO_4^{3-} ions To balance PO_4^{3-}, put a 2 before $Li_3PO_4(aq)$. $3\ SrCl_2(aq) + 2\ Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ 2 PO_4^{3-} ions $\longrightarrow$ 2 PO_4^{3-} ions **Balance Cl^-:** $3\ SrCl_2(aq) + 2\ Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ 6 Cl^- ions $\longrightarrow$ 3 Cl^- ion To balance Cl^-, replace the 3 before $LiCl(aq)$ with a 6. This also corrects the balance for Li^+, which was thrown off in the previous step. $3\ SrCl_2(aq) + 2\ Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 6\ LiCl(aq)$ 6 Cl^- ions $\longrightarrow$ 6 Cl^- ions

Check to make certain the equation is balanced by summing the total number of each type of ion on both sides of the equation.

$$3\,SrCl_2(aq) + 2\,Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 6\,LiCl(aq)$$

Left	Right
3 Sr^{2+} ions	3 Sr^{2+} ions
6 Li^{+} ions	6 Li^{+} ions
2 PO_4^{3-} ions	2 PO_4^{3-} ions
6 Cl^{-} ions	6 Cl^{-} ions

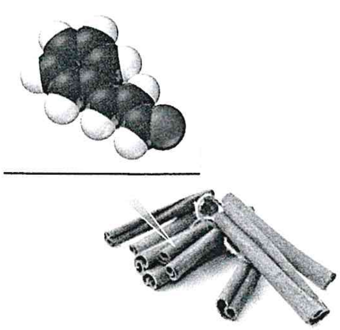
The equation is balanced.

FOR PRACTICE 3.24 Write a balanced equation for the reaction between aqueous lead(II) nitrate and aqueous potassium chloride to form solid lead(II) chloride and aqueous potassium nitrate.

3.12 Organic Compounds

Early chemists divided compounds into two types: organic and inorganic. They designated organic compounds as those that originate from living things. Sugar—from sugar-cane or the sugar beet—is a common example of an organic compound. Inorganic compounds, on the other hand, originate from the earth. Salt—mined from the ground or from the ocean—is a common example of an inorganic compound.

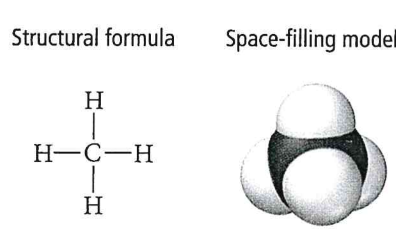
Not only did early chemists view organic and inorganic compounds as different in their origin, but also they recognized organic and inorganic compounds to be different in their properties. Organic compounds are easily decomposed. Inorganic compounds, however, are typically more difficult to decompose. Eighteenth-century chemists could synthesize inorganic compounds in the laboratory, but they could not synthesize organic compounds. This was considered to be another great difference between the two different types of compounds. Today, chemists can synthesize both organic and inorganic compounds, and even though organic chemistry is a subfield of chemistry, the differences between organic and inorganic compounds are now viewed as primarily organizational (not fundamental).



▲ The organic compound cinnamaldehyde is largely responsible for the taste and smell of cinnamon.

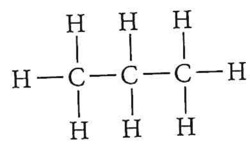
Organic compounds are common in everyday substances. Many smells—such as those in perfumes, spices, and foods—are caused by organic compounds. When you sprinkle cinnamon onto your French toast, some cinnamaldehyde—an organic compound present in cinnamon—evaporates into the air. As you inhale cinnamaldehyde molecules, you experience the unique smell of cinnamon. Organic compounds are the major components of living organisms. They are also the main components of most fuels, such as gasoline, oil, and natural gas, and they are the active ingredients in most pharmaceuticals, such as aspirin and ibuprofen.

Organic compounds are composed of carbon and hydrogen and a few other elements, including nitrogen, oxygen, and sulfur. The key element in organic chemistry, however, is carbon. In its compounds, carbon always forms four bonds. The simplest organic compound is methane, or CH_4.

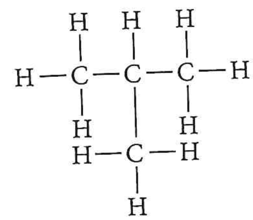


Methane, CH_4

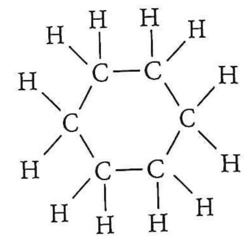
The chemistry of carbon is unique and complex because carbon frequently bonds to itself to form chain, branched, and ring structures.



Propane (C_3H_8)

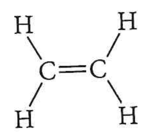


Isobutane (C_4H_{10})

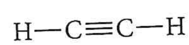


Cyclohexane (C_6H_{12})

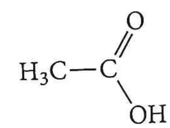
Carbon can also form double bonds and triple bonds with itself and with other elements.



Ethene (C_2H_4)



Ethyne (C_2H_2)

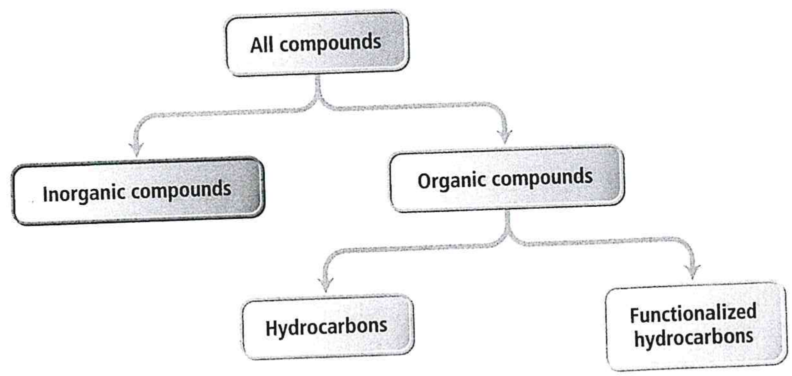


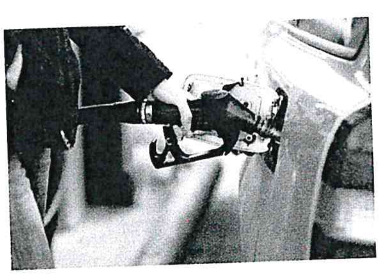
Acetic acid (CH_3COOH)

This versatility allows carbon to serve as the backbone of millions of different chemical compounds, which is why a general survey of organic chemistry is a yearlong course.

Hydrocarbons

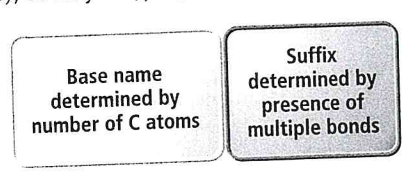
We can begin to scratch the surface of organic chemistry by categorizing organic compounds into types: hydrocarbons and functionalized hydrocarbons.





▲ Gasoline is composed mostly of hydrocarbons.

Hydrocarbons are organic compounds that contain only carbon and hydrogen. Hydrocarbons compose common fuels such as oil, gasoline, liquid propane gas, and natural gas. Hydrocarbons containing only single bonds are **alkanes**, while those containing double or triple bonds are **alkenes** and **alkynes**, respectively. The names of simple, straight-chain hydrocarbons consist of a base name, which is determined by the number of carbon atoms in the chain, and a suffix, determined by whether the hydrocarbon is an alkane (*-ane*), alkene (*-ene*), or alkyne (*-yne*).

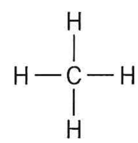
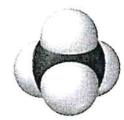
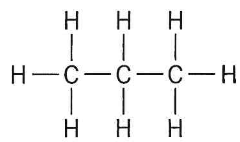
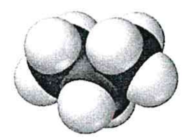
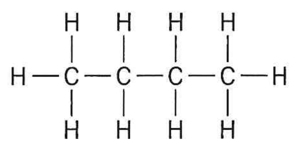
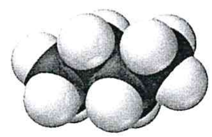
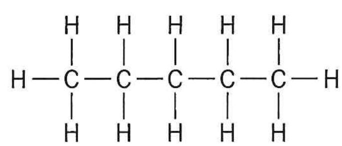
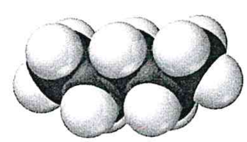
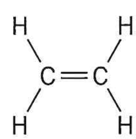
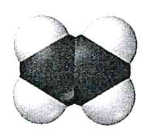


The base names for a number of hydrocarbons are listed here:

meth = 1	hex = 6
eth = 2	hept = 7
prop = 3	oct = 8
but = 4	non = 9
pent = 5	dec = 10

Table 3.6 lists some common hydrocarbons, their names, and their uses.

TABLE 3.6 Common Hydrocarbons

Name	Molecular Formula	Structural Formula	Space-filling Model	Common Uses
Methane	CH_4			Primary component of natural gas
Propane	C_3H_8			LP gas for grills and outdoor stoves
n-Butane*	C_4H_{10}			Common fuel for lighters
n-Pentane*	C_5H_{12}			Component of gasoline
Ethene	C_2H_4			Ripening agent in fruit
Ethyne	C_2H_2	$H-C\equiv C-H$		Fuel for welding torches

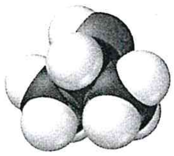
*The "*n*" in the names of these hydrocarbons stands for "normal," which means straight chain.

Functionalized Hydrocarbons

Functionalized hydrocarbons are hydrocarbons in which a **functional group**—a characteristic atom or group of atoms—is incorporated into the hydrocarbon. For example, **alcohols** are organic compounds that have an —OH functional group. We designate the hydrocarbon portion of a molecule as "R," and we write the general formula for an alcohol

The term *functional group* derives from the functionality or chemical character that a specific atom or group of atoms imparts to an organic compound. Even a carbon–carbon double bond can justifiably be called a "functional group."

as R—OH. Some examples of alcohols include methanol (also known as methyl alcohol or wood alcohol) and isopropanol (also known as isopropyl alcohol or rubbing alcohol).



Hydrocarbon (R) group → CH_3OH ← OH functional group

Methanol

Rubbing alcohol is isopropyl alcohol.

Hydrocarbon (R) group → $CH_3\underset{}{\overset{CH_3}{\overset{|}{C}}}$HOH ← OH functional group

Isopropanol
(2-propanol)

A group of organic compounds with the same functional group forms a **family**. Methanol and isopropyl alcohol are both members of the alcohol family of compounds.

The addition of a functional group to a hydrocarbon usually alters the properties of the compound significantly. Take *methanol*, which can be thought of as methane with an —OH group substituted for one of the hydrogen atoms. Methanol is a liquid at room temperature, whereas *methane* is a gas. Although each member of a family is unique, the common functional group bestows some chemical similarities on members of the same family. The names of functional groups have suffixes or endings unique to that functional group. Alcohols, for example, always have names that end in *-ol*. Table 3.7 provides examples of some common functional groups, their general formulas, and their characteristic suffixes or endings.

TABLE 3.7 Families of Organic Compounds

Family	Name Ending	General Formula	Example	Name	Occurrence/Use
Alcohols	-ol	R—OH	CH_3CH_2—OH	Ethanol (ethyl alcohol)	Alcohol in fermented beverages
Ethers	ether	R—O—R′	CH_3CH_2—O—CH_2CH_3	Diethyl ether	Anesthetic; laboratory solvent
Aldehydes	-al	$R-\overset{O}{\overset{\|}{C}}-H$	$H_3C-\overset{O}{\overset{\|}{C}}-H$	Ethanal (acetaldehyde)	Perfumes; flavors
Ketones	-one	$R-\overset{O}{\overset{\|}{C}}-R'$	$H_3C-\overset{O}{\overset{\|}{C}}-CH_3$	Propanone (acetone)	Fingernail polish remover
Carboxylic acids	acid	$R-\overset{O}{\overset{\|}{C}}-OH$	$H_3C-\overset{O}{\overset{\|}{C}}-OH$	Acetic acid	Vinegar
Esters	-ate	$R-\overset{O}{\overset{\|}{C}}-OR'$	$H_3C-\overset{O}{\overset{\|}{C}}-OCH_3$	Methyl acetate	Laboratory solvent
Amines	amine	RNH_2	$CH_3CH_2-\overset{H}{\overset{|}{N}}-H$	Ethyl amine	Smell of rotten fish

Self-Assessment Quiz

Q1. What is the empirical formula of a compound with the molecular formula $C_{10}H_8$?

a) C_5H_3 b) C_2H_4 c) C_5H_4 d) CH

Q2. Which substance is an ionic compound?

a) SrI_2 b) N_2O_4 c) He d) CCl_4

Q3. What is the correct formula for the compound formed between calcium and sulfur?

a) CaS b) Ca_2S c) CaS_2 d) CaS_3

Q4. Name the compound SrI_2.

a) strontium iodide b) strontium diiodide
c) strontium(II) iodide d) strontium(II) diiodide

Q5. What is the formula for manganese(IV) oxide?

a) Mn_4O b) MnO_4 c) Mn_2O d) MnO_2

Q6. Name the compound $Pb(C_2H_3O_2)_2$.

a) lead(II) carbonate b) lead(II) acetate
c) lead bicarbonate d) lead diacetate

Q7. Name the compound P_2I_4.

a) phosphorus iodide b) phosphorus diiodide
c) phosphorus(II) iodide d) diphosphorus tetraiodide

Q8. Name the compound $HNO_2(aq)$.

a) hydrogen nitrogen dioxide
b) hydrogen nitrate
c) nitric acid
d) nitrous acid

Q9. Determine the number of CH_2Cl_2 molecules in 25.0 g CH_2Cl_2.

a) 0.294 molecules b) 1.77×10^{23} molecules
c) 1.28×10^{27} molecules d) 1.51×10^{25} molecules

Q10. List the elements in the compound CF_2Cl_2 in order of decreasing mass percent composition.

a) C > F > Cl b) F > Cl > C
c) Cl > C > F d) Cl > F > C

Q11. Determine the mass of potassium in 35.5 g of KBr.

a) 17.4 g b) 0.298 g c) 11.7 g d) 32.9 g

Q12. A compound is 52.14% C, 13.13% H, and 34.73% O by mass. What is the empirical formula of the compound?

a) $C_2H_8O_3$ b) C_2H_6O c) C_4HO_3 d) C_3HO_6

Q13. A compound has the empirical formula CH_2O and a formula mass of 120.10 amu. What is the molecular formula of the compound?

a) CH_2O b) $C_2H_4O_2$ c) $C_3H_6O_3$ d) $C_4H_8O_4$

Q14. Combustion of 30.42 g of a compound containing only carbon, hydrogen, and oxygen produces 35.21 g CO_2 and 14.42 g H_2O. What is the empirical formula of the compound?

a) $C_4H_8O_6$ b) $C_2H_4O_3$
c) $C_2H_2O_3$ d) C_6HO_{12}

Q15. What are the correct coefficients (reading from left to right) when the chemical equation is balanced?

$$\_PCl_3(l) + \_H_2O(l) \longrightarrow \_H_3PO_3(aq) + \_HCl(aq)$$

a) 1, 3, 1, 3 b) 1, 2, 1, 1
c) 1, 3, 2, 1 d) 3, 6, 1, 9

Answers: **1.** (c) **2.** (a) **3.** (a) **4.** (a) **5.** (d) **6.** (b) **7.** (d) **8.** (d) **9.** (b) **10.** (d) **11.** (c) **12.** (b) **13.** (d) **14.** (b) **15.** (a)

Chapter 3 IN REVIEW

Key Terms

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ionic compound (90)
covalent bond (90)
molecular compound (90)

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empirical formula (90)
molecular formula (90)
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space-filling molecular model (92)

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formula unit (94)
polyatomic ion (95)

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common name (97)
systematic name (97)
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oxyanion (100)
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alkyne (124)
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alcohol (125)
family (126)

Key Concepts

Chemical Bonds (3.2)

- Chemical bonds, the forces that hold atoms together in compounds, arise from the interactions between nuclei and electrons in atoms.
- In an ionic bond, one or more electrons are *transferred* from one atom to another, forming a cation (positively charged) and an anion (negatively charged). The two ions are drawn together by the attraction between the opposite charges.
- In a covalent bond, one or more electrons are *shared* between two atoms. The atoms are held together by the attraction between their nuclei and the shared electrons.

Representing Molecules and Compounds (3.3, 3.4)

- A compound is represented with a chemical formula, which indicates the elements present and the number of atoms of each.
- An empirical formula gives only the *relative* number of atoms, while a molecular formula gives the *actual* number of atoms present in the molecule.
- Structural formulas show how atoms are bonded together, while molecular models portray the geometry of the molecule.
- Compounds can be divided into two types: molecular compounds, formed between two or more covalently bonded nonmetals, and ionic compounds, usually formed between a metal ionically bonded to one or more nonmetals. The smallest identifiable unit of a molecular compound is a molecule, and the smallest identifiable unit of an ionic compound is a formula unit: the smallest electrically neutral collection of ions.
- Elements can also be divided into two types: molecular elements, which occur as (mostly diatomic) molecules, and atomic elements, which occur as individual atoms.

Naming Inorganic Ionic and Molecular Compounds and Acids (3.5–3.7)

- A flowchart for naming simple inorganic compounds is provided in Section 3.7.

Formula Mass and Mole Concept for Compounds (3.8)

- The formula mass of a compound is the sum of the atomic masses of all the atoms in the chemical formula. Like the atomic masses of elements, the formula mass characterizes the average mass of a molecule (or a formula unit).
- The mass of one mole of a compound is the molar mass of i compound and equals its formula mass (in grams).

Chemical Composition (3.9, 3.10)

- The mass percent composition of a compound indicates each ment's percentage of the total compound's mass. We can de mine the mass percent composition from the compound's che cal formula and the molar masses of its elements.
- The chemical formula of a compound provides the relative number of atoms (or moles) of each element in a compoun and we can therefore use it to determine numerical relation ships between moles of the compound, and moles of its constituent elements. We can extend this relationship to m by using the molar masses of the compound and its constitu ent elements.
- If the mass percent composition and molar mass of a c pound are known, we can determine its empirical and mol lar formulas.

Writing and Balancing Chemical Equations (3.11)

- In chemistry, we represent chemical reactions with chem equations. The substances on the left-hand side of a chem equation are the reactants, and the substances on the right-h side are the products.
- Chemical equations are balanced when the number of each t of atom on the left side of the equation is equal to the numbe the right side.

Organic Compounds (3.12)

- Organic compounds are composed of carbon, hydrogen, and a other elements such as nitrogen, oxygen, and sulfur.
- The simplest organic compounds are hydrocarbons, compoun composed of only carbon and hydrogen.
- Hydrocarbons are categorized into three types based on the bonds they contain: alkanes contain single bonds, alkenes co tain double bonds, and alkynes contain triple bonds.
- All other organic compounds can be thought of as hydrocarbc with one or more functional groups—characteristic atoms or groups of atoms.
- Common functionalized hydrocarbons include alcohols, eth aldehydes, ketones, carboxylic acids, esters, and amines.

Key Equations and Relationships

Formula Mass (3.8)

$$\left(\begin{matrix}\text{No. of atoms of 1st element}\\\text{in chemical formula}\end{matrix} \times \begin{matrix}\text{atomic mass}\\\text{of 1st element}\end{matrix}\right) + \left(\begin{matrix}\text{No. of atoms of 2nd element}\\\text{in chemical formula}\end{matrix} \times \begin{matrix}\text{atomic mass}\\\text{of 2nd element}\end{matrix}\right) + \ldots$$

Mass Percent Composition (3.9)

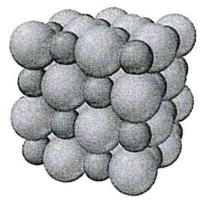
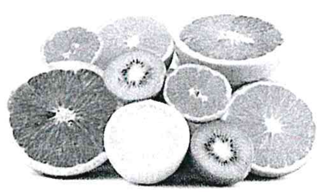
$$\text{Mass \% of element X} = \frac{\text{mass of X in 1 mol compound}}{\text{mass of 1 mol compound}} \times 100\%$$

Empirical Formula Molar Mass (3.10)

$$\text{Molecular formula} = n \times (\text{empirical formula})$$

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

Key Learning Outcomes

CHAPTER OBJECTIVES	ASSESSMENT
Writing Molecular and Empirical Formulas (3.3)	Example 3.1 For Practice 3.1 Exercises 4, 23–26
Classifying Substances as Atomic Elements, Molecular Elements, Molecular Compounds, or Ionic Compounds (3.4) 	Example 3.2 For Practice 3.2 Exercises 27–32
Writing Formulas for Ionic Compounds (3.5)	Examples 3.3, 3.4 For Practice 3.3, 3.4 Exercises 33–36, 43, 44
Naming Ionic Compounds (3.5) 	Examples 3.5, 3.6 For Practice 3.5, 3.6 For More Practice 3.5, 3.6 Exercises 37–40
Naming Ionic Compounds Containing Polyatomic Ions (3.5)	Example 3.7 For Practice 3.7 For More Practice 3.7 Exercises 41–44
Naming Molecular Compounds (3.6)	Example 3.8 For Practice 3.8 For More Practice 3.8 Exercises 47–50
Naming Acids (3.6) 	Examples 3.9, 3.10 For Practice 3.9, 3.10 For More Practice 3.10 Exercises 51–54
Naming Uncategorized Inorganic Compounds (3.7)	Example 3.11 For Practice 3.11 Exercises 55–58
Calculating Formula Mass (3.8)	Example 3.12 For Practice 3.12 Exercises 59, 60
Using Formula Mass to Count Molecules by Weighing (3.8)	Example 3.13 For Practice 3.13 For More Practice 3.13 Exercises 65–70
Calculating Mass Percent Composition (3.9)	Example 3.14 For Practice 3.14 For More Practice 3.14 Exercises 71–76
Using Mass Percent Composition as a Conversion Factor (3.9)	Example 3.15 For Practice 3.15 For More Practice 3.15 Exercises 77, 78
Using Chemical Formulas as Conversion Factors (3.9)	Example 3.16 For Practice 3.16 For More Practice 3.16 Exercises 83, 84
Obtaining an Empirical Formula from Experimental Data (3.10)	Examples 3.17, 3.18 For Practice 3.17, 3.18 Exercises 87–94
Calculating a Molecular Formula from an Empirical Formula and Molar Mass (3.10)	Example 3.19 For Practice 3.19 For More Practice 3.19 Exercises 95–96
Obtaining an Empirical Formula from Combustion Analysis (3.10)	Examples 3.20, 3.21 For Practice 3.20, 3.21 Exercises 97–100
Balancing Chemical Equations (3.11)	Examples 3.22, 3.23, 3.24 For Practice 3.22, 3.23, 3.24 Exercises 101–112

EXERCISES

Review Questions

1. How do the properties of compounds compare to the properties of the elements from which the compounds are composed?
2. What is a chemical bond? Explain the difference between an ionic bond and a covalent bond.
3. Explain the different ways to represent compounds. Why are there so many?
4. What is the difference between an empirical formula and a molecular formula?
5. Define and provide an example for each of the following: atomic element, molecular element, ionic compound, molecular compound.

6. Explain how to write a formula for an ionic compound given the names of the metal and nonmetal (or polyatomic ion) in the compound.
7. Explain how to name binary ionic compounds. How do you name an ionic compound if it contains a polyatomic ion?
8. Why do the names of some ionic compounds include the charge of the metal ion while others do not?
9. Explain how to name molecular inorganic compounds.
10. How many atoms are specified by each of these prefixes: *mono-*, *di-*, *tri-*, *tetra-*, *penta-*, *hexa-*?
11. Explain how to name binary and oxyacids.
12. What is the formula mass for a compound? Why is it useful?
13. Explain how you can use the information in a chemical formula to determine how much of a particular element is present in a given amount of a compound. Provide some examples of why this might be important.
14. What is mass percent composition? Why is it useful?
15. What kinds of conversion factors are inherent in chemical formulas? Provide an example.
16. What kind of chemical formula can be obtained from experimental data showing the relative masses of the elements in a compound?
17. How can a molecular formula be obtained from an empirical formula? What additional information is required?
18. What is combustion analysis? What is it used for?
19. Which elements are normally present in organic compounds?
20. What is the difference between an alkane, an alkene, and an alkyne?
21. What are functionalized hydrocarbons? Cite an example of a functionalized hydrocarbon.
22. Write a generic formula for each of the families of organic compounds.
    a. alcohols
    b. ethers
    c. aldehydes
    d. ketones
    e. carboxylic acids
    f. esters
    g. amines

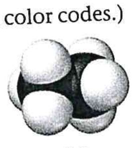
Problems by Topic

Note: Answers to all odd-numbered Problems, numbered in blue, can be found in Appendix III. Exercises in the Problems by Topic section are paired, with each odd-numbered problem followed by a similar even-numbered problem. Exercises in the Cumulative Problems section are also paired, but somewhat more loosely. (Challenge Problems and Conceptual Problems, because of their nature, are unpaired.)

Chemical Formulas and Molecular View of Elements and Compounds

23. Determine the number of each type of atom in each formula.
    a. $Mg_3(PO_4)_2$ b. $BaCl_2$ c. $Fe(NO_2)_2$ d. $Ca(OH)_2$
24. Determine the number of each type of atom in each formula.
    a. $Ca(NO_2)_2$ b. $CuSO_4$ c. $Al(NO_3)_3$ d. $Mg(HCO_3)_2$
25. Write a chemical formula for each molecular model. (See Appendix IIA for color codes.)

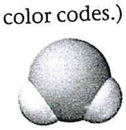

(a)


(b)


(c)

26. Write a chemical formula for each molecular model. (See Appendix IIA for color codes.)


(a)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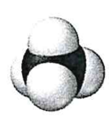
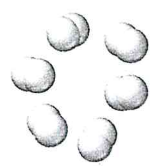
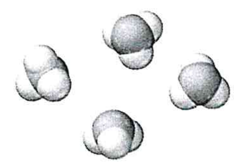

(c)

27. Classify each element as atomic or molecular.
    a. neon b. fluorine c. potassium d. nitrogen
28. Identify the elements that have molecules as their basic units.
    a. hydrogen b. iodine c. lead d. oxygen
29. Classify each compound as ionic or molecular.
    a. CO_2 b. $NiCl_2$ c. NaI d. PCl_3
30. Classify each compound as ionic or molecular.
    a. CF_2Cl_2 b. CCl_4 c. PtO_2 d. SO_3
31. Based on the molecular views, classify each substance as an atomic element, a molecular element, an ionic compound, or a molecular compound.


(a)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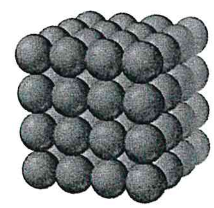
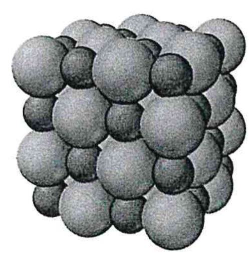
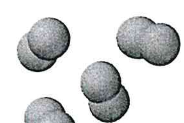

(c)

32. Based on the molecular views, classify each substance as an atomic element, a molecular element, an ionic compound, or a molecular compound.


(a)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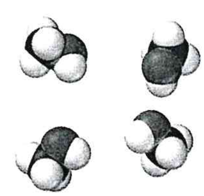

(c)

Formulas and Names for Ionic Compounds

33. Write a formula for the ionic compound that forms between each pair of elements.
a. calcium and oxygen b. zinc and sulfur
c. rubidium and bromine d. aluminum and oxygen

34. Write a formula for the ionic compound that forms between each pair of elements.
a. silver and chlorine b. sodium and sulfur
c. aluminum and sulfur d. potassium and chlorine

35. Write a formula for the compound that forms between calcium and each polyatomic ion.
a. hydroxide b. chromate
c. phosphate d. cyanide

36. Write a formula for the compound that forms between potassium and each polyatomic ion.
a. carbonate b. phosphate
c. hydrogen phosphate d. acetate

37. Name each ionic compound.
a. Mg_3N_2 b. KF c. Na_2O
d. Li_2S e. CsF f. KI

38. Name each ionic compound.
a. $SnCl_4$ b. PbI_2 c. Fe_2O_3
d. CuI_2 e. $HgBr_2$ f. $CrCl_2$

39. Give each ionic compound an appropriate name.
a. SnO b. Cr_2S_3 c. RbI d. $BaBr_2$

40. Give each ionic compound an appropriate name.
a. BaS b. $FeCl_3$ c. PbI_4 d. $SrBr_2$

41. Name each ionic compound containing a polyatomic ion.
a. $CuNO_2$ b. $Mg(C_2H_3O_2)_2$
c. $Ba(NO_3)_2$ d. $Pb(C_2H_3O_2)_2$

42. Name each ionic compound containing a polyatomic ion.
a. $Ba(OH)_2$ b. NH_4I c. $NaBrO_4$ d. $Fe(OH)_3$

43. Write the formula for each ionic compound.
a. sodium hydrogen sulfite
b. lithium permanganate
c. silver nitrate
d. potassium sulfate
e. rubidium hydrogen sulfate
f. potassium hydrogen carbonate

44. Write the formula for each ionic compound.
a. copper(II) chloride b. copper(I) iodate
c. lead(II) chromate d. calcium fluoride
e. potassium hydroxide f. iron(II) phosphate

45. Write the name from the formula or the formula from the name for each hydrated ionic compound.
a. $CoSO_4 \cdot 7\,H_2O$ b. iridium(III) bromide tetrahydrate
c. $Mg(BrO_3)_2 \cdot 6\,H_2O$ d. potassium carbonate dihydrate

46. Write the name from the formula or the formula from the name for each hydrated ionic compound.
a. cobalt(II) phosphate octahydrate b. $BeCl_2 \cdot 2\,H_2O$
c. chromium(III) phosphate trihydrate d. $LiNO_2 \cdot H_2O$

Formulas and Names for Molecular Compounds and Acids

47. Name each molecular compound.
a. CO b. NI_3 c. $SiCl_4$ d. N_4Se_4

48. Name each molecular compound.
a. SO_3 b. SO_2 c. BrF_5 d. NO

49. Write the formula for each molecular compound.
a. phosphorus trichloride b. chlorine monoxide
c. disulfur tetrafluoride d. phosphorus pentafluoride

50. Write the formula for each molecular compound.
a. boron tribromide b. dichlorine monoxide
c. xenon tetrafluoride d. carbon tetrabromide

51. Name each acid.
a. $HI(aq)$ b. $HNO_3(aq)$ c. $H_2CO_3(aq)$

52. Name each acid.
a. $HCl(aq)$ b. $HClO_2(aq)$ c. $H_2SO_4(aq)$

53. Write the formula for each acid.
a. hydrofluoric acid b. hydrobromic acid
c. sulfurous acid

54. Write the formula for each acid.
a. phosphoric acid b. hydrocyanic acid
c. chlorous acid

Using the Nomenclature Flowchart

55. Refer to the nomenclature flowchart (Figure 3.11) to name each compound.
a. $SrCl_2$ b. SnO_2 c. P_2S_5 d. $HC_2H_3O_2(aq)$

56. Refer to the nomenclature flowchart (Figure 3.11) to name each compound.
a. $HNO_2(aq)$ b. B_2Cl_2 c. $BaCl_2$ d. $CrCl_3$

57. Refer to the nomenclature flowchart (Figure 3.11) to name each compound.
a. $KClO_3$ b. I_2O_5 c. $PbSO_4$

58. Refer to the nomenclature flowchart (Figure 3.11) to name each compound.
a. XeO_3 b. KClO c. $CoSO_4$

Formula Mass and the Mole Concept for Compounds

59. Calculate the formula mass for each compound.
a. NO_2 b. C_4H_{10} c. $C_6H_{12}O_6$ d. $Cr(NO_3)_3$

60. Calculate the formula mass for each compound.
a. $MgBr_2$ b. HNO_2 c. CBr_4 d. $Ca(NO_3)_2$

61. Calculate the number of moles in each sample.
a. 72.5 g CCl_4 b. 12.4 g $C_{12}H_{22}O_{11}$
c. 25.2 kg C_2H_2 d. 12.3 g dinitrogen monoxide

62. Calculate the mass of each sample.
a. 15.7 mol HNO_3
b. 1.04×10^{-3} mol H_2O_2
c. 72.1 mmol SO_2
d. 1.23 mol xenon difluoride

63. Determine the number of moles (of molecules or formula units) in each sample.
a. 25.5 g NO_2 b. 1.25 kg CO_2
c. 38.2 g KNO_3 d. 155.2 kg Na_2SO_4

64. Determine the number of moles (of molecules or formula units) in each sample.
a. 55.98 g CF_2Cl_2 b. 23.6 kg $Fe(NO_3)_2$
c. 0.1187 g C_8H_{18} d. 195 kg CaO

65. How many molecules are in each sample?
a. 6.5 g H_2O b. 389 g CBr_4 c. 22.1 g O_2 d. 19.3 g C_8H_{10}

66. How many molecules (or formula units) are in each sample?
a. 85.26 g CCl_4 b. 55.93 kg $NaHCO_3$
c. 119.78 g C_4H_{10} d. 4.59×10^5 g Na_3PO_4

67. Calculate the mass (in g) of each sample.
 a. 5.94×10^{20} SO_3 molecules
 b. 2.8×10^{22} H_2O molecules
 c. 1 glucose molecule ($C_6H_{12}O_6$)

68. Calculate the mass (in g) of each sample.
 a. 4.5×10^{25} O_3 molecules
 b. 9.85×10^{19} CCl_2F_2 molecules
 c. 1 water molecule

69. A sugar crystal contains approximately 1.8×10^{17} sucrose ($C_{12}H_{22}O_{11}$) molecules. What is its mass in mg?

70. A salt crystal has a mass of 0.12 mg. How many NaCl formula units does it contain?

Composition of Compounds

71. Calculate the mass percent composition of carbon in each carbon-containing compound.
 a. CH_4 b. C_2H_6 c. C_2H_2 d. C_2H_5Cl

72. Calculate the mass percent composition of nitrogen in each nitrogen-containing compound.
 a. N_2O b. NO c. NO_2 d. HNO_3

73. Most fertilizers consist of nitrogen-containing compounds such as NH_3, $CO(NH_2)_2$, NH_4NO_3, and $(NH_4)_2SO_4$. Plants use the nitrogen content in these compounds for protein synthesis. Calculate the mass percent composition of nitrogen in each of the fertilizers listed. Which fertilizer has the highest nitrogen content?

74. Iron in the earth is in the form of iron ore. Common ores include Fe_2O_3 (hematite), Fe_3O_4 (magnetite), and $FeCO_3$ (siderite). Calculate the mass percent composition of iron for each of these iron ores. Which ore has the highest iron content?

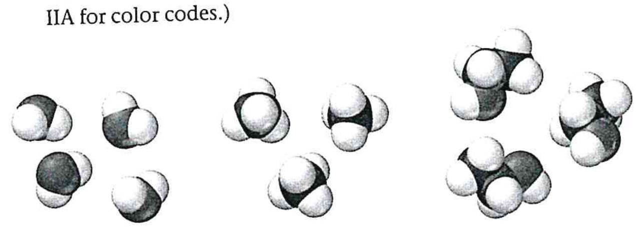
75. Copper(II) fluoride contains 37.42% F by mass. Calculate the mass of fluorine (in g) in 55.5 g of copper(II) fluoride.

76. Silver chloride, often used in silver plating, contains 75.27% Ag by mass. Calculate the mass of silver chloride required to plate 155 mg of pure silver.

77. The iodide ion is a dietary mineral essential to good nutrition. In countries where potassium iodide is added to salt, iodine deficiency (or goiter) has been almost completely eliminated. The recommended daily allowance (RDA) for iodine is 150 μg/day. How much potassium iodide (76.45% I) should you consume if you want to meet the RDA?

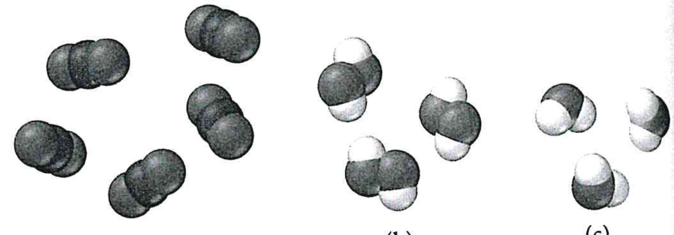
78. The American Dental Association recommends that an adult female should consume 3.0 mg of fluoride (F^-) per day to prevent tooth decay. If the fluoride is consumed in the form of sodium fluoride (45.24% F), what amount of sodium fluoride contains the recommended amount of fluoride?

79. Write a ratio showing the relationship between the molar amounts of each element for each compound. (See Appendix IIA for color codes.)



(a) (b) (c)

80. Write a ratio showing the relationship between the molar amounts of each element for each compound. (See Appendix IIA for color codes.)



(a) (b) (c)

81. Determine the number of moles of hydrogen atoms in each sample.
 a. 0.0885 mol C_4H_{10} b. 1.3 mol CH_4
 c. 2.4 mol C_6H_{12} d. 1.87 mol C_8H_{18}

82. Determine the number of moles of oxygen atoms in each sample.
 a. 4.88 mol H_2O_2 b. 2.15 mol N_2O
 c. 0.0237 mol H_2CO_3 d. 24.1 mol CO_2

83. Calculate mass (in grams) of sodium in 8.5 g of each sodium-containing food additive.
 a. NaCl (table salt)
 b. Na_3PO_4 (sodium phosphate)
 c. $NaC_7H_5O_2$ (sodium benzoate)
 d. $Na_2C_6H_6O_7$ (sodium hydrogen citrate)

84. Calculate the mass (in kilograms) of chlorine in 25 kg of each chlorofluorocarbon (CFC).
 a. CF_2Cl_2 b. $CFCl_3$ c. $C_2F_3Cl_3$ d. CF_3Cl

85. How many fluorine atoms are present in 5.85 g of C_2F_4?

86. How many bromine atoms are present in 35.2 g of CH_2Br_2?

Chemical Formulas from Experimental Data

87. A chemist decomposes samples of several compounds; the masses of their constituent elements are listed. Calculate the empirical formula for each compound.
 a. 1.651 g Ag, 0.1224 g O
 b. 0.672 g Co, 0.569 g As, 0.486 g O
 c. 1.443 g Se, 5.841 g Br

88. A chemist decomposes samples of several compounds; the masses of their constituent elements are listed. Calculate the empirical formula for each compound.
 a. 1.245 g Ni, 5.381 g I
 b. 2.677 g Ba, 3.115 g Br
 c. 2.128 g Be, 7.557 g S, 15.107 g O

89. Calculate the empirical formula for each stimulant based on its elemental mass percent composition.
 a. nicotine (found in tobacco leaves): C 74.03%, H 8.70%, N 17.27%
 b. caffeine (found in coffee beans): C 49.48%, H 5.19%, N 28.85%, O 16.48%

90. Calculate the empirical formula for each natural flavor based on its elemental mass percent composition.
 a. methyl butyrate (component of apple taste and smell): C 58.80%, H 9.87%, O 31.33%
 b. vanillin (responsible for the taste and smell of vanilla): C 63.15%, H 5.30%, O 31.55%

91. The elemental mass percent composition of ibuprofen (a nonsteroidal anti-inflammatory drug [NSAID]) is 75.69% C, 8.80% H, and 15.51% O. Determine the empirical formula of ibuprofen.

92. The elemental mass percent composition of ascorbic acid (vitamin C) is 40.92% C, 4.58% H, and 54.50% O. Determine the empirical formula of ascorbic acid.

93. A 0.77-mg sample of nitrogen reacts with chlorine to form 6.61 mg of the chloride. Determine the empirical formula of nitrogen chloride.

94. A 45.2-mg sample of phosphorus reacts with selenium to form 131.6 mg of the selenide. Determine the empirical formula of phosphorus selenide.

95. From the given empirical formula and molar mass, find the molecular formula of each compound.
a. C_6H_7N, 186.24 g/mol b. C_2HCl, 181.44 g/mol
c. $C_5H_{10}NS_2$, 296.54 g/mol

96. From the given molar mass and empirical formula of several compounds, find the molecular formula of each compound.
a. C_4H_9, 114.22 g/mol b. CCl, 284.77 g/mol
c. C_3H_2N, 312.29 g/mol

97. Combustion analysis of a hydrocarbon produces 33.01 g CO_2 and 13.51 g H_2O. Calculate the empirical formula of the hydrocarbon.

98. Combustion analysis of naphthalene, a hydrocarbon used in mothballs, produces 8.80 g CO_2 and 1.44 g H_2O. Calculate the empirical formula of naphthalene.

99. The foul odor of rancid butter is due largely to butyric acid, a compound containing carbon, hydrogen, and oxygen. Combustion analysis of a 4.30-g sample of butyric acid produces 8.59 g CO_2 and 3.52 g H_2O. Determine the empirical formula of butyric acid.

100. Tartaric acid is the white, powdery substance that coats tart candies such as Sour Patch Kids. Combustion analysis of a 12.01-g sample of tartaric acid—which contains only carbon, hydrogen, and oxygen—produces 14.08 g CO_2 and 4.32 g H_2O. Determine the empirical formula of tartaric acid.

Writing and Balancing Chemical Equations

101. Sulfuric acid is a component of acid rain formed when gaseous sulfur dioxide pollutant reacts with gaseous oxygen and liquid water to form aqueous sulfuric acid. Write the balanced chemical equation this reaction. (Note: this is a simplified representation of this reaction.)

102. Nitric acid is a component of acid rain that forms when gaseous nitrogen dioxide pollutant reacts with gaseous oxygen and liquid water to form aqueous nitric acid. Write the balanced chemical equation for this reaction. (Note: this is a simplified representation of this reaction.)

103. In a popular classroom demonstration, solid sodium is added to liquid water and reacts to produce hydrogen gas and aqueous sodium hydroxide. Write the balanced chemical equation for this reaction.

104. When iron rusts, solid iron reacts with gaseous oxygen to form solid iron(III) oxide. Write the balanced chemical equation for this reaction.

105. Write the balanced chemical equation for the fermentation of sucrose ($C_{12}H_{22}O_{11}$) by yeasts in which the aqueous sugar reacts with water to form aqueous ethanol (C_2H_5OH) and carbon dioxide gas.

106. Write the balanced equation for the photosynthesis reaction in which gaseous carbon dioxide and liquid water react in the presence of chlorophyll to produce aqueous glucose ($C_6H_{12}O_6$) and oxygen gas.

107. Write the balanced chemical equation for each reaction.
a. Solid lead(II) sulfide reacts with aqueous hydrobromic acid to form solid lead(II) bromide and dihydrogen monosulfide gas.
b. Gaseous carbon monoxide reacts with hydrogen gas to form gaseous methane (CH_4) and liquid water.
c. Aqueous hydrochloric acid reacts with solid manganese(IV) oxide to form aqueous manganese(II) chloride, liquid water, and chlorine gas.
d. Liquid pentane (C_5H_{12}) reacts with gaseous oxygen to form carbon dioxide and liquid water.

108. Write the balanced chemical equation for each reaction.
a. Solid copper reacts with solid sulfur to form solid copper(I) sulfide.
b. Solid iron(III) oxide reacts with hydrogen gas to form solid iron and liquid water.
c. Sulfur dioxide gas reacts with oxygen gas to form sulfur trioxide gas.
d. Gaseous ammonia (NH_3) reacts with gaseous oxygen to form gaseous nitrogen monoxide and gaseous water.

109. Write the balanced chemical equation for the reaction of aqueous sodium carbonate with aqueous copper(II) chloride to form solid copper(II) carbonate and aqueous sodium chloride.

110. Write the balanced chemical equation for the reaction of aqueous potassium hydroxide with aqueous iron(III) chloride to form solid iron(III) hydroxide and aqueous potassium chloride.

111. Balance each chemical equation.
a. $CO_2(g) + CaSiO_3(s) + H_2O(l) \longrightarrow SiO_2(s) + Ca(HCO_3)_2(aq)$
b. $Co(NO_3)_3(aq) + (NH_4)_2S(aq) \longrightarrow Co_2S_3(s) + NH_4NO_3(aq)$
c. $Cu_2O(s) + C(s) \longrightarrow Cu(s) + CO(g)$
d. $H_2(g) + Cl_2(g) \longrightarrow HCl(g)$

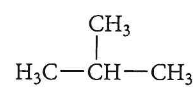
112. Balance each chemical equation.
a. $Na_2S(aq) + Cu(NO_3)_2(aq) \longrightarrow NaNO_3(aq) + CuS(s)$
b. $N_2H_4(l) \longrightarrow NH_3(g) + N_2(g)$
c. $HCl(aq) + O_2(g) \longrightarrow H_2O(l) + Cl_2(g)$
d. $FeS(s) + HCl(aq) \longrightarrow FeCl_2(aq) + H_2S(g)$

Organic Compounds

113. Classify each compound as organic or inorganic.
a. $CaCO_3$ b. C_4H_8 c. $C_4H_6O_6$ d. LiF

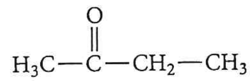
114. Classify each compound as organic or inorganic.
a. C_8H_{18} b. CH_3NH_2 c. CaO d. $FeCO_3$

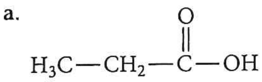
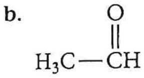
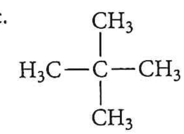
115. Classify each hydrocarbon as an alkane, alkene, or alkyne.
a. $H_2C{=}CH{-}CH_3$
b. $H_3C{-}CH_2{-}CH_3$
c. $HC{\equiv}C{-}CH_3$
d. $H_3C{-}CH_2{-}CH_2{-}CH_3$

116. Classify each hydrocarbon as an alkane, alkene, or alkyne.
a. $HC{\equiv}CH$ b. $H_3C{-}CH{=}CH{-}CH_3$
c. 
d. $H_3C{-}C{\equiv}C{-}CH_3$

117. Write the formula based on the name, or the name based on the formula, for each hydrocarbon.
a. propane b. $CH_3CH_2CH_3$
c. octane d. $CH_3CH_2CH_2CH_2CH_3$

118. Write the formula based on the name, or the name based on the formula, for each hydrocarbon.
a. CH_3CH_3 b. pentane
c. $CH_3CH_2CH_2CH_2CH_2CH_3$ d. heptane

119. Classify each organic compound as a hydrocarbon or a functionalized hydrocarbon. For functionalized hydrocarbons, identify the compound's family.
   a. $H_3C—CH_2OH$
   b. $H_3C—CH_3$
   c. 
   d. $H_3C—NH_2$

120. Classify each organic compound as a hydrocarbon or a functionalized hydrocarbon. For functionalized hydrocarbons, identify the compound's family.
   a. 
   b. 
   c. 
   d. $H_3C—CH_2—O—CH_3$

Cumulative Problems

121. How many molecules of ethanol (C_2H_5OH) (the alcohol in alcoholic beverages) are present in 145 mL of ethanol? The density of ethanol is 0.789 g/cm^3.

122. A drop of water has a volume of approximately 0.05 mL. How many water molecules does it contain? The density of water is 1.0 g/cm^3.

123. Determine the chemical formula of each compound and then use it to calculate the mass percent composition of each constituent element.
   a. potassium chromate
   b. lead(II) phosphate
   c. sulfurous acid
   d. cobalt(II) bromide

124. Determine the chemical formula of each compound and then use it to calculate the mass percent composition of each constituent element.
   a. perchloric acid
   b. phosphorus pentachloride
   c. nitrogen triiodide
   d. carbon dioxide

125. A Freon leak in the air-conditioning system of an old car releases 25 g of CF_2Cl_2 per month. What mass of chlorine does this car emit into the atmosphere each year?

126. A Freon leak in the air-conditioning system of a large building releases 12 kg of CHF_2Cl per month. If the leak is allowed to continue, how many kilograms of Cl will be emitted into the atmosphere each year?

127. A metal (M) forms a compound with the formula MCl_3. If the compound contains 65.57% Cl by mass, what is the identity of the metal?

128. A metal (M) forms an oxide with the formula M_2O. If the oxide contains 16.99% O by mass, what is the identity of the metal?

129. Estradiol is a female sexual hormone that is responsible for the maturation and maintenance of the female reproductive system. Elemental analysis of estradiol gives the following mass percent composition: C 79.37%, H 8.88%, O 11.75%. The molar mass of estradiol is 272.37 g/mol. Find the molecular formula of estradiol.

130. Fructose is a common sugar found in fruit. Elemental analysis of fructose gives the following mass percent composition: C 40.00%, H 6.72%, O 53.28%. The molar mass of fructose is 180.16 g/mol. Find the molecular formula of fructose.

131. Combustion analysis of a 13.42-g sample of equilin (which contains only carbon, hydrogen, and oxygen) produces 39.61 g CO_2 and 9.01 g H_2O. The molar mass of equilin is 268.34 g/mol. Find its molecular formula.

132. Estrone, which contains only carbon, hydrogen, and oxygen, is a female sexual hormone in the urine of pregnant women. Combustion analysis of a 1.893-g sample of estrone produces 5.545 g of CO_2 and 1.388 g H_2O. The molar mass of estrone is 270.36 g/mol. Find its molecular formula.

133. Epsom salts is a hydrated ionic compound with the following formula: $MgSO_4 \cdot x\,H_2O$. A 4.93-g sample of Epsom salts is heated to drive off the water of hydration. The mass of the sample after complete dehydration is 2.41 g. Find the number of waters of hydration (x) in Epsom salts.

134. A hydrate of copper(II) chloride has the following formula: $CuCl_2 \cdot x\,H_2O$. The water in a 3.41-g sample of the hydrate is driven off by heating. The remaining sample has a mass of 2.69 g. Find the number of waters of hydration (x) in the hydrate.

135. A compound of molar mass 177 g/mol contains only carbon, hydrogen, bromine, and oxygen. Analysis reveals that the compound contains eight times as much carbon as hydrogen by mass. Find the molecular formula.

136. Researchers obtained the following data from experiments tc find the molecular formula of benzocaine, a local anesthetic which contains only carbon, hydrogen, nitrogen, and oxygen Complete combustion of a 3.54-g sample of benzocaine witł excess O_2 forms 8.49 g of CO_2 and 2.14 g H_2O. Another 2.35-g sample contains 0.199 g of N. The molar mass of benzocaine i: 165 g/mol. Find the molar formula of benzocaine.

137. Find the total number of atoms in a sample of cocaine hydro chloride, $C_{17}H_{22}ClNO_4$, of mass 23.5 mg.

138. Vanadium forms four different oxides in which the percent by mass of vanadium is, respectively, (a) 76%, (b) 68%, (c) 61%, anc (d) 56%. Determine the formula and the name of each oxide.

139. The chloride of an unknown metal is believed to have the for mula MCl_3. A 2.395-g sample of the compound contain 3.606×10^{-2} mol Cl. Find the atomic mass of M.

140. Write the structural formulas of three different compounds tha each have the molecular formula C_5H_{12}.

141. A chromium-containing compound has the formula Fe_xCr_yO and is 28.59% oxygen by mass. Find x and y.

142. A phosphorus compound that contains 34.00% phosphorus b mass has the formula X_3P_2. Identify the element X.

143. A particular brand of beef jerky contains 0.0552% sodium n trite by mass and is sold in an 8.00-oz bag. What mass of sodiur does the sodium nitrite contribute to the sodium content of th bag of beef jerky?

144. Phosphorus is obtained primarily from ores containing calciur phosphate. If a particular ore contains 57.8% calcium pho phate, what minimum mass of the ore must be processed to ol tain 1.00 kg of phosphorus?

Challenge Problems

145. A mixture of NaCl and NaBr has a mass of 2.00 g and contains 0.75 g of Na. What is the mass of NaBr in the mixture?

146. Three pure compounds form when 1.00-g samples of element X combine with, respectively, 0.472 g, 0.630 g, and 0.789 g of element Z. The first compound has the formula X_2Z_3. Find the empirical formulas of the other two compounds.

147. A mixture of $CaCO_3$ and $(NH_4)_2CO_3$ is 61.9% CO_3 by mass. Find the mass percent of $CaCO_3$ in the mixture.

148. A mixture of 50.0 g of S and 1.00×10^2 g of Cl_2 reacts completely to form S_2Cl_2 and SCl_2. Find the mass of S_2Cl_2 formed.

149. Because of increasing evidence of damage to the ozone layer, chlorofluorocarbon (CFC) production was banned in 1996. However, many older cars still have air conditioners that use CFC-12 (CF_2Cl_2). These air conditioners are recharged from stockpiled supplies of CFC-12. Suppose that 100 million automobiles each contain 1.1 kg of CFC-12 and leak 25% of their CFC-12 into the atmosphere per year. How much chlorine, in kg, is added to the atmosphere each year due to these air conditioners? (Assume two significant figures in your calculations.)

150. A particular coal contains 2.55% sulfur by mass. When the coal is burned, it produces SO_2 emissions, which combine with rainwater to produce sulfuric acid. Use the formula of sulfuric acid to calculate the mass percent of S in sulfuric acid. Then determine how much sulfuric acid (in metric tons) is produced by the combustion of 1.0 metric ton of this coal. (A metric ton is 1000 kg.)

151. Lead is found in Earth's crust as several different lead ores. Suppose a certain rock is 38.0% PbS (galena), 25.0% $PbCO_3$ (cerussite), and 17.4% $PbSO_4$ (anglesite). The remainder of the rock is composed of substances containing no lead. How much of this rock (in kg) must be processed to obtain 5.0 metric tons of lead? (A metric ton is 1000 kg.)

152. A 2.52-g sample of a compound containing only carbon, hydrogen, nitrogen, oxygen, and sulfur is burned in excess oxygen to yield 4.23 g of CO_2 and 1.01 g of H_2O. Another sample of the same compound, of mass 4.14 g, yields 2.11 g of SO_3. A third sample, of mass 5.66 g, yields 2.27 g of HNO_3. Calculate the empirical formula of the compound.

153. A compound of molar mass 229 g/mol contains only carbon, hydrogen, iodine, and sulfur. Analysis shows that a sample of the compound contains six times as much carbon as hydrogen, by mass. Calculate the molecular formula of the compound.

154. The elements X and Y form a compound that is 40% X and 60% Y by mass. The atomic mass of X is twice that of Y. What is the empirical formula of the compound?

155. A compound of X and Y is $\frac{1}{3}$ X by mass. The atomic mass of element X is $\frac{1}{3}$ the atomic mass of element Y. Find the empirical formula of the compound.

156. A mixture of carbon and sulfur has a mass of 9.0 g. Complete combustion with excess O_2 gives 23.3 g of a mixture of CO_2 and SO_2. Find the mass of sulfur in the original mixture.

Conceptual Problems

157. When molecules are represented by molecular models, what does each sphere represent? How big is the nucleus of an atom in comparison to the sphere used to represent an atom in a molecular model?

158. Without doing any calculations, determine which element in each compound has the highest mass percent composition.
   a. CO
   b. N_2O
   c. $C_6H_{12}O_6$
   d. NH_3

159. Explain the problem with the following statement and correct it: "The chemical formula for ammonia (NH_3) indicates that ammonia contains three grams of hydrogen for each gram of nitrogen."

160. Explain the problem with the following statement and correct it. "When a chemical equation is balanced, the number of molecules of each type on both sides of the equation is equal."

161. Without doing any calculations, arrange the elements in H_2SO_4 in order of decreasing mass percent composition.

162. Element A is an atomic element, and element B is a diatomic molecular element. Using circles to represent atoms of A and squares to represent atoms of B, draw molecular-level views of each element.

Active Classroom Learning

Questions for Group Work

Discuss these questions with the group and record your consensus answer.

163. With group members playing the roles of nuclei and electrons, demonstrate the formation of an ionic bond between Na and Cl. Demonstrate the formation of the covalent bonds in H_2O.

164. Create a flowchart with a series of simple questions that can be used to determine whether a chemical formula is that of an atomic element, a molecular element, a molecular compound, or an ionic compound. Use your flowchart to identify the correct category for P_4, KCl, CH_4, Ne, and NH_4NO_3.

165. Have each member of your group list one similarity or difference between the naming conventions for ionic and molecular compounds.

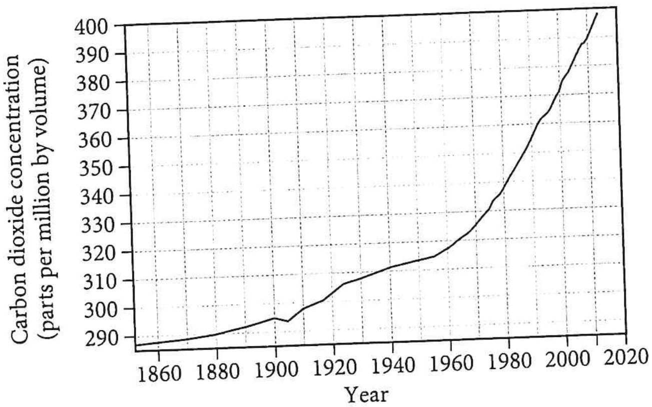
166. A compound isolated from the rind of lemons is found to be 88.14% carbon and 11.86% hydrogen by mass. How many grams of C and H are there in a 100.0-g sample of this substance? How many moles of C and H? What is the empirical formula? The molar mass is determined to be 136.26 g/mol. What is the molecular formula? Which step of the process just described does your group understand the least? Which step will be hardest for the members of your group to remember?

167. Octane (C_8H_{18}), a component of gasoline, reacts with oxygen to form carbon dioxide and water. Write the balanced chemical reaction for this process by passing a single piece of paper around your group, having each group member complete the next logical step. As you each complete your step, explain your reasoning to the group.

Data Interpretation and Analysis

Carbon and Fossil Fuel Consumption

168. Since the 1800s, the concentration of carbon dioxide in Earth's atmosphere has been increasing (Figure a ▼). The scientific consensus is that the increase in CO_2 is due to fossil fuel combustion.



▲ FIGURE a Carbon Dioxide Concentration in the Atmosphere

Use the information provided in the figure to answer the following questions (Note that 1 ppmv = 1 gas volume/1 × 10^6 air volumes = 1 μL gas/1 L air):

a. Globally, in 2010 and 2012, respectively, there were about 9.20 billion metric tons and 9.60 billion metric tons of carbon in Earth's atmosphere.
   i. Determine the increase in carbon in the atmosphere in units of metric tons between 2010 and 2012.
   ii. Determine the increase of carbon dioxide in the atmosphere in units of metric tons between 2010 and 2012.

b. Figure a lists the concentration of CO_2 in air in terms of parts per million by volume (ppmv).
   i. Compare the volume (in mL) of CO_2 in 1.00 m^3 of air in 1880 to the volume (in mL) of CO_2 in 1.00 m^3 of air in 2014.
   ii. Determine the increase in volume (in mL) of CO_2 in 1.00 m^3 of air between 1880 and 2014.
   iii. What is the percent increase of CO_2 in 1.00 m^3 of air between 1880 and 2014?
   iv. Assuming that the total volume of air in the atmosphere around Earth is 5.1×10^9 km^3 and that the concentration of CO_2 is uniform throughout the atmosphere, estimate the volume of CO_2 in the atmosphere in 2014 in units of km^3.

c. Based on the information in Figure a, determine the average yearly increase in the concentration of CO_2 in the atmosphere for the period 1960 to 2014. Estimate the concentration of CO_2 in the atmosphere, in units of ppmv, in 2030.

Answers to Conceptual Connections

Types of Chemical Bonds

3.1 The bond is covalent because it is forming between two nonmetals.

Structural Formulas

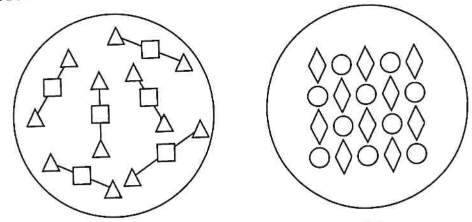
3.2 H—O—H

Representing Molecules

3.3 The spheres represent the electron cloud of the atom. It would be nearly impossible to draw a nucleus to scale on any of the space-filling molecular models in this book—the nucleus would be too small to see.

A Molecular View of Elements and Compounds

3.4



A₂B CD

Ionic and Molecular Compounds

3.5 Choice **(a)** best describes the difference between ionic and molecular compounds. Answer **(b)** is incorrect because there are no "new" forces in bonding (just rearrangements that result in lower potential energy) and because ions do not group together in pairs in the solid phase. Answer **(c)** is incorrect because the main difference between ionic and molecular compounds is the way that the atoms bond. Answer **(d)** is incorrect because ionic compounds do not contain molecules.

Nomenclature

3.6 This conceptual connection addresses one of the main errors you can make in nomenclature: the failure to correctly categorize the compound. Remember that you must first determine whether the compound is an ionic compound, a molecular compound, or an acid, and then you must name it accordingly. NCl_3 is a molecular compound (two or more nonmetals), and therefore in its name prefixes indicate the number of each type of atom—so NCl_3 is nitrogen trichloride. The compound $AlCl_3$, however, is an ionic compound (metal and nonmetal), and therefore does not require prefixes—so $AlCl_3$ is aluminum chloride.

Molecular Models and the Size of Molecules

3.7 **(c)** Atomic radii range in the hundreds of picometers, while the spheres in these models have radii of about a centimeter. The scaling factor is therefore about 10^8 (100 million).

Chemical Formula and Mass Percent Composition

3.8 C > O > H Since carbon and oxygen differ in atomic mass by only 4 amu, and since there are six carbon atoms in the formula, we can conclude that carbon constitutes the greatest fraction of the mass. Oxygen is next because its mass is 16 times that of hydrogen and there are only six hydrogen atoms for every one oxygen atom.

Chemical Formulas and Elemental Composition

3.9 **(c)** The chemical formula for a compound gives relationships between *atoms* or *moles of atoms*. The chemical formula for water states that water molecules contain two H atoms to every one O atom or 2 mol H to every 1 mol H_2O. This *does not* imply a two-to-one relationship between *masses* of hydrogen and oxygen because these atoms have different masses. It also does not imply a two-to-one relationship between volumes.

Balanced Chemical Equations

3.10 Both **(a)** and **(d)** are correct. When the number of atoms of each type is balanced, the sum of the masses of the substances will be the same on both sides of the equation. Since molecules change during a chemical reaction, the number of molecules is not the same on both sides, nor is the number of moles necessarily the same.