Read this carefully. It contains a lot of very important information.

Welcome to AP Chemistry: The AP Class you have signed up for is the equivalent of a 2 semester (1 year) general level college Chemistry course. This course is dedicated to teaching you to think like a chemist. The laboratory experiments will emphasize real-world chemical applications and will also help you on the Chemistry AP exam.

The **REQUIRED** textbook for this class is an eText with Online Mastering Chemistry. The information for purchasing this is on the JSerra website in the Parents and Students Den under textbooks. Follow the link for AP Chemistry in the downloadable list. It is the $55 option.

OPTIONAL Hardcover book: Tro; Chemistry: A Molecular Approach AP® Edition 5e 2019. Also available through the same link as above. $200.

First week of class: We will very briefly review the summer assignment topics. There will be two quizzes on this material the first week of class. We will have an exam on chapters 3.1-3.10 and 4.2-4.5 and 5.2-5.8 the second week of class and begin chapter 5.9 immediately after.

Zero Period: To be successful and fair to everyone, it is imperative you come to the zero period class on time. Class policy is if you are late for zero period, there will be a 20% penalty on any activity. If you miss half the lab, you get 50% off etc. No excuses are accepted for tardies or absences.

ABSENCE POLICY FOR WEDNESDAY/THURSDAY CLASS ZERO PERIOD LABORATORY CLASS:

A MAXIMUM OF ONE LABORTORY CLASS PERIOD CAN BE MISSED EACH SEMESTER. ADDITIONAL ABSENCES WILL RESULT IN A ZERO FOR THE LAB. NO EXCEPTIONS.

NO EXCUSES ARE ACCEPTABLE. ABSENCES DUE TO ILLNESS, TRAFFIC, CHOIR, CARITAS CHRISTI, ATHLETICS, VACATIONS, DENTIST AND DOCTOR APPOINTMENTS, FIELD TRIPS AND COLLEGE VISITS ARE NOT ACCEPTED.

IF YOU CAN'T BE ON TIME TO ZERO PERIOD ON WED/THURS, PLEASE DROP THE CLASS NOW!!!!!!

This is an AP class and the normal rules don't apply and you have been warned!!!!!!!!!

If you can be on time, we are going to have fun and learn a lot of chemistry.

After the AP exam, there will be no more zero periods.

Summer Assignment: To properly prepare you for the Chemistry AP Exam, I am providing you with a summer assignment consisting of chapters 3, 4 and 5 problems from your text book. This assignment is due the first day of school, at the beginning of class. These chapters are mostly review from chemistry honors. Answers to some problems have been provided. Remember, for credit you must show your calculations with all units.

Read chapters 3, 4 and 5 (attached). Take COMPLETE and AMAZING NOTES and complete the following problems with all work shown for each problem. No work, no credit, period.

Chapter 3 problems: p. 130-135 #27, 29, 33-53 odd, 67, 69, 89, 97, 99, 119, 121, 123

Chapter 4 problems: p 160-163 #15, 19, 41, 45, 47, 61, 67

Chapter 5 problems: p 204-207 #23, 29, 37, 41, 45, 47, 53, 59, 77, 87

Sincerely, Mr. Fick

dfick@jserra.org

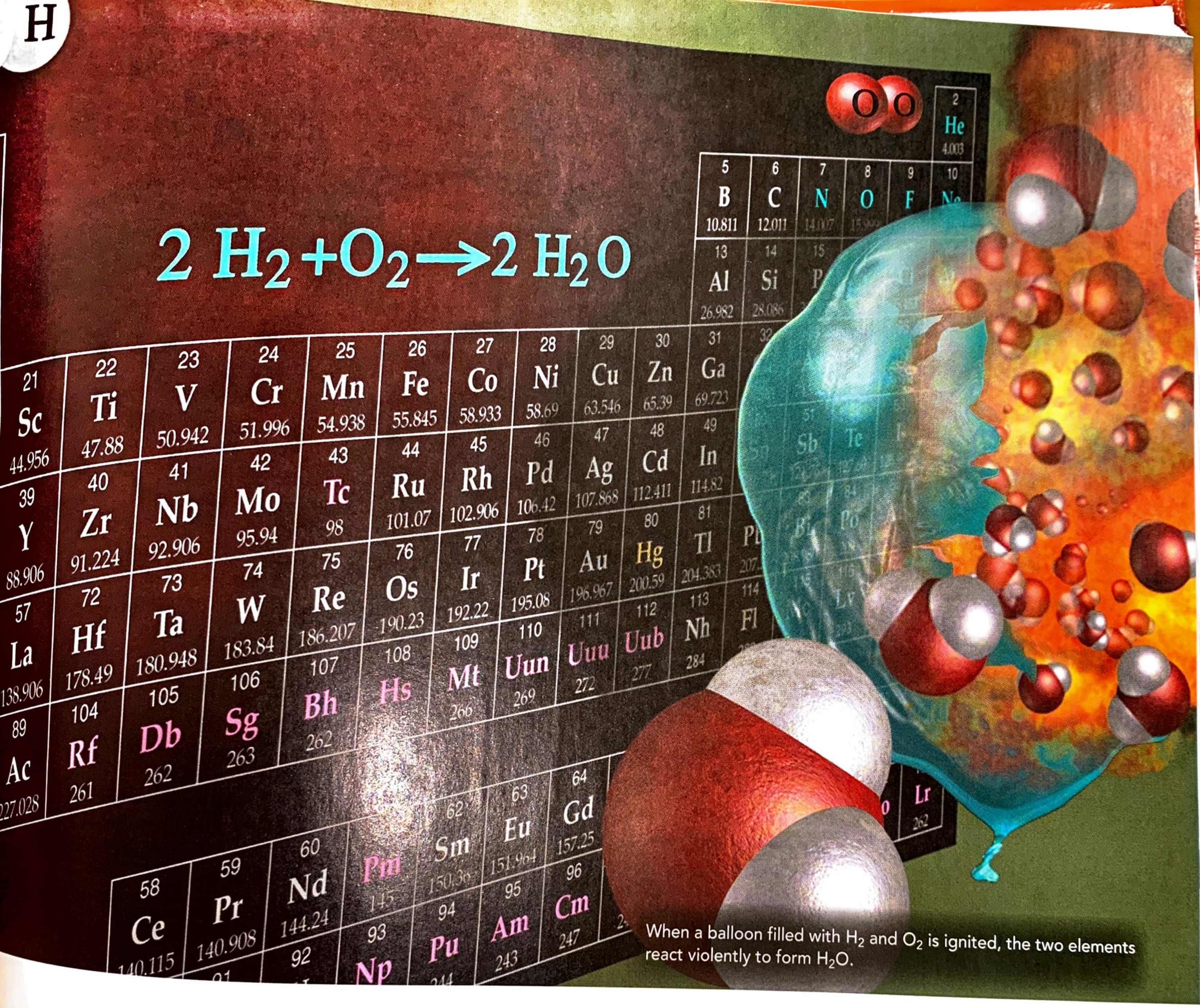
Almost all aspects of life are engineered at the molecular level, and without understanding molecules we can only have a very sketchy understanding of life itself.

—FRANCIS HARRY COMPTON CRICK (1916–2004)

CHAPTER 3

Molecules and Compounds

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.



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

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 that follows.

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

When hydrogen and oxygen combine to form the compound water (H_2O), however, a dramatically different substance results. 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.

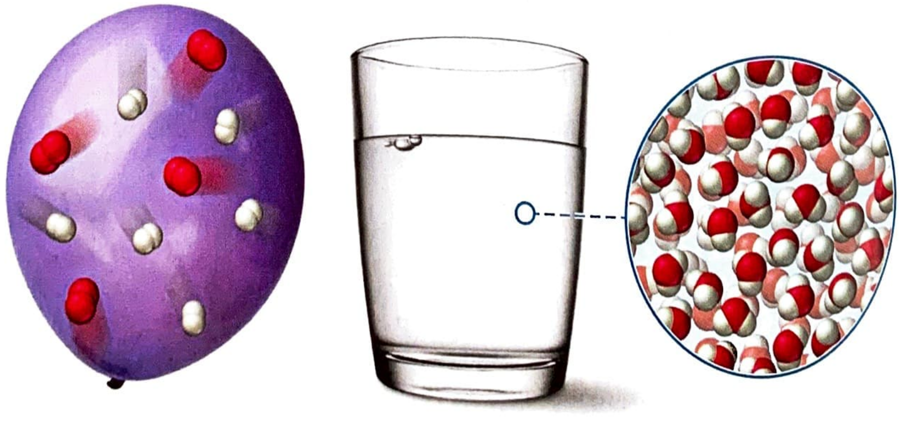
Although some of the substances that we encounter in everyday life are elements, most are compounds. As we discussed in Chapter 1, a compound is different from a mixture of elements. In a compound, elements combine in fixed, definite proportions; in a mixture, elements can mix in any proportions whatsoever. 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

Mixtures and Compounds

Hydrogen and Oxygen Mixture: This mixture can have any ratio of hydrogen to oxygen.

Water (a compound): Water molecules have a fixed ratio of atoms—2 hydrogens to 1 oxygen.



▶ **FIGURE 3.1 Mixtures and Compounds**

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 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 10 (see Section 10.2). 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.

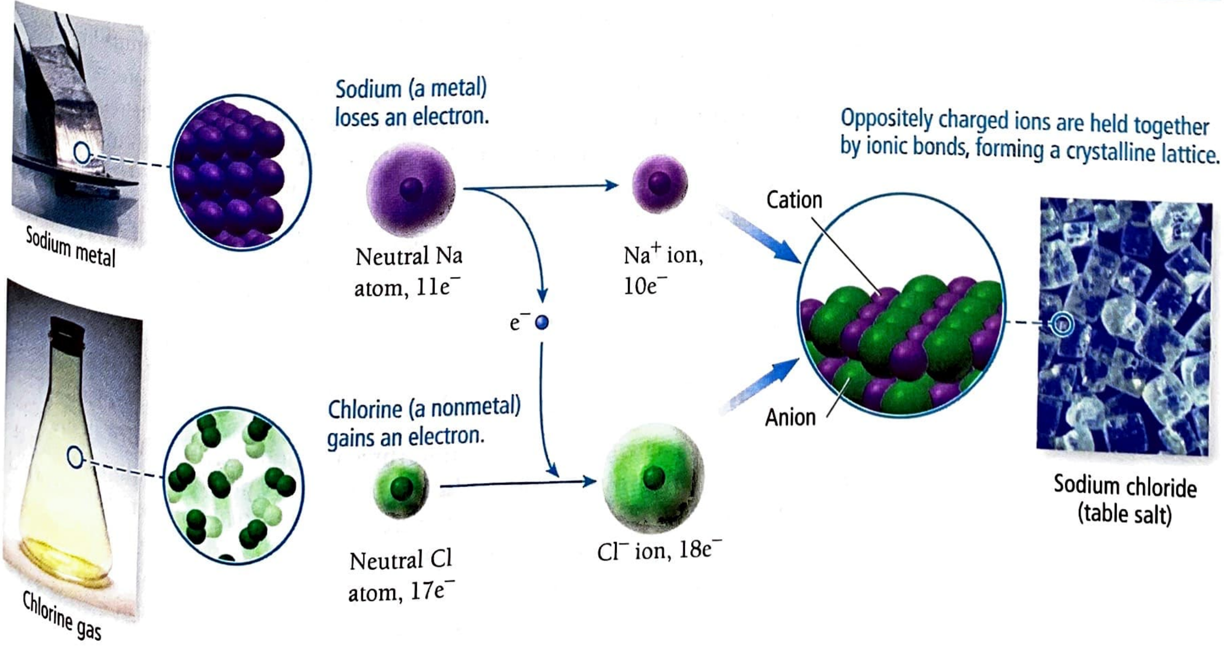
We can broadly classify most chemical bonds into two types: ionic and covalent. *Ionic bonds*—which occur between metals and nonmetals—involve the *transfer* of electrons from one atom to another. *Covalent bonds*—which occur between two or more nonmetals—involve the *sharing* of electrons between two 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 by electrostatic forces and form an **ionic bond**. The result is an **ionic compound**, which in the solid phase is composed of a lattice—a regular three-dimensional array—of alternating cations and anions.

▼ FIGURE 3.2 The Formation of an Ionic Compound

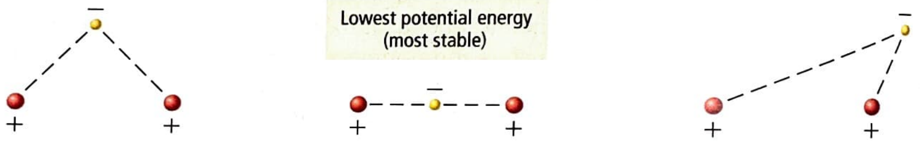
The Formation of an Ionic Compound



Covalent Bonds

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

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

ANSWER **NOW!**

3.1

Cc

Conceptual Connection

TYPES OF CHEMICAL BONDS What type of bond—ionic or covalent—forms between nitrogen and oxygen?

(a) Ionic
(b) Covalent

3.3 Representing Compounds: Chemical Formulas and Molecular Models

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

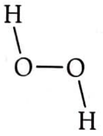
Types of Chemical Formulas

We can categorize chemical formulas into three different types: empirical, molecular, and structural. An **empirical formula** gives the *relative* number of atoms of each element in a compound. A **molecular formula** gives the *actual* number of atoms of each element in a molecule of a compound. For example, the empirical formula for hydrogen peroxide is HO, but its molecular formula is H_2O_2. The molecular formula is always a 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:

$$\mathrm{H{-}O{-}O{-}H}$$

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:

$$\mathrm{O{=}C{=}O}$$

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 10.

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 Select the structural formula for water.

(a) $\mathrm{H{-}O}$

(b) $\mathrm{H{-}H}$

(c) $\mathrm{H{-}O{-}H}$

(d) H_2O

3.2
Cc
Conceptual Connection

ANSWER **NOW!**

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.

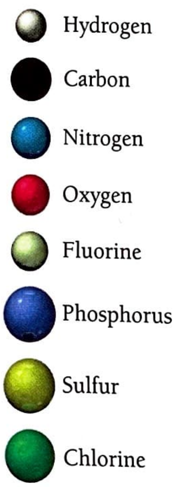
(b) For B_2H_6, the greatest common factor is 2. The empirical formula is therefore BH_3.

(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)** $C_2H_4O_2$

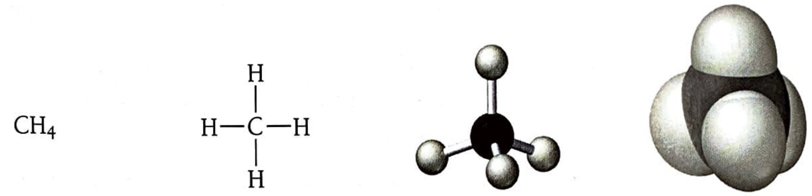
Answers to For Practice and For More Practice problems can be found in Appendix IV.



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:



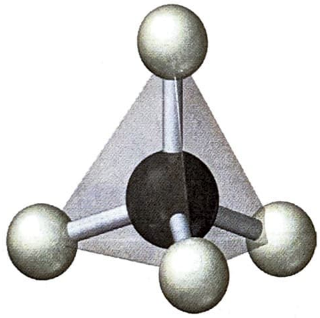
CH_4

$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

Molecular formula Structural formula Ball-and-stick model Space-filling model

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: that 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.



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

ANSWER **NOW!**

3.3
Cc
Conceptual Connection

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?

(a) Each sphere represents the hard outer shell of an atom. The nucleus would be too small to see on the same scale.

(b) Each sphere represents the electron cloud of the atom. The nucleus would be too small to see on the same scale.

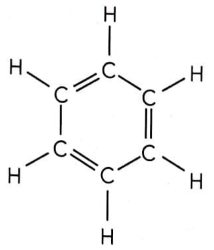
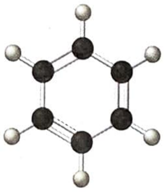
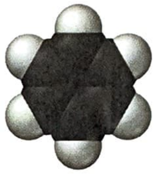
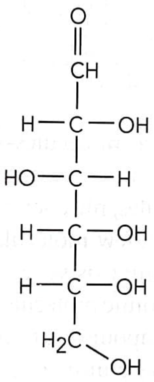
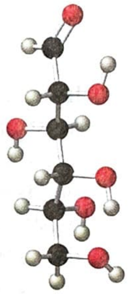
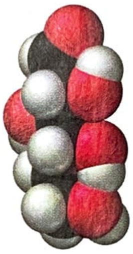
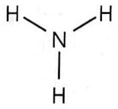
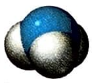
(c) Each sphere represents the nucleus of an atom. The nucleus is the same size as the sphere.

3.4 An Atomic-Level View of Elements and Compounds

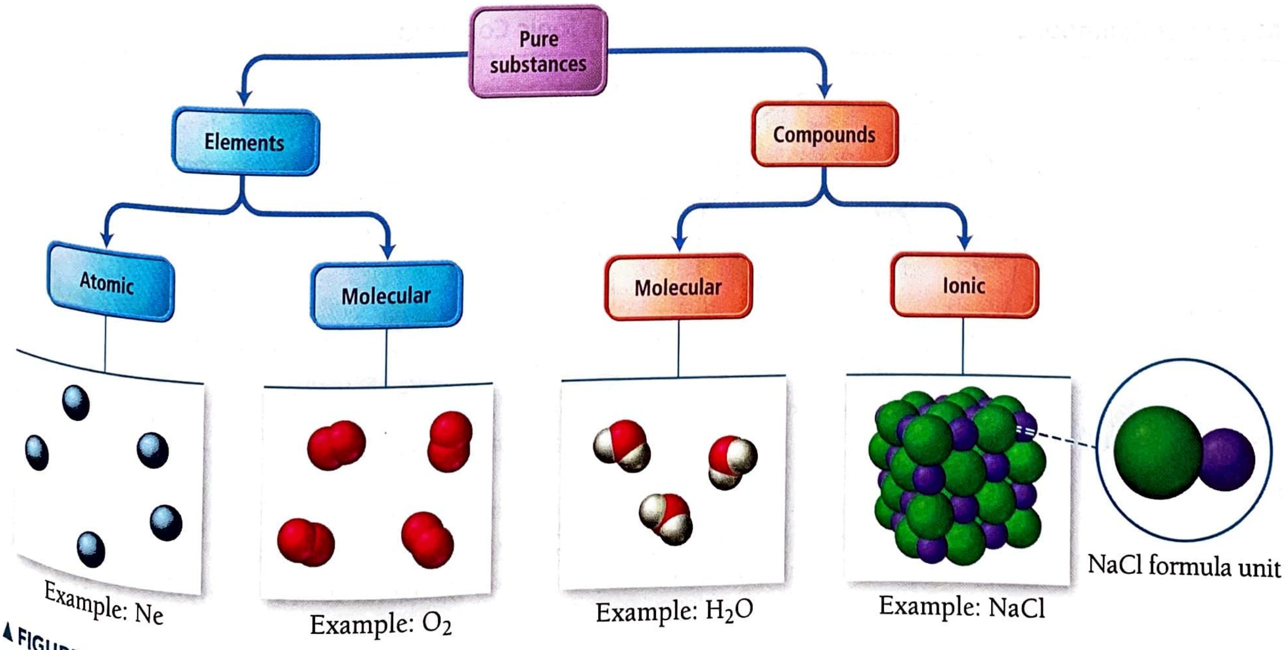
Recall from Chapter 1 that we categorize pure substances as either elements or compounds. We can subcategorize elements and compounds according to the basic units that compose them, as shown in Figure 3.4►. Elements may be either atomic or molecular. Compounds may be either molecular or ionic.

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

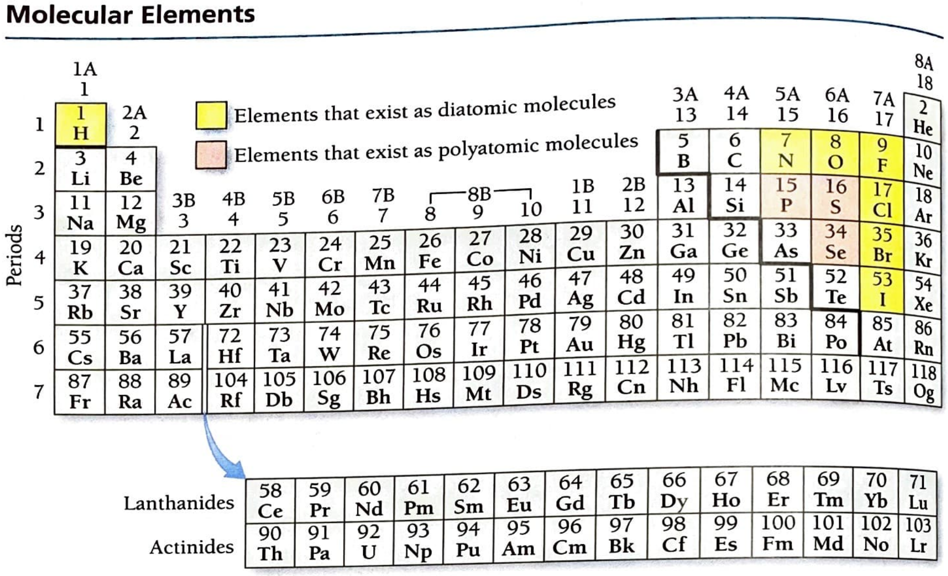
Name of Compound	Empirical Formula	Molecular Formula	Structural Formula	Ball-and-Stick Model	Space-Filling Model
Benzene	CH	C_6H_6			
Acetylene	CH	C_2H_2	$H-C\equiv C-H$		
Glucose	CH_2O	$C_6H_{12}O_6$			
Ammonia	NH_3	NH_3			

Classification of Elements and Compounds



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

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



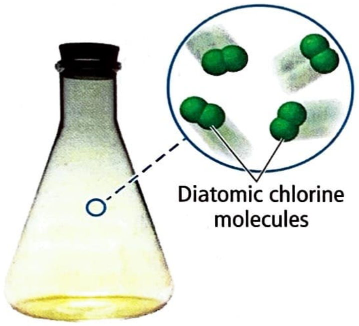
Molecular Elements

🟨 Elements that exist as diatomic molecules

🟧 Elements that exist as polyatomic molecules

Periods	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8	8B 9	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 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

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



▲ The basic units that compose chlorine gas are diatomic molecules (Cl_2).

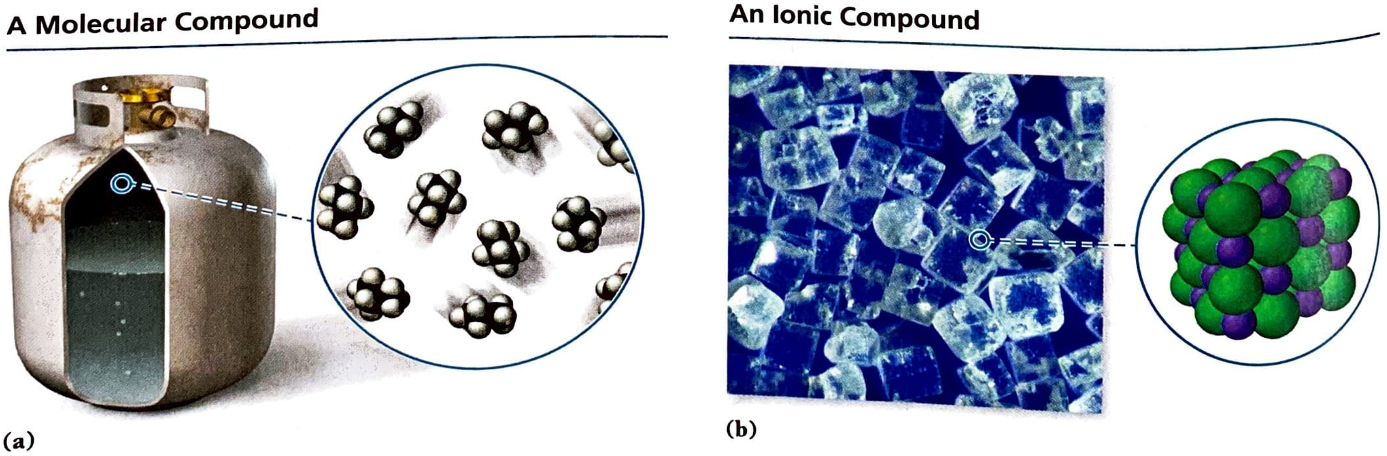
Instead, they exist as molecules—two or more atoms of the element bonded together. Most molecular elements exist as *diatomic* molecules. For example, hydrogen is composed of H_2 molecules, nitrogen is composed of N_2 molecules, and chlorine is composed of Cl_2 molecules. A few molecular elements exist 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

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.



A Molecular Compound

An Ionic Compound

▲ 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.

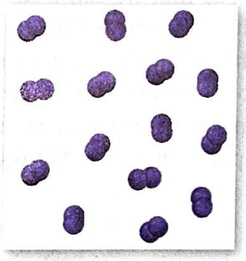
Many common ionic compounds contain ions that are themselves composed of a group of covalently bonded atoms with an overall charge. For example, the active ingredient in household bleach is sodium hypochlorite, which acts to chemically alter color-causing molecules in clothes (bleaching action) and to 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

Classify the substance represented by the molecular view shown here.

(a) Atomic element
(b) Molecular element
(c) Molecular compound
(d) Ionic compound



3.4 Conceptual Connection

ANSWER **NOW!**

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

ANSWER **NOW!**

3.5

Cc

Conceptual Connection

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. Ionic compounds contain nondirectional ionic bonds, which result (in the solid state) in the formation of ionic lattices.

(b) Molecular compounds and ionic compounds both contain molecules as their smallest identifiable unit, but in ionic compounds the molecules are smaller.

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

WATCH **NOW!**

KEY CONCEPT VIDEO 3.5

Naming Ionic Compounds

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.

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

▲ 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 ($Na_2CO_3 \cdot NaHCO_3 \cdot 2\,H_2O$).

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

Writing Formulas for Ionic Compounds

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.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE VIDEO 3.3

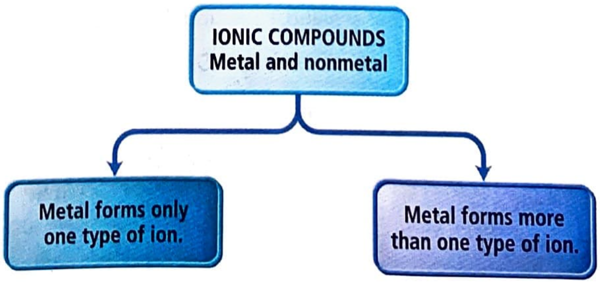
HOW TO: *Write Formulas for Ionic Compounds*	**EXAMPLE 3.3** **Writing Formulas for Ionic Compounds**	**EXAMPLE 3.4** **Writing Formulas for Ionic Compounds**
	Write the formula for the ionic compound that forms between aluminum and oxygen.	Write the formula for the ionic compound that forms between 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.

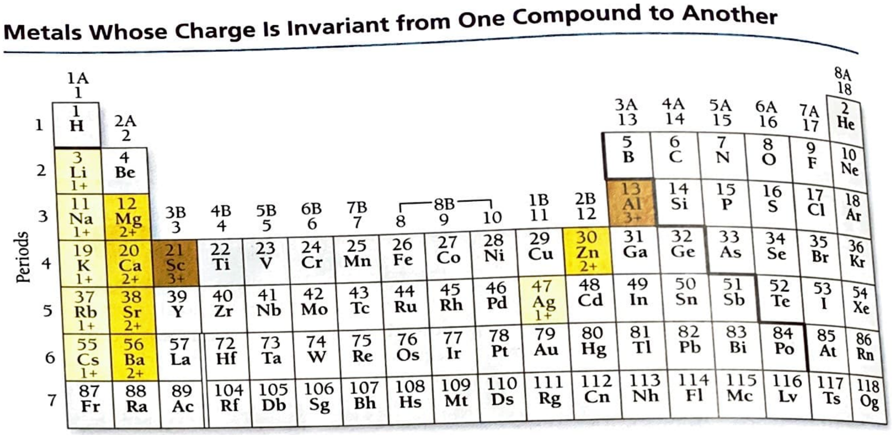
Naming Ionic Compounds

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.

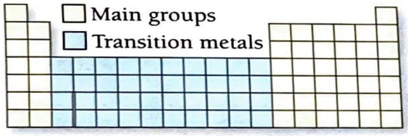
Before naming an ionic compound you must identify it as one. Remember, *ionic compounds are usually composed of metals and nonmetals*; any time you see a metal and one or more nonmetals together in a chemical formula, assume that you have an ionic compound. Ionic compounds can be categorized into two 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 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.





▲ **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.

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.

ANSWER **NOW!**

3.6

Cc

Conceptual Connection

TYPES OF METALS Which metal has the same charge in all of its compounds?

(a) Fe **(b)** Mo
(c) Pb **(d)** Sr

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^-	*chlorite*
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.

3.7 Conceptual Connection ANSWER **NOW!**

POLYATOMIC IONS Identify the polyatomic ion and its charge in each compound: KNO_2, $CaSO_4$, $Mg(NO_3)_2$.

(a) NO_2^-, SO_4^{2-}, and NO_3^-

(b) K^+, Ca^{2+}, and Mg^{2+}

(c) K^+, Ca^{2+}, Mg^{2+}, NO_2^-, SO_4^{2-}, and NO_3^-

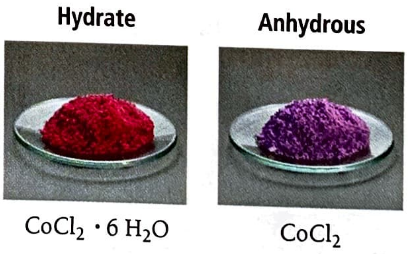
(d) NO_2^{2-}, SO_4^-, and NO_3^{2-}

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



▲ **FIGURE 3.9 Hydrates** Heating pink cobalt(II) chloride hexahydrate removes the waters of hydration to produce blue 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

WATCH **NOW!**

KEY CONCEPT VIDEO 3.6
Naming Molecular Compounds

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 10, we will discuss the stability of these various combinations of the same elements. For now, we focus on naming a molecular compound based on its formula and writing its formula based on its name.

Naming Molecular Compounds

Like ionic compounds, many molecular compounds have common names. For example, H_2O and NH_3 have the common names *water* and *ammonia*. However, the sheer number of existing molecular compounds—numbering in the millions—necessitates a systematic approach to naming them.

The first step in naming a molecular compound is identifying it as one. Remember, *molecular compounds are composed of two or more nonmetals*. In this section, we discuss how to name binary (two-element) molecular compounds. Their names have the form:

prefix | **name of 1st element** | **prefix** | **base name of 2nd element + *-ide***

When writing the name of a molecular compound, as when writing the formula, we first list the more metal-like element (toward the left and bottom of the periodic table). Generally, we write the name of the element with the smaller group number first. If the two elements lie in the same group, then we write the element with the greater row number first. The prefixes given to each element indicate the number of atoms present:

These prefixes are the same as those used in naming hydrates.

mono = 1
di = 2
tri = 3
tetra = 4
penta = 5
hexa = 6
hepta = 7
octa = 8
nona = 9
deca = 10

If there is only one atom of the *first element* in the formula, the prefix *mono-* is normally omitted. For example, we name NO_2 according to the first element, *nitrogen*, with no prefix because *mono-* is omitted for the first element, followed by the prefix *di-*, to indicate two oxygen atoms, and the base name of the second element, *ox*, with the ending *-ide*. Its full name is *nitrogen dioxide*.

NO_2 nitrogen dioxide

When a prefix ends with "o" and the base name begins with "o," the first "o" is often dropped. For example, mono-oxide becomes *monoxide*.

We name the compound N_2O, sometimes called laughing gas, similarly except that we use the prefix *di-* before nitrogen to indicate two nitrogen atoms and the prefix *mono-* before oxide to indicate one oxygen atom. Its full name is *dinitrogen monoxide*.

N_2O dinitrogen monoxide

EXAMPLE 3.8 Naming Molecular Compounds

Name each compound.

(a) NI_3 **(b)** PCl_5 **(c)** P_4S_{10}

SOLUTION

(a) The name of the compound is the name of the first element, *nitrogen*, followed by the base name of the second element, *iod*, prefixed by *tri-* to indicate three and given the suffix *-ide*.

NI_3 nitrogen triiodide

(b) The name of the compound is the name of the first element, *phosphorus*, followed by the base name of the second element, *chlor*, prefixed by *penta-* to indicate five and given the suffix *-ide*.

PCl_5 phosphorus pentachloride

(c) The name of the compound is the name of the first element, *phosphorus*, prefixed by *tetra-* to indicate four, followed by the base name of the second element, *sulf*, prefixed by *deca-* to indicate ten and given the suffix *-ide*.

P_4S_{10} tetraphosphorus decasulfide

FOR PRACTICE 3.8 Name the compound N_2O_5.

FOR MORE PRACTICE 3.8 Write the formula for phosphorus tribromide.

3.8 Cc Conceptual Connection ANSWER **NOW!**

NOMENCLATURE The compound NCl_3 is nitrogen trichloride, but $AlCl_3$ is simply aluminum chloride. Why?

(a) The name forms differ because NCl_3 is an ionic compound and $AlCl_3$ is a molecular compound. Prefixes such as *mono-*, *di-*, and *tri-* are used for ionic compounds but not for molecular compounds.

(b) The name forms differ because NCl_3 is a molecular compound and $AlCl_3$ is an ionic compound. Prefixes such as *mono-*, *di-*, and *tri-* are used for molecular compounds but not for ionic compounds.

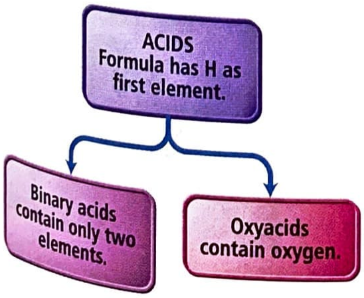
Naming Acids

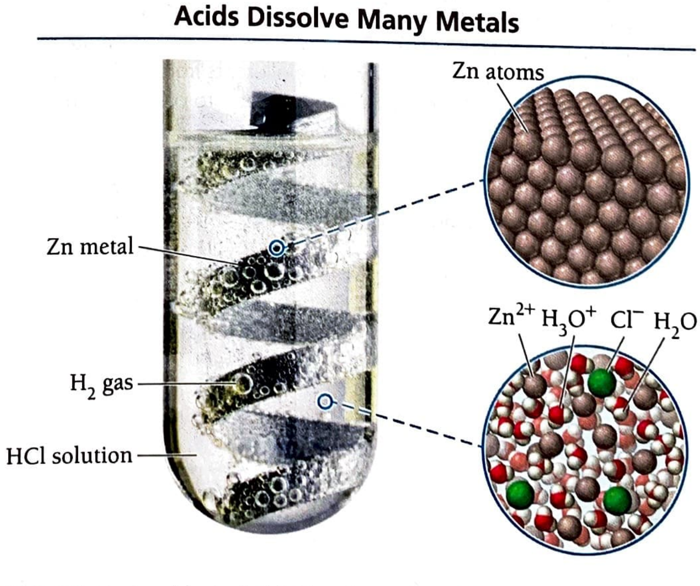
We can define acids in a number of ways, as we will discuss in Chapter 17. For now we define **acids** as molecular compounds that release hydrogen ions (H^+) when dissolved in water. Acids are composed of hydrogen, usually written first in their formula, and one or more nonmetals, written second. For example, HCl is a molecular compound that, when dissolved in water, forms $H^+(aq)$ and $Cl^-(aq)$ ions, where *aqueous* (aq) means *dissolved in water*. Therefore, HCl is an acid when dissolved in water. To distinguish between gaseous HCl (which is named hydrogen monochloride because it is a molecular compound) and HCl in solution (which is named hydrochloric acid because it is an acid), we write the former as $HCl(g)$ and the latter as $HCl(aq)$.

▲ Many fruits are acidic and have the characteristically sour taste of acids.

Acids are characterized by their sour taste and their ability to dissolve many metals. For example, hydrochloric acid is present in stomach fluids, and its sour taste becomes painfully obvious during vomiting. Hydrochloric acid also dissolves some metals. For example, if we put a strip of zinc into a test tube of hydrochloric acid, it slowly dissolves as the $H^+(aq)$ ions convert the zinc metal into $Zn^{2+}(aq)$ cations (Figure 3.10▶).

Acids are present in foods such as lemons and limes and are used in household products such as toilet bowl cleanser and Lime-A-Way. In this section, we discuss how to name them; in Chapter 17 you will learn more about their properties. We categorize acids into two types: binary acids and oxyacids.





▲ **FIGURE 3.10 Hydrochloric Acid Dissolving Zinc Metal** The zinc atoms are ionized to zinc ions, which dissolve in the water. The HCl forms H_2 gas, which is responsible for the bubbles you can see in the test tube.

Naming Binary Acids

Binary acids are composed of hydrogen and a nonmetal. Names for binary acids have the form:

hydro | **base name of nonmetal + *-ic*** | **acid**

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 and 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 sulfite (SO_3^{2-}) ion, and $H_2SO_4(aq)$ contains the sulfate (SO_4^{2-}) ion. Oxyacids are a combination of one or more H^+ ions with an oxyanion. The number of H^+ ions depends on the

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

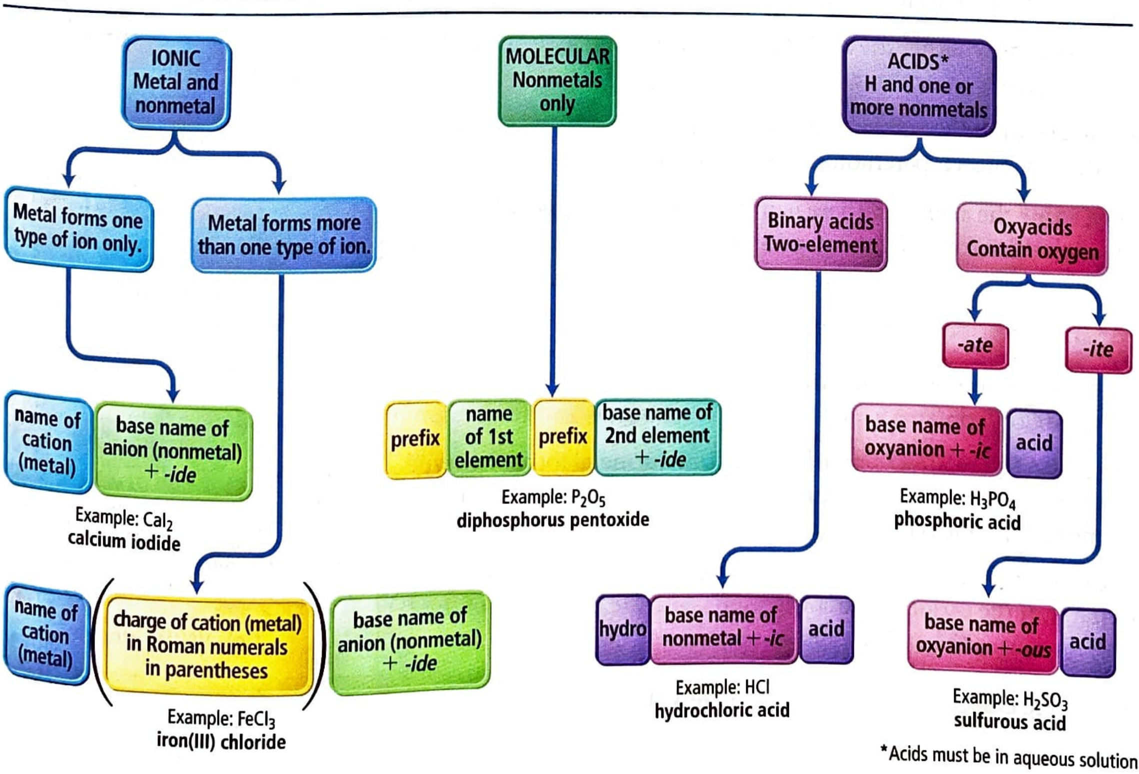
FOR PRACTICE 3.10 Name the acid $HNO_2(aq)$.

FOR MORE PRACTICE 3.10 Write the formula for perchloric acid.

3.7 Summary of Inorganic Nomenclature

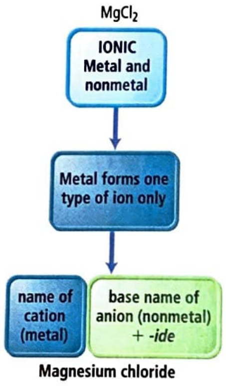
In Sections 3.5 and 3.6, we discussed naming inorganic compounds, specifically ionic compounds, molecular compounds, and acids. However, we often have to name a compound without initially knowing the category into which it falls. In other words, real-life nomenclature is a bit messier than the categorized 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.

Inorganic Nomenclature Flowchart



*Acids must be in aqueous solution.

▼ **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 trying to name. For example, to name the compound $MgCl_2$, you need to decide if the compound is ionic, molecular, or an acid. In this case, since $MgCl_2$ is composed of a metal and nonmetal, it is ionic. Therefore, you begin at the box labeled "IONIC" at the far left side of the flowchart.

Next, decide whether the metal in the compound forms only one type of ion or more than one type. You can determine this by looking for the metal (in this case magnesium) in Figure 3.7. Since magnesium is listed in the figure, it forms only one type of 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 the flowchart. In this case, write the name of the cation (the metal) followed by the base name of the anion (the nonmetal) appended with the ending *-ide*. Its full name is magnesium chloride.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 3.11

EXAMPLE 3.11 Using the Nomenclature Flowchart to Name Compounds

Use the flowchart in Figure 3.11 to name each compound.

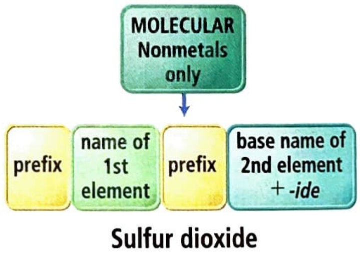
(a) SO_2 **(b)** $HClO_4(aq)$ **(c)** CoF_2

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*.



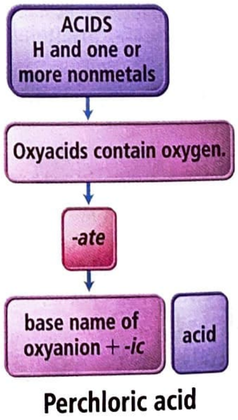
(b) $HClO_4(aq)$

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*.

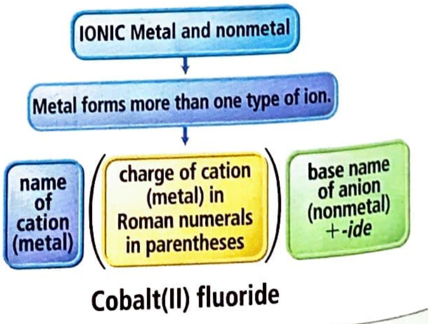


(c) CoF_2

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.

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

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

$$\begin{aligned}\text{Formula mass} &= 12.01\text{ amu} + \overset{\swarrow\text{ Multiply by 2 because formula has two oxygen atoms.}}{2(16.00\text{ amu})}\\ &= 44.01\text{ amu}\end{aligned}$$

and that of sodium oxide, Na_2O, is:

$$\begin{aligned}\text{Formula mass} &= \overset{\swarrow\text{ Multiply by 2 because formula has two sodium atoms.}}{2(22.99\text{ amu})} + 16.00\text{ amu}\\ &= 61.98\text{ amu}\end{aligned}$$

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.

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

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 use Avogadro's number to convert to number of molecules. The conceptual plan is as follows:

Conceptual Plan

$$\begin{tikzcd}
\boxed{\text{g } CO_2} \arrow[r, "\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}"'] & \boxed{\text{mol } CO_2} \arrow[r, "\frac{6.022 \times 10^{23}\ CO_2 \text{ molecules}}{1 \text{ mol } CO_2}"'] & \boxed{CO_2 \text{ molecules}}
\end{tikzcd}$$

To solve the problem, we follow the conceptual plan, beginning with 10.8 g CO_2, converting to moles, and then to molecules.

Solution

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

$$= 1.48 \times 10^{23}\ CO_2 \text{ molecules}$$

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 3.13

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

$$\begin{tikzcd}
\boxed{\text{mg } C_9H_8O_4} \arrow[r, "\frac{10^{-3}\text{ g}}{1 \text{ mg}}"'] & \boxed{\text{g } C_9H_8O_4} \arrow[r, "\frac{1 \text{ mol } C_9H_8O_4}{180.15 \text{ g } C_9H_8O_4}"'] & {} \\
 & \boxed{\text{mol } C_9H_8O_4} \arrow[r, "\frac{6.022 \times 10^{23}\ C_9H_8O_4 \text{ molecules}}{1 \text{ mol } C_9H_8O_4}"'] & \boxed{\text{number of } C_9H_8O_4 \text{ molecules}}
\end{tikzcd}$$

RELATIONSHIPS USED

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

$$6.022 \times 10^{23} = 1 \text{ mol}$$
$$1 \text{ mg} = 10^{-3} \text{ g}$$

SOLVE Follow the conceptual plan to solve the problem.

SOLUTION

$$325\ \cancel{\text{mg } C_9H_8O_4} \times \frac{10^{-3}\ \cancel{\text{g}}}{1\ \cancel{\text{mg}}} \times \frac{1\ \cancel{\text{mol } C_9H_8O_4}}{180.15\ \cancel{\text{g } C_9H_8O_4}} \times$$

$$\frac{6.022 \times 10^{23}\ C_9H_8O_4 \text{ molecules}}{1\ \cancel{\text{mol } C_9H_8O_4}} = 1.09 \times 10^{21}\ C_9H_8O_4 \text{ molecules}$$

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.

MOLECULAR MODELS AND THE SIZE OF MOLECULES

3.9
Cc
Conceptual Connection

ANSWER **NOW!**

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 here?

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

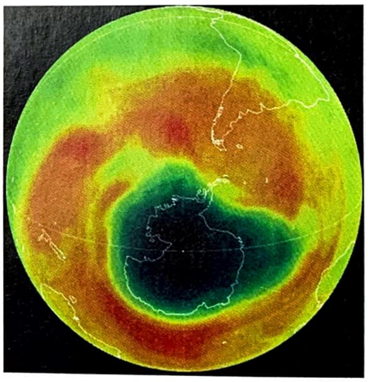
3.9 Composition of Compounds

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 40 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?

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:

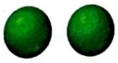
$$\text{Mass percent of element X} = \frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100\%$$



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

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:

$$\mathrm{CCl_{\textcircled{2}}F_2}$$



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

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:

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

EXAMPLE 3.14 Mass Percent Composition

Calculate the mass percent of Cl in Freon-112 ($C_2Cl_4F_2$), a CFC refrigerant.

SORT You are given the molecular formula of Freon-112 and asked to find the mass percent of Cl.

GIVEN: $C_2Cl_4F_2$

FIND: mass percent Cl

STRATEGIZE The molecular formula tells you that there are 4 mol of Cl in each mole of Freon-112. Find the mass percent composition from the chemical formula by using the equation that defines mass percent. The conceptual plan shows you how to use the mass of Cl in 1 mol of $C_2Cl_4F_2$ and the molar mass of $C_2Cl_4F_2$ to find the mass percent of Cl.

CONCEPTUAL PLAN

$$\text{Mass \% Cl} = \frac{4 \times \text{molar mass Cl}}{\text{molar mass } C_2Cl_4F_2} \times 100\%$$

RELATIONSHIPS USED

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

SOLVE Calculate the necessary parts of the equation and substitute the values into the equation to find mass percent Cl.

SOLUTION

$$4 \times \text{molar mass Cl} = 4(35.45\text{ g/mol}) = 141.8\text{ g/mol}$$

$$\begin{aligned}\text{Molar mass } C_2Cl_4F_2 &= 2(12.01\text{ g/mol}) + 4(35.45\text{ g/mol}) + 2(19.00\text{ g/mol}) \\ &= 24.02\text{ g/mol} + 141.8\text{ g/mol} + 38.00\text{ g/mol} = 203.8\text{ g/mol}\end{aligned}$$

$$\begin{aligned}\text{Mass \% Cl} &= \frac{4 \times \text{molar mass Cl}}{\text{molar mass } C_2Cl_4F_2} \times 100\% \\ &= \frac{141.8\ \cancel{\text{g/mol}}}{203.8\ \cancel{\text{g/mol}}} \times 100\% \\ &= 69.58\%\end{aligned}$$

CHECK The units of the answer (%) are correct. The magnitude is reasonable because it is between 0 and 100% and chlorine is the heaviest atom in the molecule and there are four atoms of it.

FOR PRACTICE 3.14 Acetic acid ($HC_2H_3O_2$) is the active ingredient in vinegar. Calculate the mass percent composition of oxygen in acetic acid.

FOR MORE PRACTICE 3.14 Calculate the mass percent composition of sodium in sodium oxide.

ANSWER **NOW!**

3.10

Cc

Conceptual Connection

CHEMICAL FORMULA AND MASS PERCENT COMPOSITION

Without doing any calculations, list the elements in C_6H_6O in order of decreasing mass percent composition.

(a) C>O>H **(b)** O>C>H **(c)** H>O>C **(d)** C>H>O

Mass Percent Composition as a Conversion Factor

The mass percent composition of an element in a compound is a conversion factor between mass of the element and mass of the compound. For example, we saw that the mass percent composition of Cl in CCl_2F_2 is 58.64%. Since percent means *per hundred*, there are 58.64 g Cl *per hundred* grams CCl_2F_2, which can be expressed as the ratio:

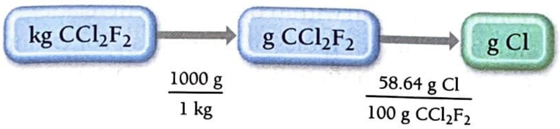
$$58.64\text{ g Cl} : 100\text{ g } CCl_2F_2$$

or, in fractional form:

$$\frac{58.64\text{ g Cl}}{100\text{ g } CCl_2F_2} \quad \text{or} \quad \frac{100\text{ g } CCl_2F_2}{58.64\text{ g Cl}}$$

These ratios can function as conversion factors between grams of Cl and grams of CCl_2F_2. For example, to calculate the mass of Cl in 1.00 kg CCl_2F_2, we use the following conceptual plan:

Conceptual Plan



Notice that the mass percent composition acts as a conversion factor between grams of the compound and grams of the constituent element. To calculate grams Cl, we follow the conceptual plan.

Solution

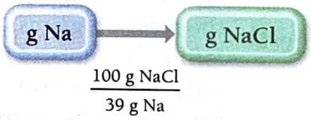
$$1.00\ \cancel{\text{kg}}\ \cancel{\text{CCl}_2\text{F}_2} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{58.64\text{ g Cl}}{100\ \cancel{\text{g}}\ \cancel{\text{CCl}_2\text{F}_2}} = 5.86 \times 10^2\text{ g Cl}$$

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 3.15

EXAMPLE 3.15 Using Mass Percent Composition as a Conversion Factor

The U.S. Food and Drug Administration (FDA) recommends that an adult consume less than 2.4 g of sodium per day. What mass of sodium chloride (in grams) can you consume and still be within the FDA guidelines? Sodium chloride is 39% sodium by mass.

SORT You are given a mass of sodium and the mass percent of sodium in sodium chloride. You are asked to find the mass of NaCl that contains the given mass of sodium.	**GIVEN:** 2.4 g Na **FIND:** g NaCl
STRATEGIZE Convert between mass of a constituent element and mass of a compound by using mass percent composition as a conversion factor.	**CONCEPTUAL PLAN**  **RELATIONSHIPS USED** 39 g Na : 100 g NaCl
SOLVE Follow the conceptual plan to solve the problem.	**SOLUTION** $2.4\text{ g}\ \cancel{\text{Na}} \times \frac{100\text{ g NaCl}}{39\text{ g}\ \cancel{\text{Na}}} = 6.2\text{ g NaCl}$ You can consume 6.2 g NaCl and still be within the FDA guidelines.

CHECK The units of the answer are correct. The magnitude seems reasonable because it is larger than the amount of sodium, as expected, because sodium is only one of the elements in NaCl.

FOR PRACTICE 3.15 What mass (in grams) of iron(III) oxide contains 58.7 g of iron? Iron(III) oxide is 69.94% iron by mass.

FOR MORE PRACTICE 3.15 If someone consumes 22 g of sodium chloride per day, what mass (in grams) of sodium does that person consume? Sodium chloride is 39% sodium by mass.

◀12.5 packets of salt contain 6.2 g of NaCl.

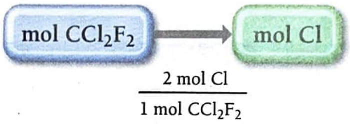
Conversion Factors from Chemical Formulas

Mass percent composition is one way to understand how much chlorine is in a particular chlorofluorocarbon or, more generally, how much of a constituent element is present in a given mass of any compound. However, we can also approach this type of problem in a different way. Chemical formulas contain within them inherent relationships between atoms (or moles of atoms) and molecules (or moles of molecules). For example, the formula for CCl_2F_2 tells us that 1 mol of CCl_2F_2 contains 2 mol of Cl atoms. We write the 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 directly determine the amounts of the constituent elements present in a given amount of a compound without having to calculate mass percent composition. For example, we calculate the number of moles of Cl in 38.5 mol of CCl_2F_2 as follows:

Conceptual Plan

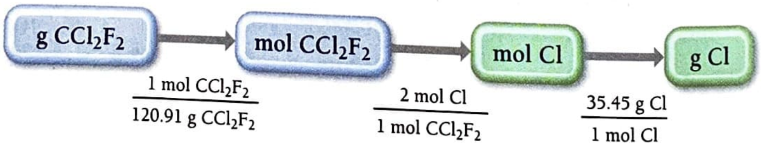


Solution

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

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

Conceptual Plan



Solution

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

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

The general form for solving problems in which we need to find the mass of an element 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 relationships inherent in the chemical to convert between moles and moles.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 3.16

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

$$\begin{tikzcd}
\text{gal } H_2O \arrow[r, "\frac{3.785\ L}{1\ gal}"'] & \text{L } H_2O \arrow[r, "\frac{1000\ mL}{1\ L}"'] & \text{mL } H_2O \arrow[r, "\frac{1.00\ g}{1\ mL}"'] & \text{g } H_2O
\end{tikzcd}$$

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.

$$\begin{tikzcd}
\text{g } H_2O \arrow[r, "\frac{1\ mol\ H_2O}{18.02\ g\ H_2O}"'] & \text{mol } H_2O \arrow[r, "\frac{2\ mol\ H}{1\ mol\ H_2O}"'] & \text{mol H} \arrow[r, "\frac{1.008\ g\ H}{1\ mol\ H}"'] & \text{g H}
\end{tikzcd}$$

RELATIONSHIPS USED

$3.785\ L = 1\ gal$

$1000\ mL = 1\ L$

$1.00\ g\ H_2O = 1\ mL\ H_2O$ (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

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

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

$$\times \frac{1.008\ g\ H}{1\ \cancel{mol\ H}} = 4.23 \times 10^2\ g\ H$$

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

3.11
Cc
Conceptual Connection

ANSWER **NOW!**

The molecular formula for water is H_2O. Which ratio can be correctly derived from this formula? Explain.

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

CHEMISTRY 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 for 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. Can we also do the reverse? Can we calculate a chemical formula from mass percent composition? This question is important because many laboratory analyses of compounds list the relative masses of each element present in the compound. For example, if we decompose water into hydrogen and oxygen in the laboratory, we can measure the masses of hydrogen and oxygen produced. Can we determine a chemical formula from these data? The answer is a qualified yes. We can determine a chemical formula, but it is an empirical formula (not a molecular formula). To get a molecular formula, we need additional information, such as the molar mass of the compound.

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

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

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

From these data, we know there are 0.849 mol H for every 0.429 mol O. We can now 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 subscripts 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.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE VIDEO 3.18

HOW TO: *Obtain 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.	$24.5\,\cancel{\text{g N}} \times \frac{1\text{ mol N}}{14.01\,\cancel{\text{g N}}} = 1.75\text{ mol N}$ $70.0\,\cancel{\text{g O}} \times \frac{1\text{ mol O}}{16.00\,\cancel{\text{g O}}} = 4.38\text{ mol O}$	$60.00\,\cancel{\text{g C}} \times \frac{1\text{ mol C}}{12.01\,\cancel{\text{g C}}} = 4.996\text{ mol C}$ $4.48\,\cancel{\text{g H}} \times \frac{1\text{ mol H}}{1.008\,\cancel{\text{g H}}} = 4.44\text{ mol H}$ $35.52\,\cancel{\text{g O}} \times \frac{1\text{ mol O}}{16.00\,\cancel{\text{g O}}} = 2.220\text{ mol O}$
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 $C_9H_8O_4$.

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

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?

Determining Molecular Formulas for Compounds

We can find the molecular formula of a compound from the empirical formula if we also know the molar mass of the compound. Recall from Section 3.3 that the molecular formula 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 fruit) from its empirical formula, CH_2O, and its molar mass, 180.2 g/mol. We know that the molecular formula is a whole-number multiple of CH_2O:

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

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

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

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

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

For fructose, the empirical formula molar mass is:

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

Therefore, *n* is:

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

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 Mass

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.	Molecular formula = empirical formula × *n* $n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$
SOLVE Calculate the empirical formula mass.	Empirical formula molar mass $= 2(12.01 \text{ g/mol}) + 3(1.008 \text{ g/mol}) + 16.00 \text{ g/mol} = 43.04 \text{ g/mol}$
Divide the molar mass by the empirical formula mass to find *n*.	$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}} = \frac{86.09 \cancel{\text{g/mol}}}{43.04 \cancel{\text{g/mol}}} = 2$
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:

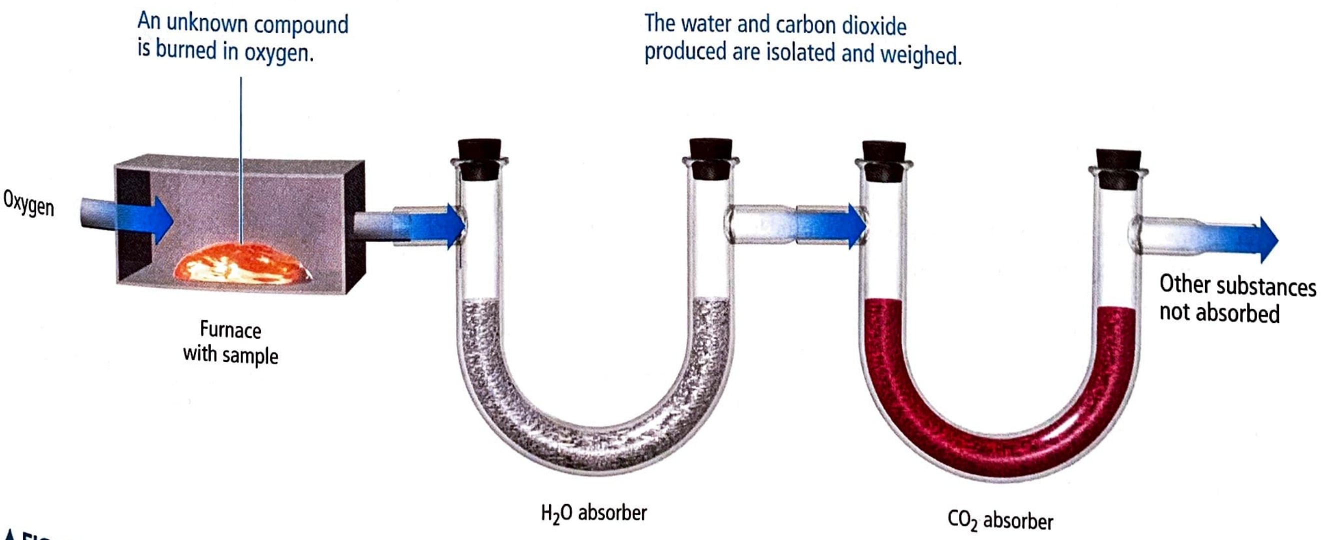
C, 39.97%
H, 13.41%
N, 46.62%

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 3.20 and 3.21 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 4.2.

Combustion Analysis



▲ **FIGURE 3.13 Combustion Analysis Apparatus**

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE VIDEO 3.21

HOW TO: Determine 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.	$1.83\ \cancel{g\ CO_2} \times \frac{1\ mol\ CO_2}{44.01\ \cancel{g\ CO_2}} = 0.0416\ mol\ CO_2$ $0.901\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 0.0500\ mol\ H_2O$	$2.445\ \cancel{g\ CO_2} \times \frac{1\ mol\ CO_2}{44.01\ \cancel{g\ CO_2}} = 0.05556\ mol\ CO_2$ $0.6003\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 0.03331\ mol\ H_2O$
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.	$0.0416\ \cancel{mol\ CO_2} \times \frac{1\ mol\ C}{1\ \cancel{mol\ CO_2}} = 0.0416\ mol\ C$ $0.0500\ \cancel{mol\ H_2O} \times \frac{2\ mol\ H}{1\ \cancel{mol\ H_2O}} = 0.100\ mol\ H$	$0.05556\ \cancel{mol\ CO_2} \times \frac{1\ mol\ C}{1\ \cancel{mol\ CO_2}} = 0.05556\ mol\ C$ $0.03331\ \cancel{mol\ H_2O} \times \frac{2\ mol\ H}{1\ \cancel{mol\ H_2O}} = 0.06662\ mol\ H$
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.	$Mass\ C = 0.05556\ \cancel{mol\ C} \times \frac{12.01\ g\ C}{\cancel{mol\ C}} = 0.6673\ g\ C$ $Mass\ H = 0.06662\ \cancel{mol\ H} \times \frac{1.008\ g\ H}{\cancel{mol\ H}} = 0.06715\ g\ H$ $Mass\ O = 0.8233\ g - (0.6673\ g + 0.06715\ g) = 0.0889\ g$ $Mol\ O = 0.0889\ \cancel{g\ O} \times \frac{mol\ O}{16.00\ \cancel{g\ O}} = 0.00556\ mol\ O$
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.)	$C_{\frac{0.0416}{0.0416}}H_{\frac{0.100}{0.0416}} \longrightarrow C_1H_{2.4}$	$C_{\frac{0.05556}{0.00556}}H_{\frac{0.0662}{0.00556}}O_{\frac{0.00556}{0.00556}} \longrightarrow C_{10}H_{12}O_1$

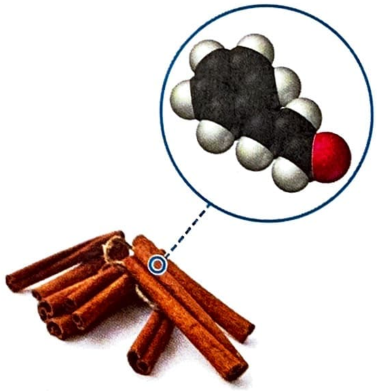
7. If the subscripts are not whole numbers, multiply all the subscripts by a small whole number to get whole-number subscripts.	$C_1H_{2.4} \times 5 \longrightarrow C_5H_{12}$ The correct empirical formula is C_5H_{12}.	The subscripts are whole numbers; no additional multiplication is needed. The correct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

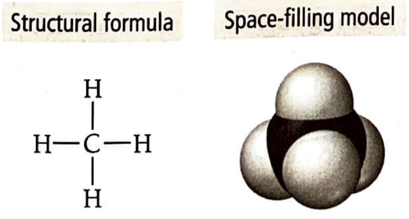
Not only did early chemists view organic and inorganic compounds as different in their origin, but also they recognized organic and inorganic compounds to be different in their properties. Organic compounds are easily decomposed. Inorganic compounds, however, are typically more difficult to decompose. 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).

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.



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

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

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 acids 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

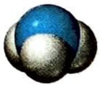
23. Determine the number of each type of atom in each formula. **MISSED THIS?** *Read Section 3.3*

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.) **MISSED THIS?** *Read Section 3.3*


a.


b.


c.

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


a.


b.


c.

27. Classify each element as atomic or molecular. **MISSED THIS?** *Read Section 3.4*

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. **MISSED THIS?** *Read Section 3.2*

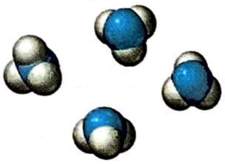
a. CO_2
b. $NiCl_2$
c. NaI
d. PCl_3

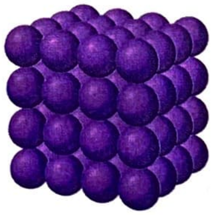
30. Classify each compound as ionic or molecular.

a. CF_2Cl_2
b. CCl_4
c. PtO_2
d. SO_3

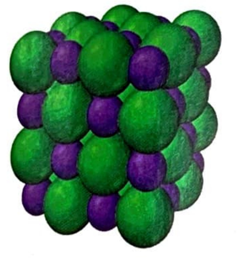
31. Based on the molecular views, classify each substance as an atomic element, a molecular element, an ionic compound, or a molecular compound. **MISSED THIS?** *Read Section 3.4*


a.

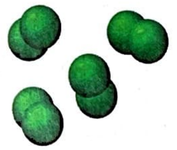

b.


c.

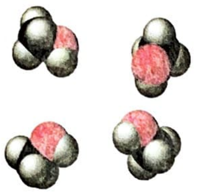
32. Based on the molecular views, classify each substance as an atomic element, a molecular element, an ionic compound, or a molecular compound.



a.



b.



c.

Formulas and Names for Ionic Compounds

33. Write a formula for the ionic compound that forms between each pair of elements.
MISSED THIS? *Read Section 3.5; Watch KCV 3.5, IWE 3.3*
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.
MISSED THIS? *Read Section 3.5; Watch KCV 3.5, IWE 3.3*
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.
MISSED THIS? *Read Section 3.5; Watch KCV 3.5, IWE 3.11*
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.
MISSED THIS? *Read Section 3.5; Watch KCV 3.5, IWE 3.11*
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.
MISSED THIS? *Read Section 3.5; Watch KCV 3.5, IWE 3.11*
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.
MISSED THIS? *Read Section 3.5; Watch KCV 3.5, IWE 3.11*
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.
MISSED THIS? *Read Section 3.5; Watch KCV 3.5, IWE 3.11*
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.
MISSED THIS? *Read Section 3.6; Watch KCV 3.6, IWE 3.11*
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.
MISSED THIS? *Read Section 3.6; Watch KCV 3.6, IWE 3.11*
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.
MISSED THIS? *Read Section 3.6; Watch IWE 3.11*
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.
MISSED THIS? *Read Section 3.6; Watch IWE 3.11*
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.
MISSED THIS? *Read Section 3.7; Watch IWE 3.11*
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.
MISSED THIS? *Read Section 3.7; Watch IWE 3.11*
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.
MISSED THIS? *Read Section 3.8*
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.
MISSED THIS? *Read Section 3.8; Watch IWE 3.13*
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.
MISSED THIS? *Read Section 3.8; Watch IWE 3.13*
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?
MISSED THIS? *Read Section 3.8; Watch IWE 3.13*
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.
MISSED THIS? *Read Section 3.8; Watch IWE 3.13*
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?
MISSED THIS? *Read Section 3.8; Watch IWE 3.13*

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.
MISSED THIS? *Read Section 3.9*
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?
MISSED THIS? *Read Section 3.9*

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.
MISSED THIS? *Read Section 3.9; Watch IWE 3.15*

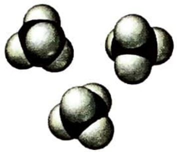
76. Silver chloride, often used in silver plating, contains 75.27% Ag by mass. Calculate the mass of silver chloride required to plate 155 mg of pure silver.

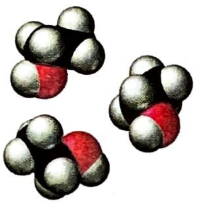
77. The iodide ion is a dietary mineral essential to good nutrition. In countries where potassium iodide is added to salt, iodine deficiency (or goiter) has been almost completely eliminated. The recommended daily allowance (RDA) for iodine is 150 μg/day. How much potassium iodide (76.45% I) should you consume if you want to meet the RDA?
MISSED THIS? *Read Section 3.9; Watch IWE 3.15*

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.) **MISSED THIS?** *Read Section 3.9*

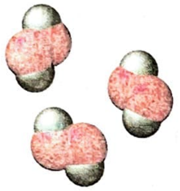

a.


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. **MISSED THIS?** *Read Section 3.9; Watch IWE 3.16*

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 the mass (in grams) of sodium in 8.5 g of each sodium-containing food additive.
MISSED THIS? *Read Section 3.9; Watch IWE 3.16*

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?
MISSED THIS? *Read Section 3.9; Watch IWE 3.16*

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.
MISSED THIS? *Read Section 3.10; Watch IWE 3.18*

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.
MISSED THIS? *Read Section 3.10; Watch IWE 3.18*

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. **MISSED THIS?** *Read Section 3.10; Watch IWE 3.18*

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.
MISSED THIS? *Read Section 3.10; Watch IWE 3.18*

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.
MISSED THIS? *Read Section 3.10*

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. **MISSED THIS?** *Read Section 3.10; Watch IWE 3.21*

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. **MISSED THIS?** *Read Section 3.10; Watch IWE 3.21*

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.

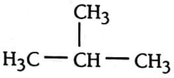
Organic Compounds

101. Classify each compound as organic or inorganic.
MISSED THIS? *Read Section 3.11*

a. $CaCO_3$
b. C_4H_8
c. $C_4H_6O_6$
d. LiF

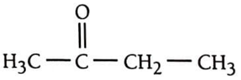
102. Classify each compound as organic or inorganic.
a. C_8H_{18}
b. CH_3NH_2
c. CaO
d. $FeCO_3$

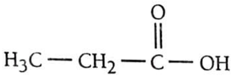
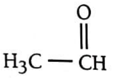
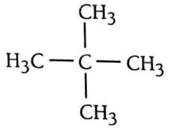
103. Classify each hydrocarbon as an alkane, alkene, or alkyne.
MISSED THIS? *Read Section 3.11*
a. $H_2C{=}CH-CH_3$
b. $H_3C-CH_2-CH_3$
c. $HC{\equiv}C-CH_3$
d. $H_3C-CH_2-CH_2-CH_3$

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

$H_3C-\overset{CH_3}{\overset{|}{CH}}-CH_3$
d. $H_3C-C{\equiv}C-CH_3$

105. Write the formula based on the name, or the name based on the formula, for each hydrocarbon.
MISSED THIS? *Read Section 3.11*
a. propane
b. $CH_3CH_2CH_3$
c. octane
d. $CH_3CH_2CH_2CH_2CH_3$

106. 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

107. Classify each organic compound as a hydrocarbon or a functionalized hydrocarbon. For functionalized hydrocarbons, identify the compound's family.
MISSED THIS? *Read Section 3.11*
a. H_3C-CH_2OH
b. H_3C-CH_3
c.

$H_3C-\overset{O}{\overset{\|}{C}}-CH_2-CH_3$
d. H_3C-NH_2

108. Classify each organic compound as a hydrocarbon or a functionalized hydrocarbon. For functionalized hydrocarbons, identify the compound's family.
a.

$H_3C-CH_2-\overset{O}{\overset{\|}{C}}-OH$
b.

$H_3C-\overset{O}{\overset{\|}{CH}}$
c.

$H_3C-\underset{CH_3}{\underset{|}{\overset{CH_3}{\overset{|}{C}}}}-CH_3$
d. $H_3C-CH_2-O-CH_3$

CUMULATIVE PROBLEMS

109. 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.

110. 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.

111. 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

112. 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

113. 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?

114. 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?

115. 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?

116. 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?

117. 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.

118. 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.

119. 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.

120. 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.

121. 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.

122. 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.

123. 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.

124. Researchers obtained the following data from experiments to 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 with 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 is 165 g/mol. Find the molar formula of benzocaine.

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

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

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

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

129. A chromium-containing compound has the formula $Fe_xCr_yO_4$ and is 28.59% oxygen by mass. Find *x* and *y*.

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

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

132. Phosphorus is obtained primarily from ores containing calcium phosphate. If a particular ore contains 57.8% calcium phosphate, what minimum mass of the ore must be processed to obtain 1.00 kg of phosphorus?

CHALLENGE PROBLEMS

133. 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?

134. 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.

135. 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.

136. 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.

137. 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.)

138. 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.)

139. 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.)

140. 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.

141. 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.

142. 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?

143. 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.

144. 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

145. 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?

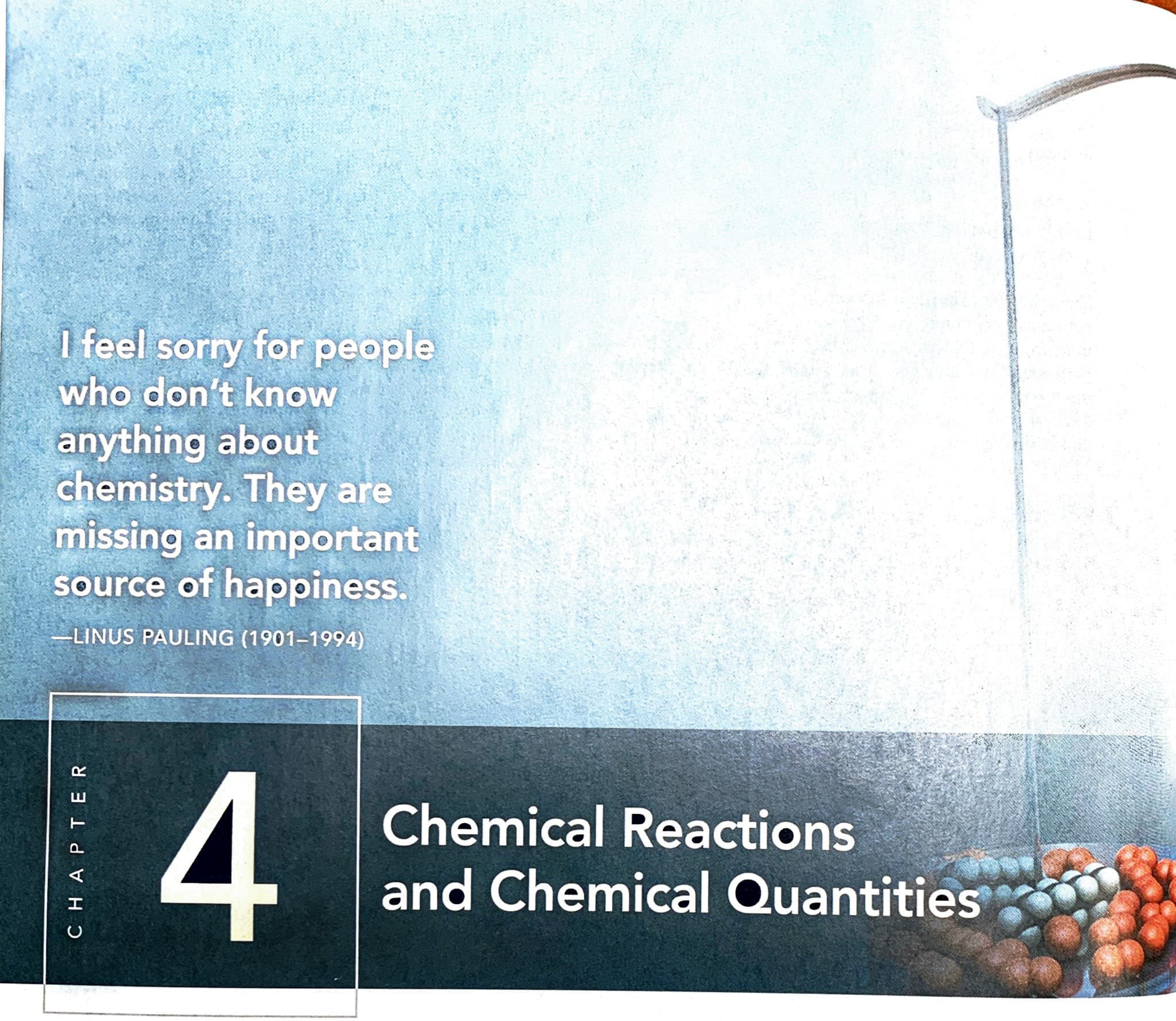
146. Without doing any calculations, determine which element in each compound has the highest mass percent composition.

a. CO
b. N_2O
c. $C_6H_{12}O_6$
d. NH_3

147. 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."

148. 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.

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



> I feel sorry for people who don't know anything about chemistry. They are missing an important source of happiness.
>
> —LINUS PAULING (1901–1994)

CHAPTER 4

Chemical Reactions and Chemical Quantities

In Chapter 3, we examined chemical compounds. We now turn to the process that can create and transform compounds: *chemical reactions*. We have seen that matter is composed of particles (atoms and molecules). When we mix certain types of particles with others, the electrons from one set of particles are attracted to the nuclei in the other set. If the conditions are right, a chemical reaction occurs and the particles are transformed. In this chapter, we learn to write chemical equations that represent these transformations. We also examine *chemical stoichiometry*—the numerical relationships between the amounts of reactants and products in chemical reactions.

The molecular models on this balance represent the reactants and products in the combustion of octane, a component of petroleum. One of the products, carbon dioxide, is the main greenhouse gas implicated in global climate change.

4.1 Climate Change and the Combustion of Fossil Fuels

The temperature outside my office today is a cool 48 °F, lower than normal for this time of year on the California coast. However, today's "chill" pales in comparison with how cold it would be without the presence of *greenhouse gases* in the atmosphere.

Greenhouse gases in the atmosphere act as a one-way filter. They allow visible light to pass through and warm Earth's surface, but they prevent heat energy from radiating back out into space.

These gases act like the glass of a greenhouse, allowing sunlight to enter the atmosphere and warm Earth's surface but preventing some of the heat generated by the

4.2 Writing and Balancing Chemical Equations

Combustion analysis (which we 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.

WATCH **NOW!**

KEY CONCEPT VIDEO 4.2
Writing and Balancing Chemical Equations

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

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

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 4.1 summarizes the common states of reactants and products and their symbols used in chemical equations.

TABLE 4.1 ■ 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.

$$CH_4(g) + \underbrace{O_2(g)}_{\text{2 O atoms}} \longrightarrow \underbrace{CO_2(g) + H_2O(g)}_{\text{2 O atoms + 1 O atom = 3 O atoms}}$$

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.

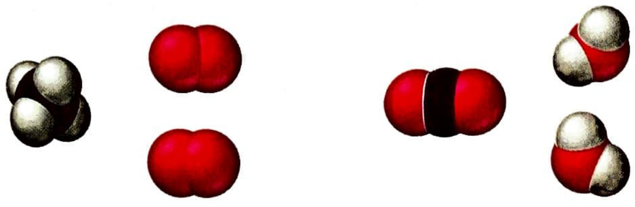
$$\underbrace{CH_4(g)}_{\text{4 H atoms}} + O_2(g) \longrightarrow CO_2(g) + \underbrace{H_2O(g)}_{\text{2 H atoms}}$$

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

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.

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

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



The equation is now balanced because the numbers of each type of atom on either side of the equation are equal. The balanced equation tells us that one CH_4 molecule reacts with two O_2 molecules to form one CO_2 molecule and two H_2O molecules. We verify that the equation is balanced by summing the number of each type of atom on 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 × $\underline{C}H_4$)	1 C atom (1 × $\underline{C}O_2$)
4 H atoms (1 × $C\underline{H_4}$)	4 H atoms (2 × $\underline{H_2}O$)
4 O atoms (2 × $\underline{O_2}$)	4 O atoms (1 × $C\underline{O_2}$ + 2 × $H_2\underline{O}$)

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

ANSWER **NOW!**

4.1 Cc Conceptual Connection

COUNTING ATOMS IN A CHEMICAL EQUATION How many oxygen atoms are on the right-hand side of the following chemical equation?

$$4\,FeCO_3(s) + O_2(g) \longrightarrow 2\,Fe_2O_3(s) + 4\,CO_2(g)$$

(a) 4 **(b)** 5 **(c)** 6 **(d)** 14

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

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 4.2

HOW TO: *Balance Chemical Equations*	**EXAMPLE 4.1** **Balancing Chemical Equations**	**EXAMPLE 4.2** **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 stoves and grills, in which it combines with gaseous oxygen to form gaseous carbon 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.)	$Co_2O_3(s) + C(s) \longrightarrow Co(s) + CO_2(g)$	$C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$

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 atoms

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 4.1

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 4.2

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.

ANSWER **NOW!**

4.2 Cc Conceptual Connection

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

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE 4.3

EXAMPLE 4.3 Balancing Chemical Equations Containing 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+}:** $SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ $1\ Sr^{2+}\text{ ion} \longrightarrow 3\ Sr^{2+}\text{ ions}$ To balance Sr^{2+}, put a 3 before $SrCl_2(aq)$. $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ $3\ Sr^{2+}\text{ ions} \longrightarrow 3\ Sr^{2+}\text{ ions}$ **Balance Li^+:** $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ $3\ Li^+\text{ ions} \longrightarrow 1\ Li^+\text{ ion}$ To balance Li^+, put a 3 before $LiCl(aq)$. $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ $3\ Li^+\text{ ions} \longrightarrow 3\ Li^+\text{ ions}$
3. Balance nonmetal ions (anions) second. If a polyatomic anion exists on both sides of the equation, balance it as a unit.	**Balance $PO_4^{\ 3-}$:** $3\ SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ $1\ PO_4^{\ 3-}\text{ ion} \longrightarrow 2\ PO_4^{\ 3-}\text{ ions}$ To balance $PO_4^{\ 3-}$, put a 2 before $Li_3PO_4(aq)$. $3\ SrCl_2(aq) + 2\ Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ $2\ PO_4^{\ 3-}\text{ ions} \longrightarrow 2\ PO_4^{\ 3-}\text{ ions}$ **Balance Cl^-:** $3\ SrCl_2(aq) + 2\ Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\ LiCl(aq)$ $6\ Cl^-\text{ ions} \longrightarrow 3\ Cl^-\text{ ions}$ 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. $3\ SrCl_2(aq) + 2\ Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 6\ LiCl(aq)$ $6\ Cl^-\text{ ions} \longrightarrow 6\ Cl^-\text{ ions}$

4. 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 4.3 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.

4.3 Reaction Stoichiometry: How Much Carbon Dioxide?

WATCH **NOW!**

KEY CONCEPT VIDEO 4.3
Reaction Stoichiometry

The balanced chemical equations for fossil fuel combustion reactions provide the exact relationships between the amount of fossil fuel burned and the amount of carbon dioxide emitted. In this discussion, we use octane (a component of gasoline) as a representative fossil fuel. The balanced equation for the combustion of octane is:

$$2\,C_8H_{18}(l) + 25\,O_2(g) \longrightarrow 16\,CO_2(g) + 18\,H_2O(g)$$

The balanced equation shows that 16 CO_2 molecules are produced for every 2 molecules of octane burned. We can extend this numerical relationship between molecules to the amounts in moles as follows:

> **The coefficients in a chemical equation specify the relative amounts in moles of each of the substances involved in the reaction.**

In other words, from the equation, we know that 16 *moles* of CO_2 are produced for every 2 *moles* of octane burned. The numerical relationships between chemical amounts in a balanced chemical equation are called reaction **stoichiometry**. Stoichiometry allows us to predict the amounts of products that will form in a chemical reaction based on the amounts of reactants that react. Stoichiometry also allows us to determine the amounts of reactants necessary to form a given amount of product. These calculations are central to chemistry, allowing chemists to plan and carry out chemical reactions to obtain products in the desired quantities.

Stoichiometry is pronounced stoy-kee-AHM-e-tree.

Making Pizza: The Relationships among Ingredients

The concepts of stoichiometry are similar to a cooking recipe. Calculating the amount of carbon dioxide produced by the combustion of a given amount of a fossil fuel is analogous to calculating the number of pizzas that can be made from a given amount of cheese. For example, suppose we use the following pizza recipe:

$$1\text{ crust} + 5\text{ ounces tomato sauce} + 2\text{ cups cheese} \longrightarrow 1\text{ pizza}$$

The recipe contains the numerical relationships between the pizza ingredients. It says that if we have two cups of cheese—and enough of everything else—we can make one pizza. We can write this relationship as a ratio between the cheese and the pizza:

$$2\text{ cups cheese} : 1\text{ pizza}$$

What if we have six cups of cheese? Assuming that we have enough of everything else, we can use the ratio as a conversion factor to calculate the number of pizzas:

$$6\,\cancel{\text{cups cheese}} \times \frac{1\text{ pizza}}{2\,\cancel{\text{cups cheese}}} = 3\text{ pizzas}$$

Six cups of cheese are sufficient to make three pizzas. The pizza recipe contains numerical ratios between other ingredients as well, including the crust and the tomato sauce:

1 crust : 1 pizza

5 ounces tomato sauce : 1 pizza

Making Molecules: Mole-to-Mole Conversions

In a balanced chemical equation, we have a "recipe" for how reactants combine to form products. From our balanced equation for the combustion of octane, for example, we can write the following stoichiometric ratio:

$$2 \text{ mol } C_8H_{18} : 16 \text{ mol } CO_2$$

We can use this ratio to determine how many moles of CO_2 form when a given number of moles of C_8H_{18} burns. Suppose that we burn 22.0 moles of C_8H_{18}; how many moles of CO_2 form? We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe. The ratio acts as a conversion factor between the amount in moles of the reactant (C_8H_{18}) and the amount in moles of the product (CO_2):

$$22.0 \, \cancel{\text{mol } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2 \, \cancel{\text{mol } C_8H_{18}}} = 176 \text{ mol } CO_2$$

The combustion of 22.0 moles of C_8H_{18} adds 176 moles of CO_2 to the atmosphere.

ANSWER NOW!

4.3 Cc Conceptual Connection

STOICHIOMETRY I Use the balanced equation for the combustion of octane to determine how many moles of H_2O are produced by the combustion of 22.0 moles of C_8H_{18}.

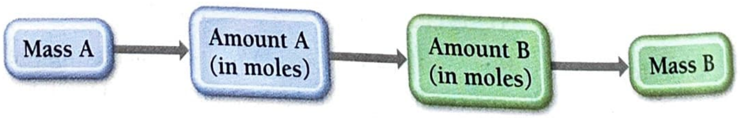
$$2\,C_8H_{18}(l) + 25\,O_2(g) \longrightarrow 16\,CO_2(g) + 18\,H_2O(g)$$

(a) 18 moles H_2O **(b)** 22 moles H_2O

(c) 176 moles H_2O **(d)** 198 moles H_2O

Making Molecules: Mass-to-Mass Conversions

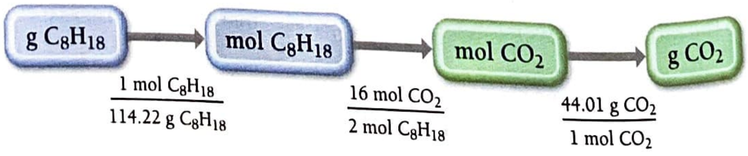
According to the U.S. Department of Energy, the world burned 3.6×10^{10} barrels of petroleum in 2017, the equivalent of approximately 4.0×10^{15} g of gasoline. We can estimate the mass of CO_2 emitted into the atmosphere from burning this much gasoline using the combustion of 4.0×10^{15} g octane as the representative reaction. This calculation is similar to the one we just did, except that we are now given the *mass* of octane instead of the *amount* of octane in moles. Consequently, we must first convert the mass (in grams) to the amount (in moles). The general conceptual plan for calculations in which we are given the mass of a reactant or product in a chemical reaction and asked to find the mass of a different reactant or product takes the form:



where A and B are two different substances involved in the reaction.

We use the molar mass of A to convert from the mass of A to the amount of A (in moles). We use the appropriate ratio from the balanced chemical equation to convert from the amount of A (in moles) to the amount of B (in moles). And finally, we use the molar mass of B to convert from the amount of B (in moles) to the mass of B. To calculate the mass of CO_2 emitted upon the combustion of 4.0×10^{15} g of octane, we use the following conceptual plan:

Conceptual Plan



Relationships used

2 mol C_8H_{18} : 16 mol CO_2 (from the chemical equation)
molar mass $C_8H_{18} = 114.22$ g/mol
molar mass $CO_2 = 44.01$ g/mol

Solution

We follow the conceptual plan to solve the problem, beginning with g C_8H_{18} and canceling units to arrive at g CO_2:

$$4.0 \times 10^{15}\ \cancel{\text{g } C_8H_{18}} \times \frac{1\ \cancel{\text{mol } C_8H_{18}}}{114.22\ \cancel{\text{g } C_8H_{18}}} \times \frac{16\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_8H_{18}}} \times \frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 1.2 \times 10^{16}\text{ g } CO_2$$

The world's petroleum combustion produces 1.2×10^{16} g CO_2 (1.2×10^{13} kg) per year. In comparison, volcanoes produce about 2×10^{11} kg CO_2 per year.* In other words, volcanoes emit only $\frac{2 \times 10^{11}\text{ kg}}{1.2 \times 10^{13}\text{ kg}} \times 100\% = 1.7\%$ as much CO_2 per year as petroleum combustion. The argument that volcanoes emit more carbon dioxide than fossil fuel combustion is clearly mistaken. Examples 4.4 and 4.5 provide additional practice with stoichiometric calculations.

The percentage of CO_2 emitted by volcanoes relative to all fossil fuels is even less than 1.7% because the combustion of coal and natural gas also emits CO_2.

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE 4.4

EXAMPLE 4.4 Stoichiometry

During photosynthesis, plants convert carbon dioxide and water into glucose ($C_6H_{12}O_6$) according to the reaction:

$$6\,CO_2(g) + 6\,H_2O(l) \xrightarrow{\text{sunlight}} 6\,O_2(g) + C_6H_{12}O_6(aq)$$

Suppose that a particular plant consumes 37.8 g of CO_2 in one week. Assuming that there is more than enough water present to react with all of the CO_2, what mass of glucose (in grams) can the plant synthesize from the CO_2?

SORT The problem provides the mass of carbon dioxide and asks you to find the mass of glucose that can be produced.

GIVEN: 37.8 g CO_2
FIND: g $C_6H_{12}O_6$

STRATEGIZE The conceptual plan follows the general pattern of mass A → amount A (in moles) → amount B (in moles) → mass B. From the chemical equation, deduce the relationship between moles of carbon dioxide and moles of glucose. Use the molar masses to convert between grams and moles.

CONCEPTUAL PLAN

$$\boxed{\text{g } CO_2} \xrightarrow[\frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2}]{} \boxed{\text{mol } CO_2} \xrightarrow[\frac{1\text{ mol } C_6H_{12}O_6}{6\text{ mol } CO_2}]{} \boxed{\text{mol } C_6H_{12}O_6} \xrightarrow[\frac{180.2\text{ g } C_6H_{12}O_6}{1\text{ mol } C_6H_{12}O_6}]{} \boxed{\text{g } C_6H_{12}O_6}$$

RELATIONSHIPS USED

molar mass $CO_2 = 44.01$ g/mol
6 mol CO_2 : 1 mol $C_6H_{12}O_6$
molar mass $C_6H_{12}O_6 = 180.2$ g/mol

SOLVE Follow the conceptual plan to solve the problem. Begin with g CO_2 and use the conversion factors to arrive at g $C_6H_{12}O_6$.

SOLUTION

$$37.8\ \cancel{\text{g } CO_2} \times \frac{1\ \cancel{\text{mol } CO_2}}{44.01\ \cancel{\text{g } CO_2}} \times \frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{6\ \cancel{\text{mol } CO_2}} \times \frac{180.2\text{ g } C_6H_{12}O_6}{1\ \cancel{\text{mol } C_6H_{12}O_6}} = 25.8\text{ g } C_6H_{12}O_6$$

—*Continued on the next page*

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*Gerlach, T. M., Present-day CO_2 emissions from volcanoes. *Eos, Transactions, American Geophysical Union* 72(1991): 249, 254–255.

Continued—

CHECK The units of the answer are correct. The magnitude of the answer (25.8 g) is less than the initial mass of CO_2 (37.8 g). This is reasonable because each carbon in CO_2 has two oxygen atoms associated with it, while in $C_6H_{12}O_6$ each carbon has only one oxygen atom associated with it and two hydrogen atoms, which are much lighter than oxygen. Therefore, the mass of glucose produced should be less than the mass of carbon dioxide for this reaction.

FOR PRACTICE 4.4 Magnesium hydroxide, the active ingredient in milk of magnesia, neutralizes stomach acid, primarily HCl, according to the reaction:

$$\text{Mg(OH)}_2(s) + 2\,\text{HCl}(aq) \longrightarrow 2\,\text{H}_2\text{O}(l) + \text{MgCl}_2(aq)$$

What mass of HCl, in grams, is neutralized by a dose of milk of magnesia containing 3.26 g Mg(OH)_2?

EXAMPLE 4.5 Stoichiometry

Sulfuric acid (H_2SO_4) is a component of acid rain that forms when SO_2, a pollutant, reacts with oxygen and water according to the simplified reaction:

$$2\,\text{SO}_2(g) + \text{O}_2(g) + 2\,\text{H}_2\text{O}(l) \longrightarrow 2\,\text{H}_2\text{SO}_4(aq)$$

The generation of the electricity used by a medium-sized home produces about 25 kg of SO_2 per year. Assuming that there is more than enough O_2 and H_2O, what mass of H_2SO_4, in kg, can form from this much SO_2?

SORT The problem gives the mass of sulfur dioxide and asks you to find the mass of sulfuric acid.

GIVEN: 25 kg SO_2

FIND: kg H_2SO_4

STRATEGIZE The conceptual plan follows the standard format of mass → amount (in moles) → amount (in moles) → mass. Since the original quantity of SO_2 is given in kilograms, you must first convert to grams. You can deduce the relationship between moles of sulfur dioxide and moles of sulfuric acid from the chemical equation. Since the final quantity is requested in kilograms, convert to kilograms at the end.

CONCEPTUAL PLAN

$$\boxed{\text{kg SO}_2} \xrightarrow[\frac{1000\text{ g}}{1\text{ kg}}]{} \boxed{\text{g SO}_2} \xrightarrow[\frac{1\text{ mol SO}_2}{64.07\text{ g SO}_2}]{} \boxed{\text{mol SO}_2} \xrightarrow[\frac{2\text{ mol H}_2\text{SO}_4}{2\text{ mol SO}_2}]{}$$

$$\boxed{\text{mol H}_2\text{SO}_4} \xrightarrow[\frac{98.09\text{ g H}_2\text{SO}_4}{1\text{ mol H}_2\text{SO}_4}]{} \boxed{\text{g H}_2\text{SO}_4} \xrightarrow[\frac{1\text{ kg}}{1000\text{ g}}]{} \boxed{\text{kg H}_2\text{SO}_4}$$

RELATIONSHIPS USED

1 kg = 1000 g

molar mass SO_2 = 64.07 g/mol

2 mol SO_2 : 2 mol H_2SO_4

molar mass H_2SO_4 = 98.09 g/mol

SOLVE Follow the conceptual plan to solve the problem. Begin with the given amount of SO_2 in kilograms and use the conversion factors to arrive at kg H_2SO_4.

SOLUTION

$$25\,\cancel{\text{kg}}\,\cancel{\text{SO}_2} \times \frac{1000\,\cancel{\text{g}}}{1\,\cancel{\text{kg}}} \times \frac{1\,\cancel{\text{mol SO}_2}}{64.07\,\cancel{\text{g SO}_2}} \times \frac{2\,\cancel{\text{mol H}_2\text{SO}_4}}{2\,\cancel{\text{mol SO}_2}}$$

$$\times \frac{98.09\,\cancel{\text{g}}\,\text{H}_2\text{SO}_4}{1\,\cancel{\text{mol H}_2\text{SO}_4}} \times \frac{1\text{ kg}}{1000\,\cancel{\text{g}}} = 38\text{ kg H}_2\text{SO}_4$$

CHECK The units of the final answer are correct. The magnitude of the final answer (38 kg H_2SO_4) is larger than the amount of SO_2 given (25 kg). This is reasonable because in the reaction each SO_2 molecule "gains weight" by reacting with O_2 and H_2O.

FOR PRACTICE 4.5 Another component of acid rain is nitric acid, which forms when NO_2, also a pollutant, reacts with oxygen and water according to the simplified equation:

$$4\,\text{NO}_2(g) + \text{O}_2(g) + 2\,\text{H}_2\text{O}(l) \longrightarrow 4\,\text{HNO}_3(aq)$$

The generation of the electricity used by a medium-sized home produces about 16 kg of NO_2 per year. Assuming that there is adequate O_2 and H_2O, what mass of HNO_3, in kg, can form from this amount of NO_2 pollutant?

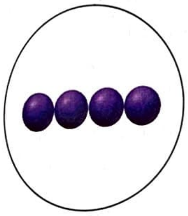
STOICHIOMETRY II Under certain conditions, sodium reacts with oxygen to form sodium oxide according to the reaction:

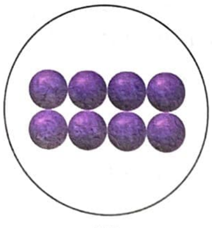
$$4\,\text{Na}(s) + \text{O}_2(g) \longrightarrow 2\,\text{Na}_2\text{O}(s)$$

A flask contains the amount of oxygen represented by the diagram shown at far right. Which of the following images best represents the amount of sodium required to completely react with all of the oxygen in the flask according to the equation?

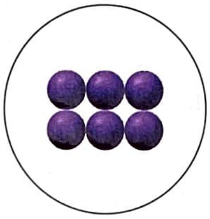
4.4
Cc
Conceptual Connection

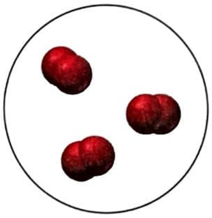
ANSWER **NOW!**


(a)


(b)


(c)

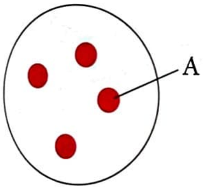

(d)



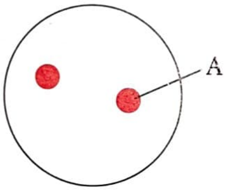
STOICHIOMETRY III Consider the generic chemical equation A + 3B → 2C. Let circles represent molecules of A and squares represent molecules of B. The diagram shown at the far right represents the amount of B available for reaction. Which diagram in the answer options accurately represents the amount of A necessary to completely react with B?

4.5
Cc
Conceptual Connection

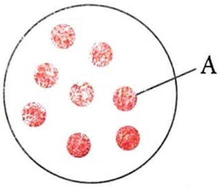
ANSWER **NOW!**



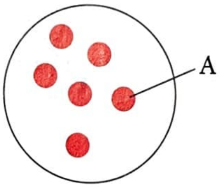
(a)



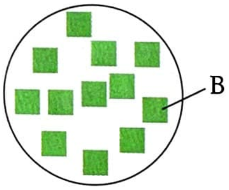
(b)



(c)



(d)



4.4 Stoichiometric Relationships: Limiting Reactant, Theoretical Yield, Percent Yield, and Reactant in Excess

WATCH **NOW!**

KEY CONCEPT VIDEO 4.4
Limiting Reactant, Theoretical Yield, and Percent Yield

Let's return to our pizza analogy to understand three more important concepts in reaction stoichiometry: *limiting reactant*, *theoretical yield*, and *percent yield*. Recall our pizza recipe from Section 4.3:

$$1\text{ crust} + 5\text{ ounces tomato sauce} + 2\text{ cups cheese} \longrightarrow 1\text{ pizza}$$

Suppose that we have four crusts, ten cups of cheese, and 15 ounces of tomato sauce. How many pizzas can we make?

We have enough crusts to make:

$$4\,\cancel{\text{crusts}} \times \frac{1\text{ pizza}}{1\,\cancel{\text{crust}}} = 4\text{ pizzas}$$

We have enough cheese to make:

$$10\,\cancel{\text{cups cheese}} \times \frac{1\text{ pizza}}{2\,\cancel{\text{cups cheese}}} = 5\text{ pizzas}$$

We have enough tomato sauce to make:

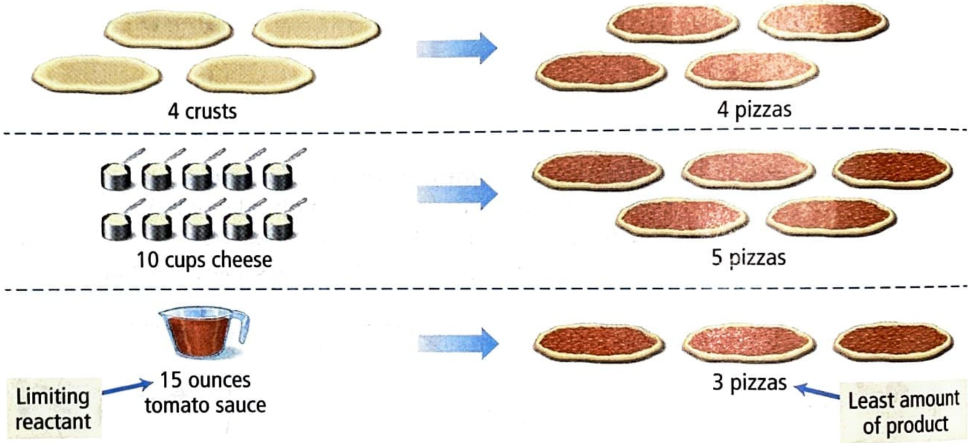
$$15\ \cancel{\text{ounces tomato sauce}} \times \frac{1\ \text{pizza}}{5\ \cancel{\text{ounces tomato sauce}}} = 3\ \text{pizzas}$$

Limiting reactant

Smallest number of pizzas

The term *limiting reagent* is sometimes used in place of *limiting reactant*.

We have enough crusts for four pizzas and enough cheese for five pizzas, but enough tomato sauce for only three pizzas. Consequently, unless we get more ingredients, we can make only three pizzas. The tomato sauce *limits* how many pizzas we can make. If the pizza recipe were a chemical reaction, the tomato sauce would be the **limiting reactant**, the reactant that limits the amount of product in a chemical reaction. Notice that the limiting reactant is the reactant that makes *the least amount of product*.



The ingredient that makes the least amount of pizza determines how many pizzas you can make.

The reactants that *do not* limit the amount of product—such as the crusts and the cheese in this example—are said to be **in excess**. If this were a chemical reaction, three pizzas would be the **theoretical yield**, the maximum amount of product that can be made in a chemical reaction based on the amount of limiting reactant.

Let us carry this analogy one step further. Suppose we go on to cook our pizzas and accidentally burn one of them. Even though we theoretically have enough ingredients for three pizzas, we end up with only two. If this were a chemical reaction, the two pizzas would be our **actual yield**, the amount of product actually produced by a chemical reaction. (The actual yield is always equal to or less than the theoretical yield because a small amount of product is usually lost to other reactions or does not form during a reaction.) Finally, our **percent yield**, the percentage of the theoretical yield that was actually attained, is calculated as the ratio of the actual yield to the theoretical yield:

Actual yield

$$\%\ \text{yield} = \frac{2\ \cancel{\text{pizzas}}}{3\ \cancel{\text{pizzas}}} \times 100\% = 67\%$$

Theoretical yield

Since one of our pizzas burned, our percent yield for pizzas is 67%.

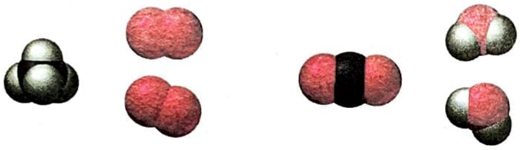
Summarizing Limiting Reactant and Yield

- **The limiting reactant** (or **limiting reagent**) is the reactant that is completely consumed in a chemical reaction and limits the amount of product.
- **The reactant in excess** is any reactant that occurs in a quantity greater than is required to completely react with the limiting reactant.
- **The theoretical yield** is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
- **The actual yield** is the amount of product actually produced by a chemical reaction.
- **The percent yield** is calculated as $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$.

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield

We can apply these concepts to a chemical reaction. Recall from Section 4.2 our balanced equation for the combustion of methane:

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

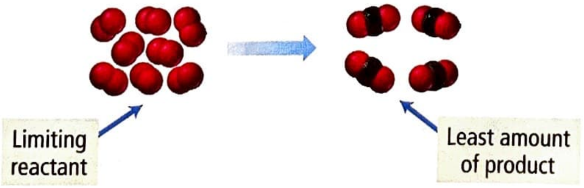


If we start out with five CH_4 molecules and eight O_2 molecules, what is our limiting reactant? What is our theoretical yield of carbon dioxide molecules? First, we calculate the number of CO_2 molecules that can be made from five CH_4 molecules:

$$5\,\cancel{CH_4} \times \frac{1\,CO_2}{1\,\cancel{CH_4}} = 5\,CO_2$$

We then calculate the number of CO_2 molecules that can be made from eight O_2 molecules:

$$8\,\cancel{O_2} \times \frac{1\,CO_2}{2\,\cancel{O_2}} = 4\,CO_2$$



We have enough CH_4 to make five CO_2 molecules and enough O_2 to make four CO_2 molecules; therefore, O_2 is the limiting reactant, and four CO_2 molecules is the theoretical yield. The CH_4 is in excess.

An alternative way to calculate the limiting reactant (which we describe here but do not use again in this book) is to pick any reactant and determine how much of the *other reactant* is necessary to completely react with it. For the reaction we just examined, we have five CH_4 molecules and eight O_2 molecules. Let's pick the five CH_4 molecules and determine how many O_2 molecules are necessary to completely react with them:

$$5\,\cancel{CH_4} \times \frac{2\,O_2}{1\,\cancel{CH_4}} = 10\,O_2$$

Since we need ten O_2 molecules to completely react with the five CH_4 molecules, and since we have only eight O_2 molecules, we know that the O_2 is the limiting reactant. The same method can be applied by comparing the amounts of reactants in moles.

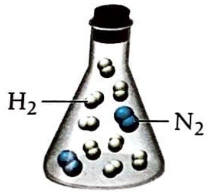
ANSWER NOW!

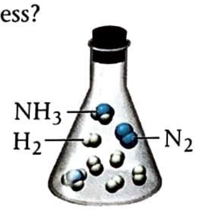
4.6 Cc Conceptual Connection

LIMITING REACTANT AND THEORETICAL YIELD Nitrogen and hydrogen gas react to form ammonia according to the reaction:

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$$

A flask contains a mixture of reactants represented by the image shown at the left. Which of the following images best represents the mixture in the flask after the reactants have reacted as completely as possible? What is the limiting reactant? Which reactant is in excess?





(a)

(b)

(c)

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Reactant Masses

When working in the laboratory, we normally measure the initial quantities of reactants in grams, not in number of molecules. To find the limiting reactant and theoretical yield from initial masses, we must first convert the masses to amounts in moles. Consider the reaction:

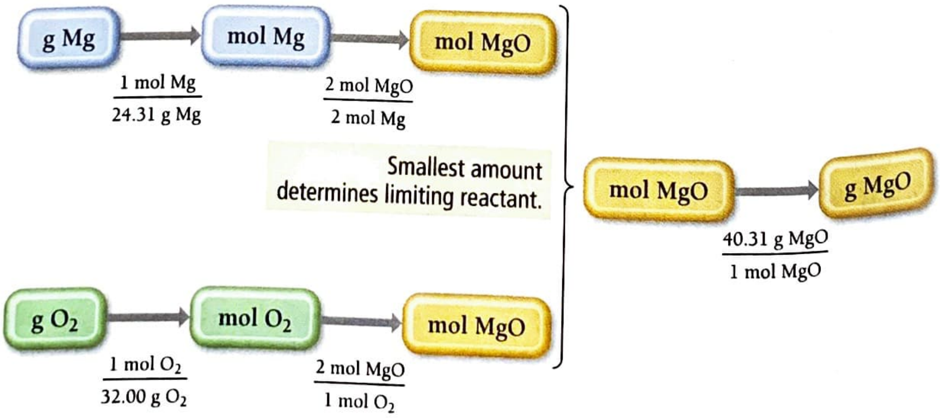
$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

A reaction mixture contains 42.5 g Mg and 33.8 g O_2; what is the limiting reactant and theoretical yield?

To solve this problem, we must determine which of the reactants makes the least amount of product.

Conceptual Plan

We can find the limiting reactant by calculating how much product can be made from each reactant. However, we are given the initial quantities in grams, and stoichiometric relationships are between moles, so we must first convert to moles. We then convert from moles of the reactant to moles of product. The reactant that makes the *least amount of product* is the limiting reactant. The conceptual plan is:



In this conceptual plan, we compare the number of moles of MgO made by each reactant and convert only the smaller amount to grams. (Alternatively, we can convert both quantities to grams and determine the limiting reactant based on the mass of the product.)

Relationships Used

molar mass Mg = 24.31 g/mol

molar mass O_2 = 32.00 g/mol

2 mol Mg : 2 mol MgO
1 mol O_2 : 2 mol MgO
molar mass MgO = 40.31 g/mol

Solution

Beginning with the masses of each reactant, we follow the conceptual plan to calculate how much product can be made from each:

$$42.5\,\cancel{\text{g Mg}} \times \frac{1\,\cancel{\text{mol Mg}}}{24.31\,\cancel{\text{g Mg}}} \times \frac{2\text{ mol MgO}}{2\,\cancel{\text{mol Mg}}} = 1.7\underline{4}83\text{ mol MgO}$$

Limiting reactant

Least amount of product

$$1.7\underline{4}83\,\cancel{\text{mol MgO}} \times \frac{40.31\text{ g MgO}}{1\,\cancel{\text{mol MgO}}} = 70.5\text{ g MgO}$$

$$33.8\,\cancel{\text{g }O_2} \times \frac{1\,\cancel{\text{mol }O_2}}{32.00\,\cancel{\text{g }O_2}} \times \frac{2\text{ mol MgO}}{1\,\cancel{\text{mol }O_2}} = 2.1\underline{1}25\text{ mol MgO}$$

Since Mg makes the least amount of product, it is the limiting reactant, and O_2 is in excess. Notice that the limiting reactant is not necessarily the reactant with the least mass. In this case, the mass of O_2 is less than the mass of Mg, yet Mg is the limiting reactant because it makes the least amount of MgO. The theoretical yield is 70.5 g of MgO, the mass of product possible based on the limiting reactant.

Suppose that after the synthesis, the actual yield of MgO is 55.9 g. What is the percent yield? We calculate the percent yield as follows:

$$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{55.9\,\cancel{\text{g}}}{70.5\,\cancel{\text{g}}} \times 100\% = 79.3\%$$

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 4.6

EXAMPLE 4.6 Limiting Reactant and Theoretical Yield

Ammonia, NH_3, can be synthesized by the reaction:

$$2\,NO(g) + 5\,H_2(g) \longrightarrow 2\,NH_3(g) + 2\,H_2O(g)$$

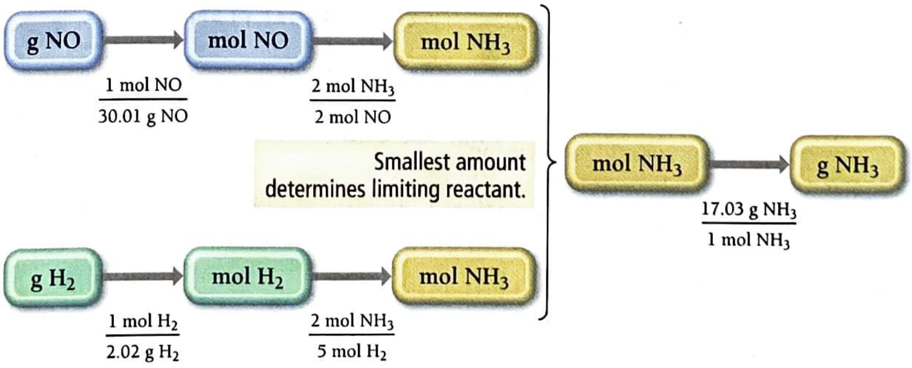
Starting with 86.3 g NO and 25.6 g H_2, find the theoretical yield of ammonia in grams.

SORT You are given the mass of each reactant in grams and asked to find the theoretical yield of a product.

GIVEN: 86.3 g NO, 25.6 g H_2
FIND: theoretical yield of $NH_3(g)$

STRATEGIZE Determine which reactant makes the least amount of product by converting from grams of each reactant to moles of the reactant to moles of the product. Use molar masses to convert between grams and moles and use the stoichiometric relationships (from the chemical equation) to convert between moles of reactant and moles of product. Remember that the reactant that makes *the least amount of product* is the limiting reactant. Convert the number of moles of product obtained using the limiting reactant to grams of product.

CONCEPTUAL PLAN



RELATIONSHIPS USED

molar mass NO = 30.01 g/mol
molar mass H_2 = 2.02 g/mol
2 mol NO : 2 mol NH_3 (from chemical equation)
5 mol H_2 : 2 mol NH_3 (from chemical equation)
molar mass NH_3 = 17.03 g/mol

—*Continued on the next page*

Continued—

SOLVE Beginning with the given mass of each reactant, calculate the amount of product that can be made in moles. Convert the amount of product made by the limiting reactant to grams—this is the theoretical yield.

SOLUTION

$$86.3\ \cancel{\text{g NO}} \times \frac{1\ \cancel{\text{mol NO}}}{30.01\ \cancel{\text{g NO}}} \times \frac{2\ \text{mol NH}_3}{2\ \cancel{\text{mol NO}}} = 2.8\underline{7}57\ \text{mol NH}_3$$

Limiting reactant

Least amount of product

$$25.6\ \cancel{\text{g H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{2.02\ \cancel{\text{g H}_2}} \times \frac{2\ \text{mol NH}_3}{5\ \cancel{\text{mol H}_2}} = 5.0\underline{6}93\ \text{mol NH}_3$$

$$2.8\underline{7}57\ \cancel{\text{mol NH}_3} \times \frac{17.03\ \text{g NH}_3}{\cancel{\text{mol NH}_3}} = 49.0\ \text{g NH}_3$$

Since NO makes the least amount of product, it is the limiting reactant, and the theoretical yield of ammonia is 49.0 g.

CHECK The units of the answer (g NH_3) are correct. The magnitude (49.0 g) seems reasonable given that 86.3 g NO is the limiting reactant. NO contains one oxygen atom per nitrogen atom, and NH_3 contains three hydrogen atoms per nitrogen atom. Since three hydrogen atoms have less mass than one oxygen atom, it is reasonable that the mass of NH_3 obtained is less than the mass of NO.

FOR PRACTICE 4.6 Ammonia can also be synthesized by the reaction:

$$3\ H_2(g) + N_2(g) \longrightarrow 2\ NH_3(g)$$

What is the theoretical yield of ammonia, in kg, that we can synthesize from 5.22 kg of H_2 and 31.5 kg of N_2?

EXAMPLE 4.7 Limiting Reactant and Theoretical Yield

We can obtain titanium metal from its oxide according to the following balanced equation:

$$TiO_2(s) + 2\ C(s) \longrightarrow Ti(s) + 2\ CO(g)$$

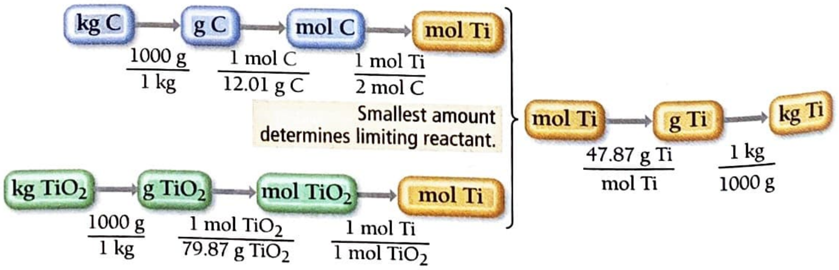
When 28.6 kg of C reacts with 88.2 kg of TiO_2, 42.8 kg of Ti is produced. Find the limiting reactant, theoretical yield (in kg), and percent yield.

SORT You are given the mass of each reactant and the mass of product formed. You are asked to find the limiting reactant, theoretical yield, and percent yield.

GIVEN: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti produced
FIND: limiting reactant, theoretical yield, % yield

STRATEGIZE Determine which of the reactants makes the least amount of product by converting from kilograms of each reactant to moles of product. Convert between grams and moles using molar mass. Convert between moles of reactant and moles of product using the stoichiometric relationships derived from the chemical equation. Remember that the reactant that makes the *least amount of product* is the limiting reactant.

CONCEPTUAL PLAN



Determine the theoretical yield (in kilograms) by converting the number of moles of product obtained with the limiting reactant to kilograms of product.

RELATIONSHIPS USED

1000 g = 1 kg

molar mass of C = 12.01 g/mol

molar mass of TiO_2 = 79.87 g/mol

1 mol TiO_2 : 1 mol Ti

2 mol C : 1 mol Ti

molar mass of Ti = 47.87 g/mol

SOLVE Beginning with the actual amount of each reactant, calculate the amount of product that can be made in moles. Convert the amount of product made by the limiting reactant to kilograms—this is the theoretical yield.

SOLUTION

$$28.6\ \cancel{\text{kg C}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \cancel{\text{mol C}}}{12.01\ \cancel{\text{g C}}} \times \frac{1\ \text{mol Ti}}{2\ \cancel{\text{mol C}}} = \mathbf{1.1\underline{9}07 \times 10^3\ mol\ Ti}$$

Limiting reactant

Least amount of product

$$88.2\ \cancel{\text{kg TiO}_2} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \cancel{\text{mol TiO}_2}}{79.87\ \cancel{\text{g TiO}_2}} \times \frac{1\ \text{mol Ti}}{1\ \cancel{\text{mol TiO}_2}} = \mathbf{1.1\underline{0}43 \times 10^3\ mol\ Ti}$$

$$1.10\underline{4}3 \times 10^3\ \cancel{\text{mol Ti}} \times \frac{47.87\ \cancel{\text{g Ti}}}{1\ \cancel{\text{mol Ti}}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = \mathbf{52.9\ kg\ Ti}$$

Calculate the percent yield by dividing the actual yield (42.8 kg Ti) by the theoretical yield.

Since TiO_2 makes the least amount of product, it is the limiting reactant, and 52.9 kg Ti is the theoretical yield.

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{42.8\ \text{kg}}{52.9\ \text{kg}} \times 100\% = 80.9\%$$

CHECK The theoretical yield has the correct units (kg Ti) and has a reasonable magnitude compared to the mass of TiO_2. Since Ti has a lower molar mass than TiO_2, the amount of Ti made from TiO_2 should have a lower mass. The percent yield is reasonable (under 100% as it should be).

FOR PRACTICE 4.7 Mining companies use this reaction to obtain iron from iron ore:

$$Fe_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Fe(s) + 3\,CO_2(g)$$

The reaction of 167 g Fe_2O_3 with 85.8 g CO produces 72.3 g Fe. Determine the limiting reactant, theoretical yield, and percent yield.

REACTANT IN EXCESS

4.7 Cc Conceptual Connection ANSWER **NOW!**

Nitrogen dioxide reacts with water to form nitric acid and nitrogen monoxide according to the equation:

$$3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(l) + NO(g)$$

Suppose that 5 mol NO_2 and 1 mol H_2O combine and react completely. How many moles of the reactant in excess are present after the reaction has completed?

(a) 1 mol NO_2 **(b)** 1 mol H_2O **(c)** 2 mol NO_2 **(d)** 2 mol H_2O

4.5 Three Examples of Chemical Reactions: Combustion, Alkali Metals, and Halogens

In this section, we examine three types of reactions. The first is combustion reactions, which we encountered in Section 4.2. The second is the reactions of the alkali metals. As we discussed in Section 2.7, alkali metals are among the most active metals. Alkali metal reactions are good examples of the types of reactions that many metals undergo. The third type of reactions involves the halogens. Halogens are among the most active nonmetals.

Combustion Reactions

A *combustion reaction* involves the reaction of a substance with O_2 to form one or more oxygen-containing compounds, often including water. Combustion reactions also emit heat. For example, as you saw earlier in this chapter, natural gas (CH_4) reacts with oxygen to form carbon dioxide and water:

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

Ethanol, the alcohol in alcoholic beverages, also reacts with oxygen in a combustion reaction to form carbon dioxide and water:

$$C_2H_5OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$$

Compounds containing carbon and hydrogen—or carbon, hydrogen, and oxygen—always form carbon dioxide and water upon complete combustion. Other combustion reactions include the reaction of carbon with oxygen to form carbon dioxide:

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

and the reaction of hydrogen with oxygen to form water:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

We can write chemical equations for most combustion reactions by noticing the pattern of reactivity. Any carbon in a combustion reaction reacts with oxygen to produce carbon dioxide, and any hydrogen reacts with oxygen to form water.

EXAMPLE 4.8 Writing Equations for Combustion Reactions

Write a balanced equation for the combustion of liquid methyl alcohol (CH_3OH).

SOLUTION

Begin by writing an unbalanced equation showing the reaction of CH_3OH with O_2 to form CO_2 and H_2O.	$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
Balance the equation using the guidelines from Section 4.2.	$2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$

FOR PRACTICE 4.8 Write a balanced equation for the complete combustion of liquid C_2H_5SH. *Hint:* The sulfur in this compound reacts to form SO_2.

Alkali Metal Reactions

The reactions of the alkali metals (group 1A in the periodic table) with nonmetals are vigorous. For example, the alkali metals (M) react with halogens (X) according to the reaction:

$$2\,M + X_2 \longrightarrow 2\,MX$$

The reaction of sodium and chlorine to form sodium chloride is typical:

$$2\,Na(s) + Cl_2(g) \longrightarrow 2\,NaCl(s)$$

This reaction emits heat and sparks as it occurs (Figure 4.4◀). Each successive alkali metal reacts even more vigorously with chlorine.

The alkali metals also react with water to form the dissolved alkali metal ion, the hydroxide ion, and hydrogen gas:

$$2\,M(s) + 2\,H_2O(l) \longrightarrow 2\,M^+(aq) + 2\,OH^-(aq) + H_2(g)$$

The reaction is highly exothermic and can be explosive because the heat from the reaction can ignite the hydrogen gas. The reaction becomes more explosive as we move down the column from one metal to the next, as shown in Figure 4.5▶.*

▲ **FIGURE 4.4 Reaction of Sodium and Chlorine to Form Sodium Chloride**

Halogen Reactions

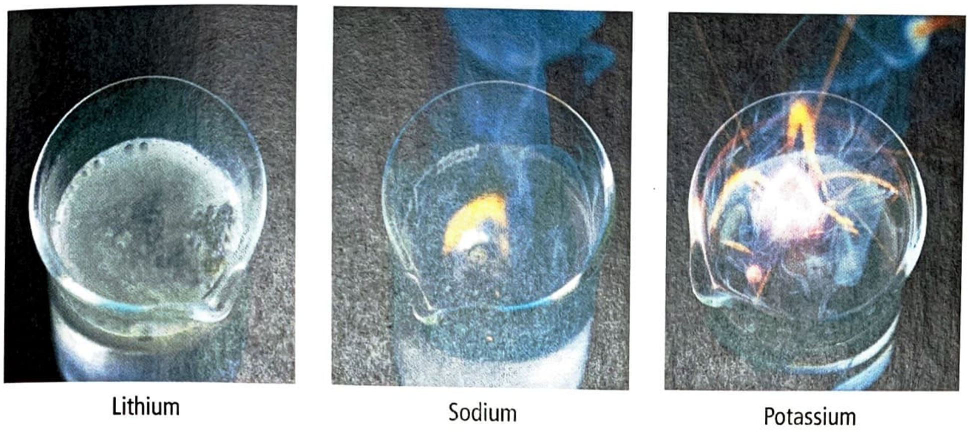
The halogens (group 7A) are among the most active nonmetals in the periodic table. The halogens all react with many metals to form *metal halides* according to the equation:

$$2\,M + n\,X_2 \longrightarrow 2\,MX_n$$

---

*The rate of the alkali metal reaction with water, and therefore its vigor, is enhanced by the successively lower melting points of the alkali metals as we move down the column. The low melting points of the heavier metals allow the emitted heat to actually melt the metal, increasing the reaction rate.

Reactions of the Alkali Metals with Water



◀ FIGURE 4.5 Reactions of the Alkali Metals with Water The reactions become progressively more vigorous as we move down the group.

where M is the metal, X is the halogen, and MX_n is the metal halide. For example, chlorine reacts with iron according to the equation:

$$2\,Fe(s) + 3\,Cl_2(g) \longrightarrow 2\,FeCl_3(s)$$

Since metals tend to lose electrons and the halogens tend to gain them, the metal halides—like all compounds that form between metals and nonmetals—contain ionic bonds.

The halogens also react with hydrogen to form *hydrogen halides* according to the equation:

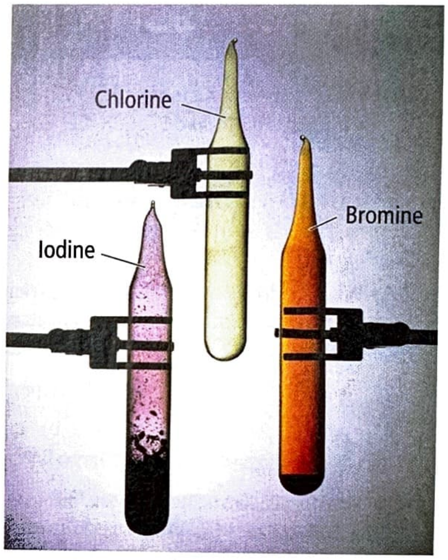
$$H_2(g) + X_2 \longrightarrow 2\,HX(g)$$

The hydrogen halides—like all compounds that form between two nonmetals—contain covalent bonds. All of the hydrogen halides form acids when combined with water.

The halogens also react with each other to form *interhalogen compounds*. For example, bromine reacts with fluorine according to the equation:

$$Br_2(l) + F_2(g) \longrightarrow 2\,BrF(g)$$

Again, like all compounds that form between two nonmetals, the interhalogen compounds contain covalent bonds.



▲ Three Halogens

EXAMPLE 4.9 Alkali Metal and Halogen Reactions

Write a balanced chemical equation for each reaction.

(a) the reaction between potassium metal and bromine gas
(b) the reaction between rubidium metal and liquid water
(c) the reaction between gaseous chlorine and solid iodine

SOLUTION

(a) Alkali metals react with halogens to form metal halides. Write the formulas for the reactants and the metal halide product (making sure to write the correct ionic chemical formula for the metal halide, as outlined in Section 3.5), and then balance the equation.	$2\,K(s) + Br_2(g) \longrightarrow 2\,KBr(s)$
(b) Alkali metals react with water to form the dissolved metal ion, the hydroxide ion, and hydrogen gas. Write the skeletal equation including these and then balance it.	$2\,Rb(s) + 2\,H_2O(l) \longrightarrow 2\,Rb^+(aq) + 2\,OH^-(aq) + H_2(g)$

—Continued on the next page

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. What is the greenhouse effect?

2. Why are scientists concerned about increases in atmospheric carbon dioxide? What is the source of the increase?

3. What is a balanced chemical equation?

4. Identify the reactants and products in this chemical equation.
$4\,NH_3(g) + 5\,O_2(g) \longrightarrow 4\,NO(g) + 6\,H_2O(g)$

5. Why must chemical equations be balanced?

6. What is reaction stoichiometry? What is the significance of the coefficients in a balanced chemical equation?

7. In a chemical reaction, what is the limiting reactant? What do we mean when we say a reactant is in excess?

8. In a chemical reaction, what is the theoretical yield and the percent yield?

9. We typically calculate the percent yield using the actual yield and theoretical yield in units of mass (grams or kilograms). Would the percent yield be different if the actual yield and theoretical yield were in units of amount (moles)?

10. What is a combustion reaction? Why are combustion reactions important? Give an example.

11. Write a general equation for the reaction of an alkali metal with:
- **a.** a halogen
- **b.** water

12. Write a general equation for the reaction of a halogen with:
- **a.** a metal
- **b.** hydrogen
- **c.** another halogen

PROBLEMS BY TOPIC

Writing and Balancing Chemical Equations

13. 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 for this reaction. (*Note:* This is a simplified representation of this reaction.)
MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2*

14. 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.)

15. 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.
MISSED THIS? *Read Sections 4.2, 4.5; Watch KCV 4.2, IWE 4.2*

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

17. 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.
MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2*

18. 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.

19. Write the balanced chemical equation for each reaction.
MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2*
- **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.

20. 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.

21. 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.
MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.3*

22. 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.

23. Balance each chemical equation.
MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2, 4.3*
- **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.** $H_2(g) + Cl_2(g) \longrightarrow HCl(g)$

24. 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.** $FeS(s) + HCl(aq) \longrightarrow FeCl_2(aq) + H_2S(g)$

Reaction Stoichiometry

25. Consider the unbalanced equation for the combustion of hexane:

$$C_6H_{14}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Balance the equation and determine how many moles of O_2 are required to react completely with 7.2 moles of C_6H_{14}.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

26. Consider the unbalanced equation for the neutralization of acetic acid:

$$HC_2H_3O_2(aq) + Ba(OH)_2(aq) \longrightarrow H_2O(l) + Ba(C_2H_3O_2)_2(aq)$$

Balance the equation and determine how many moles of $Ba(OH)_2$ are required to completely neutralize 0.461 mole of $HC_2H_3O_2$.

27. Calculate how many moles of NO_2 form when each quantity of reactant completely reacts.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

$$2\,N_2O_5(g) \longrightarrow 4\,NO_2(g) + O_2(g)$$

a. 2.5 mol N_2O_5
b. 6.8 mol N_2O_5
c. 15.2 g N_2O_5
d. 2.87 kg N_2O_5

28. Calculate how many moles of NH_3 form when each quantity of reactant completely reacts.

$$3\,N_2H_4(l) \longrightarrow 4\,NH_3(g) + N_2(g)$$

a. 2.6 mol N_2H_4
b. 3.55 mol N_2H_4
c. 65.3 g N_2H_4
d. 4.88 kg N_2H_4

29. Consider the balanced equation:

$$SiO_2(s) + 3\,C(s) \longrightarrow SiC(s) + 2\,CO(g)$$

Complete the table showing the appropriate number of moles of reactants and products. If the number of moles of a *reactant* is provided, fill in the required amount of the other reactant, as well as the moles of each product that forms. If the number of moles of a *product* is provided, fill in the required amount of each reactant to make that amount of product, as well as the amount of the other product that forms.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

Mol SiO_2	Mol C	Mol SiC	Mol CO
3	____	____	____
____	6	____	____
____	____	____	10
2.8	____	____	____
____	1.55	____	____

30. Consider the balanced equation:

$$2\,N_2H_4(g) + N_2O_4(g) \longrightarrow 3\,N_2(g) + 4\,H_2O(g)$$

Complete the table showing the appropriate number of moles of reactants and products. If the number of moles of a *reactant* is provided, fill in the required amount of the other reactant, as well as the moles of each product that forms. If the number of moles of a *product* is provided, fill in the required amount of each reactant to make that amount of product, as well as the amount of the other product that forms.

Mol N_2H_4	Mol N_2O_4	Mol N_2	Mol H_2O
2	____	____	____
____	5	____	____
____	____	____	10
2.5	____	____	____
____	4.2	____	____
____	____	11.8	____

31. Hydrobromic acid dissolves solid iron according to the reaction:

$$Fe(s) + 2\,HBr(aq) \longrightarrow FeBr_2(aq) + H_2(g)$$

What mass of HBr (in g) do you need to dissolve a 3.2-g pure iron bar on a padlock? What mass of H_2 would the complete reaction of the iron bar produce?

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

32. Sulfuric acid dissolves aluminum metal according to the reaction:

$$2\,Al(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3\,H_2(g)$$

Suppose you want to dissolve an aluminum block with a mass of 15.2 g. What minimum mass of H_2SO_4 (in g) do you need? What mass of H_2 gas (in g) does the complete reaction of the aluminum block produce?

33. For each of the reactions, calculate the mass (in grams) of the product that forms when 3.67 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

a. $\underline{Ba(s)} + Cl_2(g) \longrightarrow BaCl_2(s)$
b. $\underline{CaO(s)} + CO_2(g) \longrightarrow CaCO_3(s)$
c. $2\,\underline{Mg(s)} + O_2(g) \longrightarrow 2\,MgO(s)$
d. $4\,\underline{Al(s)} + 3\,O_2(g) \longrightarrow 2\,Al_2O_3(s)$

34. For each of the reactions, calculate the mass (in grams) of the product that forms when 15.39 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

a. $2\,K(s) + \underline{Cl_2(g)} \longrightarrow 2\,KCl(s)$
b. $2\,K(s) + \underline{Br_2(l)} \longrightarrow 2\,KBr(s)$
c. $4\,Cr(s) + 3\,\underline{O_2(g)} \longrightarrow 2\,Cr_2O_3(s)$
d. $2\,\underline{Sr(s)} + O_2(g) \longrightarrow 2\,SrO(s)$

Limiting Reactant, Theoretical Yield, and Percent Yield

35. Find the limiting reactant for each initial amount of reactants.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

$$2\,Na(s) + Br_2(g) \longrightarrow 2\,NaBr(s)$$

a. 2 mol Na, 2 mol Br_2
b. 1.8 mol Na, 1.4 mol Br_2
c. 2.5 mol Na, 1 mol Br_2
d. 12.6 mol Na, 6.9 mol Br_2

36. Find the limiting reactant for each initial amount of reactants.

$$4\,Al(s) + 3\,O_2(g) \longrightarrow 2\,Al_2O_3(s)$$

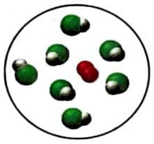
a. 1 mol Al, 1 mol O_2
b. 4 mol Al, 2.6 mol O_2
c. 16 mol Al, 13 mol O_2
d. 7.4 mol Al, 6.5 mol O_2

37. Consider the reaction:

$$4\,HCl(g) + O_2(g) \longrightarrow 2\,H_2O(g) + 2\,Cl_2(g)$$

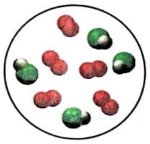
Each molecular diagram represents an initial mixture of reactants. Which mixture produces the greatest amount of products? How many molecules of Cl_2 form from the reaction mixture that produces the greatest amount of products?

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*



a.

b.



c.

38. Consider the reaction:

$$2\,CH_3OH(g) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

Each of the molecular diagrams represents an initial mixture of the reactants. Which reaction mixture produces the greatest amount of products? How many CO_2 molecules form from the reaction mixture that produces the greatest amount of products?

a.

b.

c.

39. Calculate the theoretical yield of the product (in moles) for each initial amount of reactants.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

$$Ti(s) + 2\,Cl_2(g) \longrightarrow TiCl_4(l)$$

a. 4 mol Ti, 4 mol Cl_2
b. 7 mol Ti, 17 mol Cl_2
c. 12.4 mol Ti, 18.8 mol Cl_2

40. Calculate the theoretical yield of product (in moles) for each initial amount of reactants.

$$3\,Mn(s) + 2\,O_2(g) \longrightarrow Mn_3O_4(s)$$

a. 3 mol Mn, 3 mol O_2
b. 4 mol Mn, 7 mol O_2
c. 27.5 mol Mn, 43.8 mol O_2

41. Zinc sulfide reacts with oxygen according to the reaction:

$$2\,ZnS(s) + 3\,O_2(g) \longrightarrow 2\,ZnO(s) + 2\,SO_2(g)$$

A reaction mixture initially contains 4.2 mol ZnS and 6.8 mol O_2. Once the reaction has occurred as completely as possible, what amount (in moles) of the excess reactant remains?

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

42. Iron(II) sulfide reacts with hydrochloric acid according to the reaction:

$$FeS(s) + 2\,HCl(aq) \longrightarrow FeCl_2(s) + H_2S(g)$$

A reaction mixture initially contains 0.223 mol FeS and 0.652 mol HCl. Once the reaction has occurred as completely as possible, what amount (in moles) of the excess reactant remains?

43. For the reaction shown, calculate the theoretical yield of product (in grams) for each initial amount of reactants.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

$$2\,Al(s) + 3\,Cl_2(g) \longrightarrow 2\,AlCl_3(s)$$

a. 2.0 g Al, 2.0 g Cl_2
b. 7.5 g Al, 24.8 g Cl_2
c. 0.235 g Al, 1.15 g Cl_2

44. For the reaction shown, calculate the theoretical yield of the product (in grams) for each initial amount of reactants.

$$Ti(s) + 2\,F_2(g) \longrightarrow TiF_4(s)$$

a. 5.0 g Ti, 5.0 g F_2
b. 2.4 g Ti, 1.6 g F_2
c. 0.233 g Ti, 0.288 g F_2

45. Iron(III) oxide reacts with carbon monoxide according to the equation:

$$Fe_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Fe(s) + 3\,CO_2(g)$$

A reaction mixture initially contains 22.55 g Fe_2O_3 and 14.78 g CO. Once the reaction has occurred as completely as possible, what mass (in g) of the excess reactant remains?

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

46. Elemental phosphorus reacts with chlorine gas according to the equation:

$$P_4(s) + 6\,Cl_2(g) \longrightarrow 4\,PCl_3(l)$$

A reaction mixture initially contains 45.69 g P_4 and 131.3 g Cl_2. Once the reaction has occurred as completely as possible, what mass (in g) of the excess reactant remains?

47. Lead ions can be precipitated from solution with KCl according to the reaction:

$$Pb^{2+}(aq) + 2\,KCl(aq) \longrightarrow PbCl_2(s) + 2\,K^+(aq)$$

When 28.5 g KCl is added to a solution containing 25.7 g Pb^{2+}, a $PbCl_2$ precipitate forms. The precipitate is filtered and dried and found to have a mass of 29.4 g. Determine the limiting reactant, theoretical yield of $PbCl_2$, and percent yield for the reaction.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

48. Magnesium oxide can be made by heating magnesium metal in the presence of oxygen. The balanced equation for the reaction is:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

When 10.1 g of Mg reacts with 10.5 g O_2, 11.9 g MgO is collected. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

49. Urea (CH_4N_2O) is a common fertilizer that is synthesized by the reaction of ammonia (NH_3) with carbon dioxide:

$$2\,NH_3(aq) + CO_2(aq) \longrightarrow CH_4N_2O(aq) + H_2O(l)$$

In an industrial synthesis of urea, a chemist combines 136.4 kg of ammonia with 211.4 kg of carbon dioxide and obtains 168.4 kg of urea. Determine the limiting reactant, theoretical yield of urea, and percent yield for the reaction.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

50. Many computer chips are manufactured from silicon, which occurs in nature as SiO_2. When SiO_2 is heated to melting, it reacts with solid carbon to form liquid silicon and carbon monoxide gas. In an industrial preparation of silicon, 155.8 kg of SiO_2 reacts with 78.3 kg of carbon to produce 66.1 kg of silicon. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

Combustion, Alkali Metal, and Halogen Reactions

51. Complete and balance each combustion reaction equation.

MISSED THIS? *Read Section 4.5*

a. $S(s) + O_2(g) \longrightarrow$
b. $C_3H_6(g) + O_2(g) \longrightarrow$
c. $Ca(s) + O_2(g) \longrightarrow$
d. $C_5H_{12}S(l) + O_2(g) \longrightarrow$

52. Complete and balance each combustion reaction equation:

a. $C_4H_6(g) + O_2(g) \longrightarrow$
b. $C(s) + O_2(g) \longrightarrow$
c. $CS_2(s) + O_2(g) \longrightarrow$
d. $C_3H_8O(l) + O_2(g) \longrightarrow$

53. Write a balanced chemical equation for the reaction of solid strontium with iodine gas.

MISSED THIS? *Read Section 4.5*

54. Write a balanced chemical equation for the reaction between lithium metal and chlorine gas.

55. Write a balanced chemical equation for the reaction of solid lithium with liquid water. **MISSED THIS?** *Read Section 4.5*

56. Write a balanced chemical equation for the reaction of solid potassium with liquid water.

57. Write a balanced equation for the reaction of hydrogen gas with bromine gas. **MISSED THIS?** *Read Section 4.5*

58. Write a balanced equation for the reaction of chlorine gas with fluorine gas.

CUMULATIVE PROBLEMS

59. The combustion of gasoline produces carbon dioxide and water. Assume gasoline to be pure octane (C_8H_{18}) and calculate the mass (in kg) of carbon dioxide that is added to the atmosphere per 1.0 kg of octane burned. (*Hint:* Begin by writing a balanced equation for the combustion reaction.)

60. Many home barbeques are fueled with propane gas (C_3H_8). What mass of carbon dioxide (in kg) is produced upon the complete combustion of 18.9 L of propane (approximate contents of one 5-gallon tank)? Assume that the density of the liquid propane in the tank is 0.621 g/mL. (*Hint:* Begin by writing a balanced equation for the combustion reaction.)

61. Aspirin can be made in the laboratory by reacting acetic anhydride ($C_4H_6O_3$) with salicylic acid ($C_7H_6O_3$) to form aspirin ($C_9H_8O_4$) and acetic acid ($C_2H_4O_2$). The balanced equation is:

$$C_4H_6O_3 + C_7H_6O_3 \longrightarrow C_9H_8O_4 + C_2H_4O_2$$

In a laboratory synthesis, a student begins with 3.00 mL of acetic anhydride (density = 1.08 g/mL) and 1.25 g of salicylic acid. Once the reaction is complete, the student collects 1.22 g of aspirin. Determine the limiting reactant, theoretical yield of aspirin, and percent yield for the reaction.

62. The combustion of liquid ethanol (C_2H_5OH) produces carbon dioxide and water. After 4.62 mL of ethanol (density = 0.789 g/mL) is allowed to burn in the presence of 15.55 g of oxygen gas, 3.72 mL of water (density = 1.00 g/mL) is collected. Determine the limiting reactant, theoretical yield of H_2O, and percent yield for the reaction. (*Hint:* Write a balanced equation for the combustion of ethanol.)

63. A loud classroom demonstration involves igniting a hydrogen-filled balloon. The hydrogen within the balloon reacts explosively with oxygen in the air to form water. If the balloon is filled with a mixture of hydrogen and oxygen, the explosion is even louder than if the balloon is filled only with hydrogen—the intensity of the explosion depends on the relative amounts of oxygen and hydrogen within the balloon. Look at the molecular views representing different amounts of hydrogen and oxygen in four different balloons. Based on the balanced chemical equation, which balloon will make the loudest explosion?

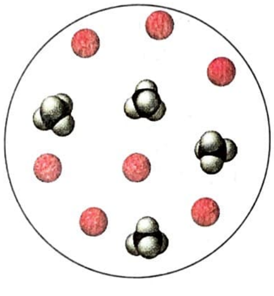
a.

b.

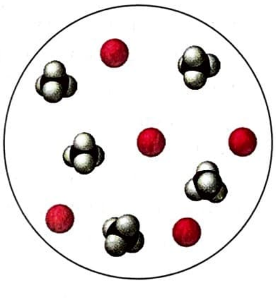
c.

d.

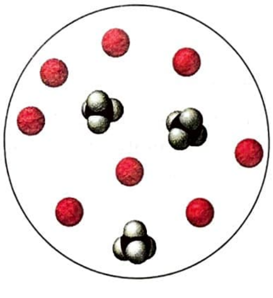
64. Gaseous methane reacts with oxygen to form carbon dioxide and water vapor. Write a balanced equation for the combustion reaction and determine which mixture has neither reactant in excess.



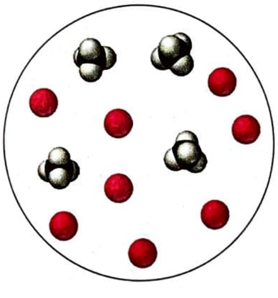
a.



b.



c.



d.

65. The reaction of NH_3 and O_2 forms NO and water. The NO can be used to convert P_4 to P_4O_6, forming N_2 in the process. The P_4O_6 can be treated with water to form H_3PO_3, which forms PH_3 and H_3PO_4 when heated. Find the mass of PH_3 that forms from the reaction of 1.00 g of NH_3.

66. An important reaction that takes place in a blast furnace during the production of iron is the formation of iron metal and CO_2 from Fe_2O_3 and CO. Determine the mass of Fe_2O_3 required to form 910 kg of iron. Determine the amount of CO_2 that forms in this process.

67. A liquid fuel mixture contains 30.35% hexane (C_6H_{14}), 15.85% heptane (C_7H_{16}), and the rest octane (C_8H_{18}). What maximum mass of carbon dioxide is produced by the complete combustion of 10.0 kg of this fuel mixture?

68. Titanium occurs in the magnetic mineral ilmenite ($FeTiO_3$), which is often found mixed with sand. The ilmenite can be separated from the sand with magnets. The titanium can then be extracted from the ilmenite by the following set of reactions:

$$FeTiO_3(s) + 3\,Cl_2(g) + 3\,C(s) \longrightarrow 3\,CO(g) + FeCl_2(s) + TiCl_4(g)$$

$$TiCl_4(g) + 2\,Mg(s) \longrightarrow 2\,MgCl_2(l) + Ti(s)$$

Suppose that an ilmenite–sand mixture contains 22.8% ilmenite by mass and that the first reaction is carried out with a 90.8% yield. If the second reaction is carried out with an 85.9% yield, what mass of titanium can be obtained from 1.00 kg of the ilmenite–sand mixture?

> Science may be described as the art of systemic oversimplification—the art of discerning what we may with advantage omit.
>
> —KARL POPPER (1902–1994)

CHAPTER 5

Introduction to Solutions and Aqueous Reactions

In this chapter, we turn to describing chemical reactions that occur in water. You have probably witnessed many of these types of reactions in your daily life because they are so common. Have you ever mixed baking soda with vinegar and observed the subsequent bubbling? Have you noticed the hard water deposits that form on plumbing fixtures? These reactions—and many others, including those that occur within the watery environment of living cells—are aqueous chemical reactions, the subject of this chapter.

The spherified cherry, a product of a cooking technique known as molecular gastronomy.

5.1 Molecular Gastronomy and the Spherified Cherry

One of my favorite chefs is the Spaniard José Andrés, who owns and operates restaurants all over the world including *The Bazaar* in Los Angeles and Miami, and *Jaleo* in Washington, D.C. Andrés was born in the province of Austurias, Spain, the same province as my grandfather (no wonder I like him so much). Andrés is not only an amazing chef, who is credited with bringing small plates (or *tapas*) to America, but he also dabbles in *molecular gastronomy*, a way of preparing food that involves chemistry.

▶ **FIGURE 5.1 The Spherified Cherry** The spherified cherry is made by precipitating an encapsulating layer around cherry juice.

A common chemical reaction in molecular gastronomy is precipitation. In a *precipitation reaction*, two *solutions*—homogeneous mixtures often containing a solid dissolved in a liquid—are mixed. Upon mixing, a solid (or *precipitate*) forms. For example, when we mix solutions of lead(II) nitrate and potassium iodide, a brilliant yellow solid forms. The solid is lead(II) iodide.

In molecular gastronomy, chefs use a similar precipitation reaction—called spherification—to encapsulate liquids. Among the most popular molecular gastronomy creations is the spherified cherry (Figure 5.1▲). To make a spherified cherry, chefs take juice from real cherries and mix it with a calcium salt (such as calcium chloride), which dissolves in the cherry juice. They then carefully pour the cherry juice into a bath of sodium alginate. Sodium alginate is a sodium salt that dissolves into water, resulting in the presence of alginate ions. When the calcium ions in the cherry juice encounter the alginate ions in the bath, a precipitation reaction occurs. In this case, the precipitation reaction forms in the area immediately surrounding the cherry juice, forming an encapsulating sphere around the juice. The result is a spherical, edible "cherry" that ruptures in the mouth and releases its juice.

In this chapter, we explore solutions, focusing especially on *aqueous* solutions (solutions in which one component is water). The cherry juice and calcium chloride mixture just discussed is an example of an aqueous solution. Other common aqueous solutions include seawater, vinegar, and the watery environment within biological cells. We will also explore the chemical reactions that occur within solutions, such as precipitation reactions, which have many common applications.

5.2 Solution Concentration

WATCH **NOW!**

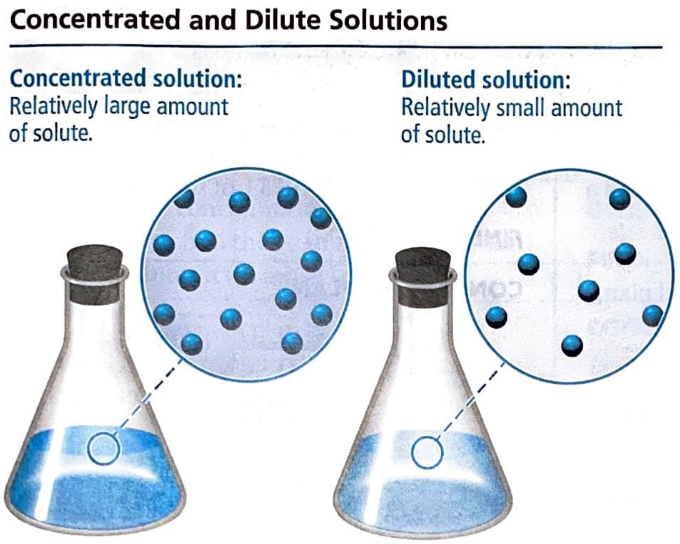
KEY CONCEPT VIDEO 5.2
Solution Concentration

The reactions that occur in lakes, streams, and oceans, as well as the reactions that occur in every cell within our bodies, take place in water. Chemical reactions involving reactants dissolved in water are among the most common and important. A homogeneous mixture of two substances—such as salt and water—is a **solution**. The majority component of the mixture is the **solvent**, and the minority component is the **solute**. An **aqueous solution** is one in which water acts as the solvent. In this section, we examine how to quantify the concentration of a solution (the amount of solute relative to solvent).

Solution Concentration

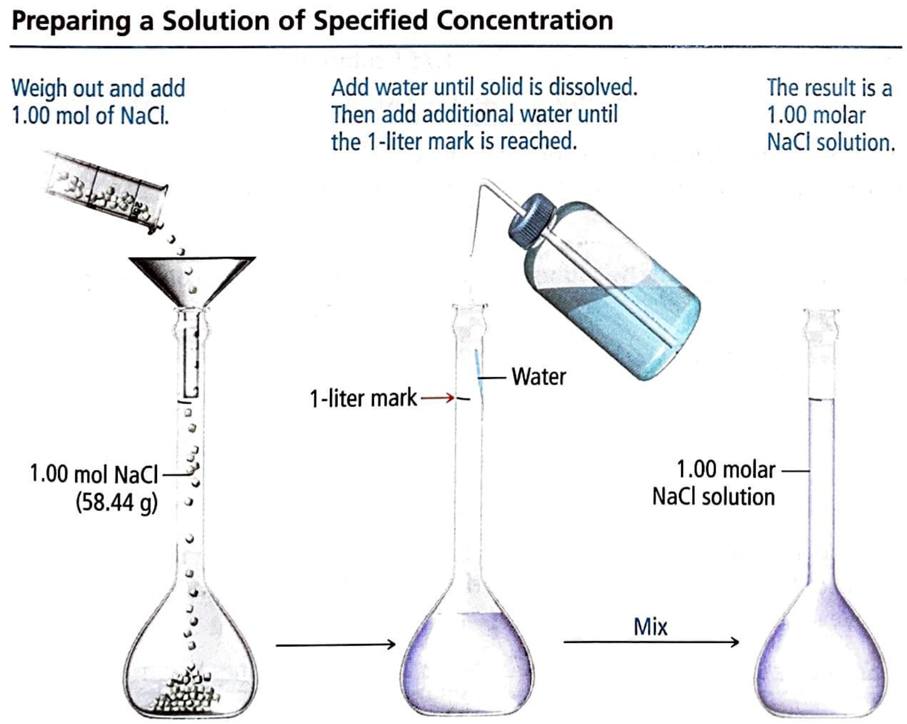
The amount of solute in a solution is variable. For example, we can add just a little salt to water to make a **dilute solution**, one that contains a small amount of solute relative to the solvent, or we can add a lot of salt to water to make a **concentrated solution**, one that contains a large amount of solute relative to the solvent (Figure 5.2▶). A common way to express solution concentration is **molarity (M)**, the amount of solute (in moles) divided by the volume of solution (in liters):

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$


Concentrated and Dilute Solutions

◀ **FIGURE 5.2 Concentrated and Dilute Solutions**

Notice that molarity is a ratio of the amount of solute per liter of *solution*, not per liter of solvent. To make an aqueous solution of a specified molarity, we usually put the solute into a flask and then add water to reach the desired volume of solution. For example, to make 1 L of a 1 M NaCl solution, we add 1 mol of NaCl to a flask and then add enough water to make 1 L of solution (Figure 5.3▼). We *do not* combine 1 mol of NaCl with 1 L of water because the resulting solution would have a total volume different from 1 L and therefore a molarity different than 1 M. To calculate molarity, divide the amount of the solute in moles by the volume of the solution (solute *and* solvent) in liters, as shown in Example 5.1.


Preparing a Solution of Specified Concentration

◀ **FIGURE 5.3 Preparing a 1 Molar NaCl Solution**

MOLARITY How many moles of solute are required to make 3.0 L of a 2.0 M solution?

(a) 2.0 mol solute (b) 3.0 mol solute
(c) 4.0 mol solute (d) 6.0 mol solute

5.1
Cc
Conceptual Connection
ANSWER **NOW!**

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 5.1

EXAMPLE 5.1 Calculating Solution Concentration

What is the molarity of a solution containing 25.5 g KBr dissolved in enough water to make 1.75 L of solution?

SORT You are given the mass of KBr and the volume of a solution and asked to find its molarity.

GIVEN: 25.5 g KBr, 1.75 L of solution

FIND: molarity (M)

STRATEGIZE When formulating the conceptual plan, think about the definition of molarity, the amount of solute *in moles* per liter of solution.

You are given the mass of KBr, so first use the molar mass of KBr to convert from g KBr to mol KBr.

Then use the number of moles of KBr and liters of solution to find the molarity.

CONCEPTUAL PLAN

$$\boxed{\text{g KBr}} \xrightarrow[\frac{1\text{ mol}}{119.00\text{ g}}]{} \boxed{\text{mol KBr}}$$

$$\boxed{\text{mol KBr, L solution}} \xrightarrow[\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}]{} \boxed{\text{Molarity}}$$

RELATIONSHIPS USED

molar mass of KBr = 119.00 g/mol

SOLVE Follow the conceptual plan. Begin with g KBr and convert to mol KBr; then use mol KBr and L solution to calculate molarity.

SOLUTION

$$25.5\ \cancel{\text{g KBr}} \times \frac{1\text{ mol KBr}}{119.00\ \cancel{\text{g KBr}}} = 0.21\underline{4}29\text{ mol KBr}$$

$$\begin{aligned}\text{molarity (M)} &= \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}\\ &= \frac{0.21\underline{4}29\text{ mol KBr}}{1.75\text{ L solution}}\\ &= 0.122\text{ M}\end{aligned}$$

CHECK The units of the answer (M) are correct. The magnitude is reasonable since common solutions range in concentration from 0 to about 18 M. Concentrations significantly above 18 M are suspect and should be double-checked.

FOR PRACTICE 5.1 Calculate the molarity of a solution made by adding 45.4 g of $NaNO_3$ to a flask and dissolving it with water to create a total volume of 2.50 L.

FOR MORE PRACTICE 5.1 What mass of KBr (in grams) do you need to make 250.0 mL of a 1.50 M KBr solution?

Using Molarity in Calculations

We can use the molarity of a solution as a conversion factor between moles of the solute and liters of the solution. For example, a 0.500 M NaCl solution contains 0.500 moles NaCl for every liter of solution:

$$\frac{0.500\text{ mol NaCl}}{\text{L solution}} \quad \textit{converts} \quad \boxed{\text{L solution}} \longrightarrow \boxed{\text{mol NaCl}}$$

This conversion factor converts from liters solution to mole NaCl. If we want to convert in the other direction, we invert the conversion factor:

$$\frac{\text{L solution}}{0.500\text{ mol NaCl}} \quad \textit{converts} \quad \boxed{\text{mol NaCl}} \longrightarrow \boxed{\text{L solution}}$$

Example 5.2 illustrates how to use molarity in this way.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 5.2

EXAMPLE 5.2 Using Molarity in Calculations

How many liters of a 0.125 M NaOH solution contain 0.255 mol of NaOH?

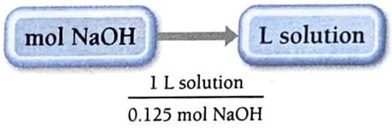
SORT You are given the concentration of a NaOH solution. You are asked to find the volume of the solution that contains a given amount (in moles) of NaOH.

GIVEN: 0.125 M NaOH solution, 0.255 mol NaOH

FIND: volume of NaOH solution (in L)

STRATEGIZE The conceptual plan begins with mol NaOH and shows the conversion to L of solution using molarity as a conversion factor.

CONCEPTUAL PLAN



RELATIONSHIPS USED

$$0.125\ \text{M NaOH} = \frac{0.125\ \text{mol NaOH}}{1\ \text{L solution}}$$

SOLVE Follow the conceptual plan. Begin with mol NaOH and convert to L solution.

SOLUTION

$$0.255\ \cancel{\text{mol NaOH}} \times \frac{1\ \text{L solution}}{0.125\ \cancel{\text{mol NaOH}}} = 2.04\ \text{L solution}$$

CHECK The units of the answer (L) are correct. The magnitude is reasonable because the solution contains 0.125 mol per liter. Therefore, roughly 2 L contains the given amount of moles (0.255 mol).

FOR PRACTICE 5.2 How many grams of sucrose ($C_{12}H_{22}O_{11}$) are in 1.55 L of a 0.758 M sucrose solution?

FOR MORE PRACTICE 5.2 How many mL of a 0.155 M KCl solution contain 2.55 g KCl?

SOLUTIONS If we dissolve 25 g of salt in 251 g of water, what is the mass of the resulting solution?

- **(a)** 251 g
- **(b)** 276 g
- **(c)** 226 g

5.2 Cc Conceptual Connection

ANSWER **NOW!**

Solution Dilution

To save space in storerooms, laboratories often store solutions in concentrated forms called **stock solutions**. For example, hydrochloric acid is frequently stored as a 12 M stock solution. However, many lab procedures call for much less concentrated hydrochloric acid solutions, so we must dilute the stock solution to the required concentration. How do we know how much of the stock solution to use? The easiest way to solve dilution problems is to use the following dilution equation:

When diluting acids, always add the concentrated acid to the water. Never add water to concentrated acid solutions, as the heat generated may cause the concentrated acid to splatter and burn your skin.

$$M_1V_1 = M_2V_2 \qquad [5.1]$$

where M_1 and V_1 are the molarity and volume of the initial concentrated solution, and M_2 and V_2 are the molarity and volume of the final diluted solution. This equation works because the molarity multiplied by the volume gives the number of moles of solute, which is the same in both solutions.

$$\begin{aligned} M_1V_1 &= M_2V_2 \\ \text{mol}_1 &= \text{mol}_2 \end{aligned}$$

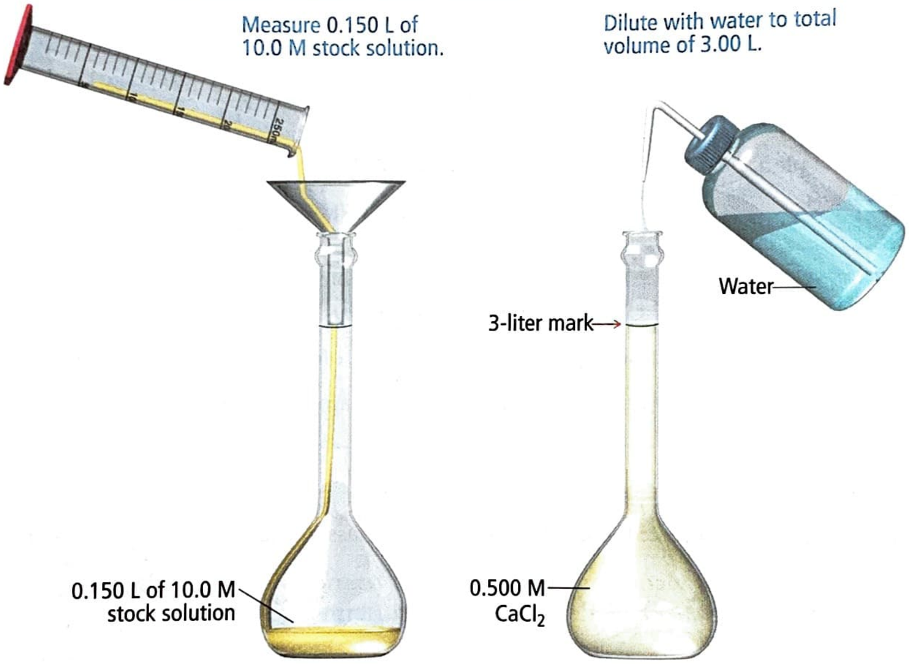
In other words, the number of moles of solute does not change when we dilute a solution.

For example, suppose a laboratory procedure calls for 3.00 L of a 0.500 M $CaCl_2$ solution. How should we prepare this solution from a 10.0 M stock solution? We solve

Equation 5.1 for V_1, the volume of the stock solution required for the dilution, and then substitute in the correct values to calculate it.

$$\begin{aligned} M_1V_1 &= M_2V_2 \\ V_1 &= \frac{M_2V_2}{M_1} \\ &= \frac{0.500\ \cancel{\text{mol/L}} \times 3.00\ \text{L}}{10.0\ \cancel{\text{mol/L}}} \\ &= 0.150\ \text{L} \end{aligned}$$

Consequently, we make the solution by diluting 0.150 L of the stock solution to a total volume of 3.00 L (V_2). The resulting solution will be 0.500 M in $CaCl_2$ (Figure 5.4▼).



Diluting a Solution

$$M_1V_1 = M_2V_2$$

$$\frac{10.0\ \text{mol}}{\cancel{\text{L}}} \times 0.150\ \cancel{\text{L}} = \frac{0.500\ \text{mol}}{\cancel{\text{L}}} \times 3.00\ \cancel{\text{L}}$$

$$1.50\ \text{mol} = 1.50\ \text{mol}$$

▶ **FIGURE 5.4 Preparing 3.00 L of 0.500 M $CaCl_2$ from a 10.0 M Stock Solution**

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 5.3

EXAMPLE 5.3 Solution Dilution

To what volume should you dilute 0.200 L of a 15.0 M NaOH solution to obtain a 3.00 M NaOH solution?

SORT You are given the initial volume, initial concentration, and final concentration of a solution. You need to determine the final volume.

GIVEN: $V_1 = 0.200$ L
$M_1 = 15.0$ M
$M_2 = 3.00$ M

FIND: V_2

STRATEGIZE Equation 5.1 relates the initial and final volumes and concentrations for solution dilution problems. You are asked to find V_2. The other quantities (V_1, M_1, and M_2) are all given in the problem.

CONCEPTUAL PLAN

$$\boxed{V_1, M_1, M_2} \xrightarrow[M_1V_1 = M_2V_2]{} \boxed{V_2}$$

RELATIONSHIPS USED

$M_1V_1 = M_2V_2$

SOLVE Begin with the solution dilution equation and solve it for V_2.

Substitute in the required quantities and calculate V_2.

Make the solution by diluting 0.200 L of the stock solution to a total volume of 1.00 L (V_2). The resulting solution will have a concentration of 3.00 M.

SOLUTION $M_1V_1 = M_2V_2$

$$V_2 = \frac{M_1V_1}{M_2}$$

$$= \frac{15.0\ \cancel{\text{mol/L}} \times 0.200\ \text{L}}{3.00\ \cancel{\text{mol/L}}}$$

$$= 1.00\ \text{L}$$

CHECK The final units (L) are correct. The magnitude of the answer is reasonable because the solution is diluted from 15.0 M to 3.00 M, a factor of five. Therefore, the volume should increase by a factor of five.

FOR PRACTICE 5.3 To what volume (in mL) should you dilute 100.0 mL of a 5.00 M $CaCl_2$ solution to obtain a 0.750 M $CaCl_2$ solution?

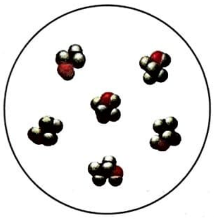
FOR MORE PRACTICE 5.3 What volume of a 6.00 M $NaNO_3$ solution should you use to make 0.525 L of a 1.20 M $NaNO_3$ solution?

5.3 **Cc** **Conceptual Connection**

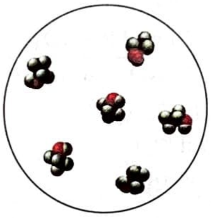
ANSWER NOW!

SOLUTION DILUTION The image shown at the far right represents a small volume within 500 mL of aqueous ethanol (CH_3CH_2OH) solution. (The water molecules have been omitted for clarity.)

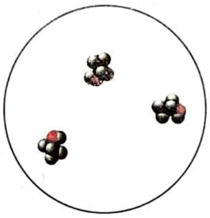
Which of the following images best represents the same volume of the solution after we add an additional 500 mL of water?



(a)



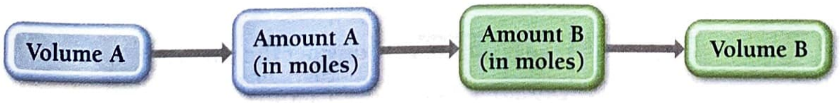
(b)



(c)

5.3 Solution Stoichiometry

In Section 4.3, we discussed how we can use the coefficients in chemical equations as conversion factors between the amounts of reactants (in moles) and the amounts of products (in moles). In aqueous reactions, quantities of reactants and products are often specified in terms of volumes and concentrations. We can use the volume and concentration of a reactant or product to calculate its amount in moles. We can then use the stoichiometric coefficients in the chemical equation to convert to the amount of another reactant or product in moles. The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product:



We make the conversions between solution volumes and amounts of solute in moles using the molarities of the solutions. We make the conversions between amounts in moles of A and B using the stoichiometric coefficients from the balanced chemical equation. Example 5.4 demonstrates solution stoichiometry.

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE 5.4

EXAMPLE 5.4 Solution Stoichiometry

What volume (in L) of a 0.150 M KCl solution will completely react with 0.150 L of a 0.175 M $Pb(NO_3)_2$ solution according to the following balanced chemical equation?

$$2\,KCl(aq) + Pb(NO_3)_2(aq) \longrightarrow PbCl_2(s) + 2\,KNO_3(aq)$$

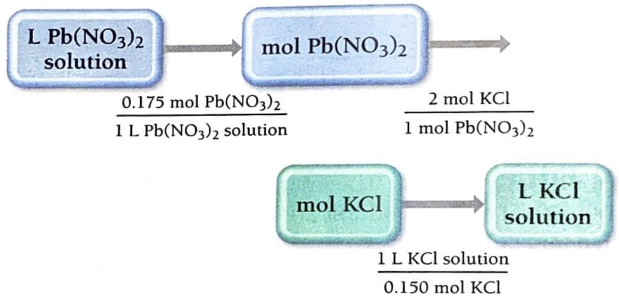
SORT You are given the volume and concentration of a $Pb(NO_3)_2$ solution. You are asked to find the volume of KCl solution (of a given concentration) required to react with it.

GIVEN: 0.150 L of $Pb(NO_3)_2$ solution, 0.175 M $Pb(NO_3)_2$ solution, 0.150 M KCl solution

FIND: volume KCl solution (in L)

STRATEGIZE The conceptual plan has the form: volume A → amount A (in moles) → amount B (in moles) → volume B. Use the molar concentrations of the KCl and $Pb(NO_3)_2$ solutions as conversion factors between the number of moles of reactants in these solutions and their volumes. Use the stoichiometric coefficients from the balanced equation to convert between number of moles of $Pb(NO_3)_2$ and number of moles of KCl.

CONCEPTUAL PLAN



RELATIONSHIPS USED

$$M\,Pb(NO_3)_2 = \frac{0.175\text{ mol }Pb(NO_3)_2}{1\text{ L }Pb(NO_3)_2\text{ solution}}$$

2 mol KCl : 1 mol $Pb(NO_3)_2$

$$M\,KCl = \frac{0.150\text{ mol KCl}}{1\text{ L KCl solution}}$$

SOLVE Begin with L $Pb(NO_3)_2$ solution and follow the conceptual plan to arrive at L KCl solution.

SOLUTION

$$0.150\text{ L }\cancel{Pb(NO_3)_2\text{ solution}} \times \frac{0.175\ \cancel{\text{mol }Pb(NO_3)_2}}{1\ \cancel{\text{L }Pb(NO_3)_2\text{ solution}}}$$

$$\times \frac{2\ \cancel{\text{mol KCl}}}{1\ \cancel{\text{mol }Pb(NO_3)_2}} \times \frac{1\text{ L KCl solution}}{0.150\ \cancel{\text{mol KCl}}} = 0.350\text{ L KCl solution}$$

CHECK The final units (L KCl solution) are correct. The magnitude (0.350 L) is reasonable because the reaction stoichiometry requires 2 mol of KCl per mole of $Pb(NO_3)_2$. Since the concentrations of the two solutions are not very different (0.150 M compared to 0.175 M), the volume of KCl required is roughly two times the 0.150 L of $Pb(NO_3)_2$ given in the problem.

FOR PRACTICE 5.4 What volume (in mL) of a 0.150 M HNO_3 solution will completely react with 35.7 mL of a 0.108 M Na_2CO_3 solution according to the following balanced chemical equation?

$$Na_2CO_3(aq) + 2\,HNO_3(aq) \longrightarrow 2\,NaNO_3(aq) + CO_2(g) + H_2O(l)$$

FOR MORE PRACTICE 5.4 In the previous reaction, what mass (in grams) of carbon dioxide forms?

ANSWER **NOW!**

5.4
Cc
Conceptual Connection

SOLUTION STOICHIOMETRY Consider the reaction:

$$2\,A(aq) + B(aq) \longrightarrow C(aq)$$

What is the limiting reactant if you mix equal volumes of a 1 M solution of A and a 1 M solution of B?

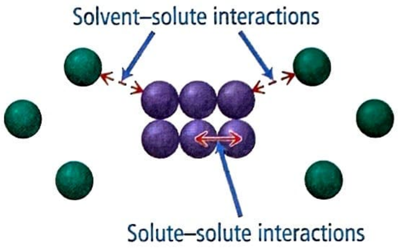
(a) A **(b)** B

5.4 Types of Aqueous Solutions and Solubility

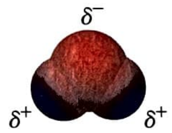
Consider two familiar aqueous solutions: salt water and sugar water. Salt water is a homogeneous mixture of NaCl and H_2O, and sugar water is a homogeneous mixture of $C_{12}H_{22}O_{11}$ and H_2O. You may have made these solutions yourself by adding table salt or sugar to water. As you stir either of these two substances into the water, the substance seems to disappear. However, you know that the original substance is still present because the mixture tastes salty or sweet. How do solids such as salt and sugar dissolve in water?

When a solid is put into a liquid solvent, the attractive forces that hold the solid together (the solute–solute interactions) compete with the attractive forces between the solvent molecules and the particles that compose the solid (the solvent–solute interactions), as shown in Figure 5.5▶. For example, when sodium chloride is put into water, there is a competition between the attraction of Na^+ cations and Cl^- anions to each other (due to their opposite charges) and the attraction of Na^+ and Cl^- to water molecules. The attraction of Na^+ and Cl^- to water is based on the *polar nature* of the water molecule. For reasons we discuss later in this book (Section 10.6), the oxygen atom in water is electron-rich, giving it a partial negative charge (δ^-), as shown in Figure 5.6▶. The hydrogen atoms, in contrast, are electron-poor, giving them a partial positive charge (δ^+). As a result, the positively charged sodium ions are strongly attracted to the oxygen side of the water molecule (which has a partial negative charge), and the negatively charged chloride ions are attracted to the hydrogen side of the water molecule (which has a partial positive charge), as shown in Figure 5.7▼. In the case of NaCl, the attraction between the separated ions and the water molecules overcomes the attraction of sodium and chloride ions to each other, and the sodium chloride dissolves in the water (Figure 5.8▼).

Solute and Solvent Interactions



▲ **FIGURE 5.5 Solute and Solvent Interactions**

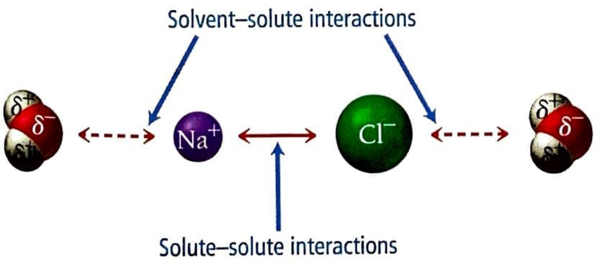


▲ **FIGURE 5.6 Charge Distribution in Water** An uneven distribution of electrons causes the oxygen side of the water molecule to have a partial negative charge and the hydrogen side to have a partial positive charge.

Electrolyte and Nonelectrolyte Solutions

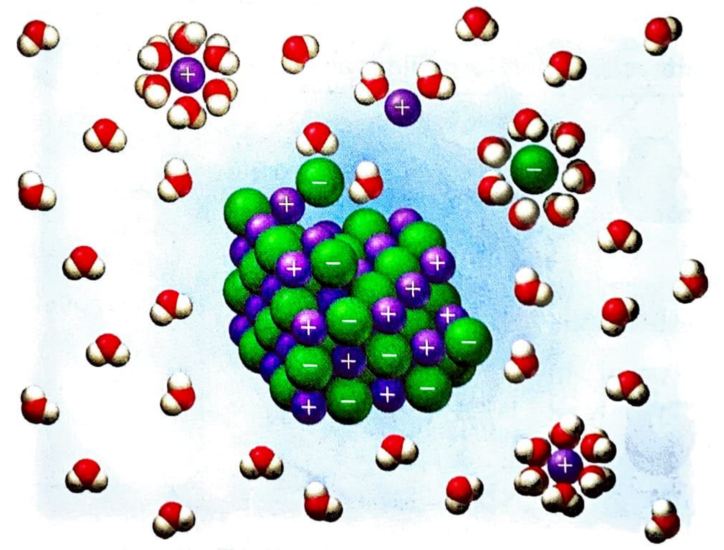
As Figure 5.9▶ illustrates, a salt solution conducts electricity while a sugar solution does not. The difference between the ways that salt (an ionic compound) and sugar (a molecular compound) dissolve in water illustrates a fundamental difference between types of solutions. Ionic compounds such as the sodium chloride in the previous example dissociate into their component ions when they dissolve in water. An NaCl solution, represented as NaCl(*aq*), does not contain any NaCl units, but rather dissolved Na^+ ions and Cl^- ions. The dissolved ions act as charge carriers, allowing the solution to conduct electricity. Substances that dissolve in water to form solutions that conduct electricity are **electrolytes**. Substances such

Interactions in a Sodium Chloride Solution



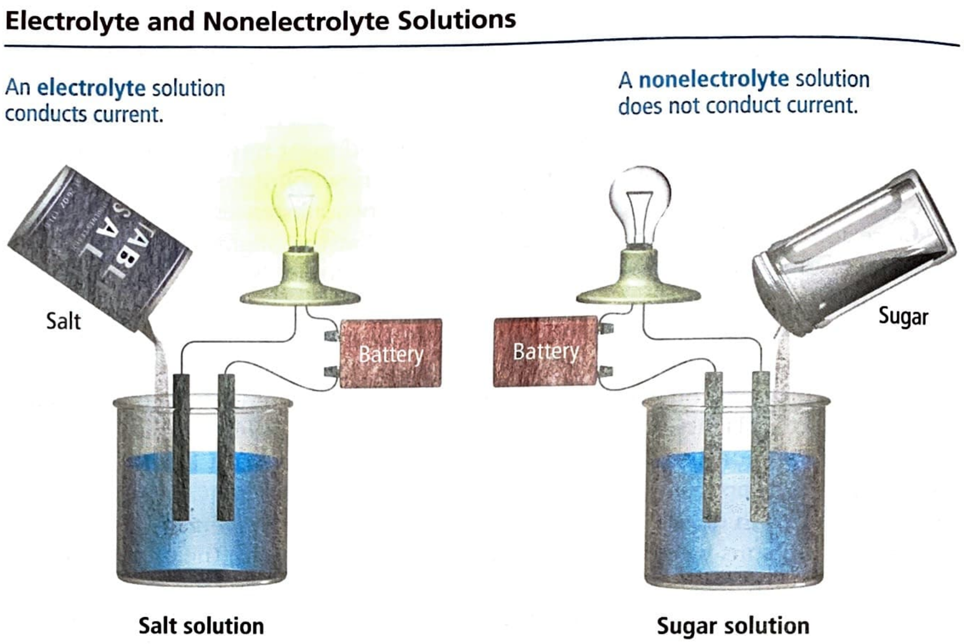
▲ **FIGURE 5.7 Solute and Solvent Interactions in a Sodium Chloride Solution** When sodium chloride is put into water, the attraction of Na^+ and Cl^- ions to water molecules competes with the attraction between the oppositely charged ions themselves.

Dissolution of an Ionic Compound



▲ **FIGURE 5.8 Sodium Chloride Dissolving in Water** The attraction between water molecules and the ions of sodium chloride causes NaCl to dissolve in the water.

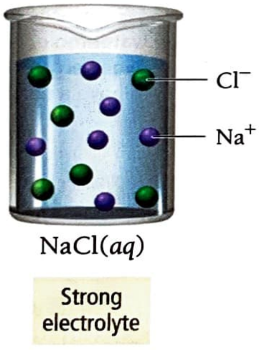
▶ **FIGURE 5.9 Electrolyte and Nonelectrolyte Solutions**

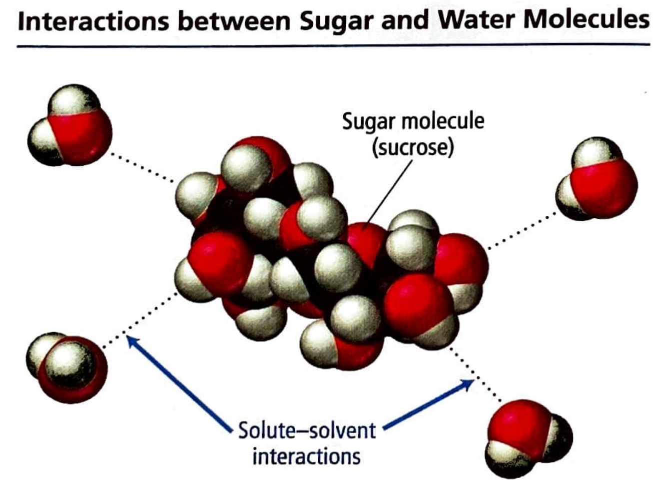


as sodium chloride that completely dissociate into ions when they dissolve in water are **strong electrolytes**, and the resulting solutions are *strong electrolyte solutions*.

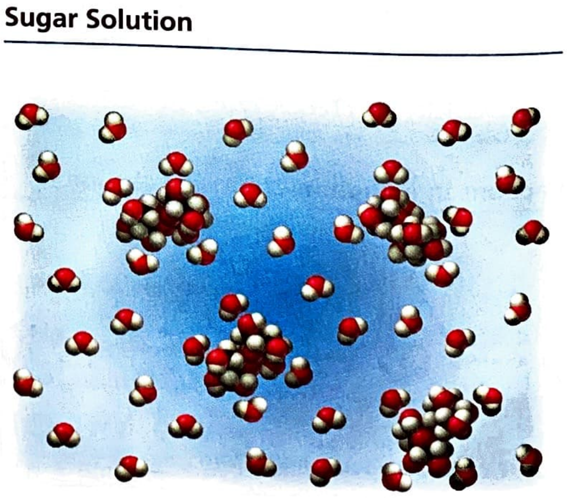
In contrast to sodium chloride, sugar is a molecular compound. Most molecular compounds—with the important exception of acids, which we discuss shortly—dissolve in water as intact molecules. Sugar dissolves because the attraction between sugar molecules and water molecules (shown in Figure 5.10▼) overcomes the attraction of sugar molecules to each other (Figure 5.11▼). So unlike a sodium chloride solution (which is composed of dissociated ions), a sugar solution is composed of intact $C_{12}H_{22}O_{11}$ molecules homogeneously mixed with the water molecules. Compounds such as sugar that do not dissociate into ions when dissolved in water are called **nonelectrolytes**, and the resulting solutions—called *nonelectrolyte solutions*—do not conduct electricity.

Acids, which we introduced in Section 3.6, are molecular compounds, but they ionize—form ions—when they dissolve in water. Hydrochloric acid (HCl) is a molecular compound that ionizes into H^+ and Cl^- when it dissolves in water. HCl is an example of a **strong acid**, one that completely ionizes in solution. Since strong acids completely





▲ **FIGURE 5.10 Sugar and Water Interactions** Partial charges on sugar molecules and water molecules (which we will discuss more fully in Chapter 12) result in attractions between the sugar molecules and water molecules.



▲ **FIGURE 5.11 A Sugar Solution** Sugar dissolves because the attractions between sugar molecules and water molecules, which both contain a distribution of electrons that results in partial positive and partial negative charges, overcome the attractions between sugar molecules to each other.

ionize in solution, they are also strong electrolytes. We represent the complete ionization of a strong acid with a single reaction arrow between the acid and its ionized form:

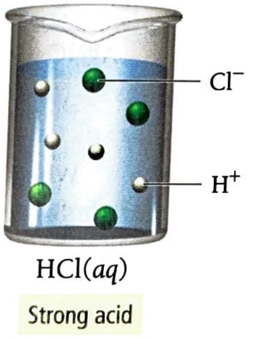
$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

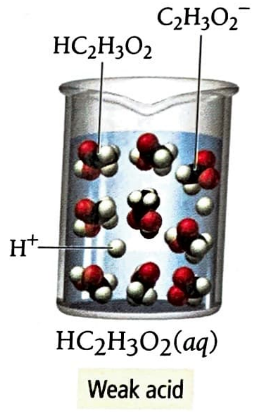
Unlike soluble ionic compounds, which contain ions and therefore *dissociate* in water, acids are molecular compounds that *ionize* in water.

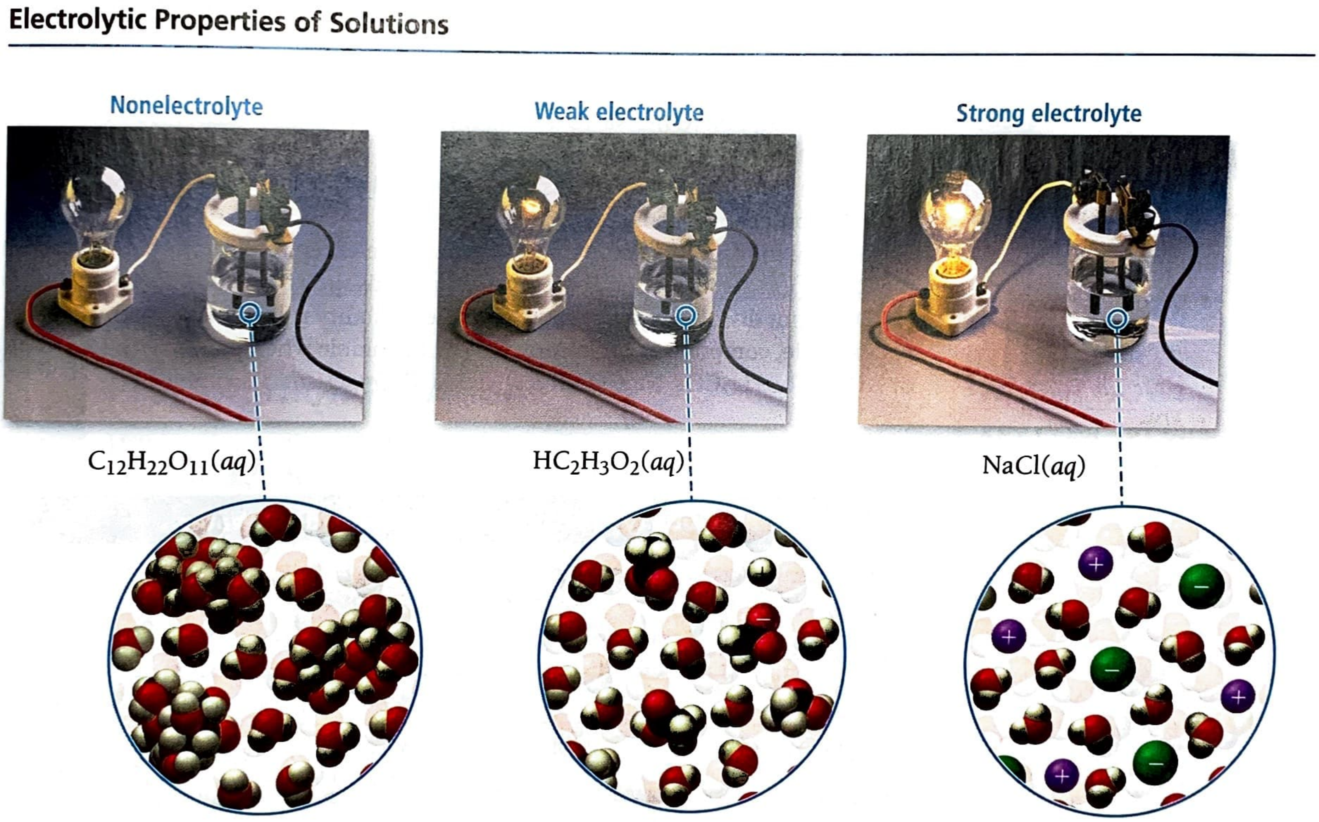
Many acids are **weak acids**; they do not completely ionize in water. For example, acetic acid ($HC_2H_3O_2$), the acid in vinegar, is a weak acid. A solution of a weak acid is composed mostly of the nonionized acid—only a small percentage of the acid molecules ionize. We represent the partial ionization of a weak acid with opposing half arrows between the reactants and products:

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

Weak acids are **weak electrolytes**, and the resulting solutions—called *weak electrolyte solutions*—conduct electricity only weakly. Figure 5.12▼ summarizes the electrolytic properties of solutions.







▲ **FIGURE 5.12 Electrolytic Properties of Solutions**

ELECTROLYTE SOLUTIONS Which aqueous solution conducts electricity?

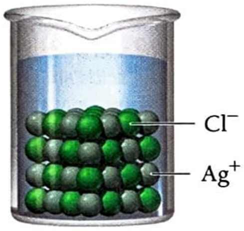
(a) 1.0 M KBr **(b)** 1.0 M $C_6H_{12}O_6$ **(c)** 1.0 M CH_3OH

5.5 Cc Conceptual Connection

ANSWER **NOW!**

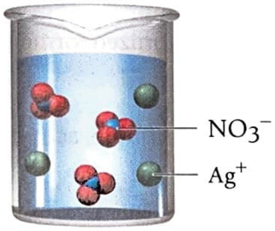
The Solubility of Ionic Compounds

As we have just discussed, when an ionic compound dissolves in water, the resulting solution contains not the intact ionic compound itself, but its component ions dissolved in water. However, not all ionic compounds dissolve in water. If we add AgCl to water, for example, it remains solid and appears as a white powder at the bottom of the water.



$AgCl(s)$

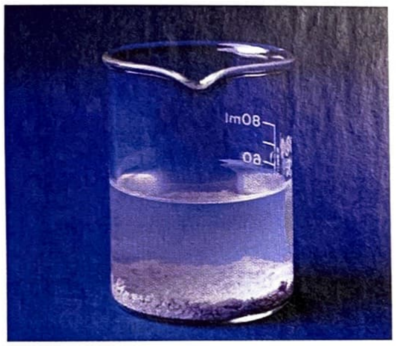
Insoluble



$AgNO_3(aq)$

Soluble

In general, a compound is termed **soluble** if it dissolves in water and **insoluble** if it does not. However, these classifications are a bit of an oversimplification. (In reality, solubility is a continuum and even "insoluble" compounds dissolve to some extent, though usually orders of magnitude less than soluble compounds.) For example, silver nitrate is soluble. If we mix solid $AgNO_3$ with water, it dissolves and forms a strong electrolyte solution. Silver chloride, as we just saw, is almost completely insoluble. If we mix solid AgCl with water, virtually all of it remains as a solid within the liquid water.



▲ AgCl does not dissolve in water; it remains as a white powder at the bottom of the beaker.

Whether a particular compound is soluble or insoluble depends on several factors. In Section 13.3, we will examine more closely the energy changes associated with solution formation. For now, we can follow a set of empirical rules that chemists have inferred from observations on many ionic compounds. Table 5.1 summarizes these *solubility rules*.

The solubility rules state that compounds containing the sodium ion are soluble. That means that compounds such as NaBr, $NaNO_3$, Na_2SO_4, NaOH, and Na_2CO_3 all dissolve in water to form strong electrolyte solutions. Similarly, the solubility rules state that compounds containing the NO_3^- ion are soluble. That means that compounds such as $AgNO_3$, $Pb(NO_3)_2$, $NaNO_3$, $Ca(NO_3)_2$, and $Sr(NO_3)_2$ all dissolve in water to form strong electrolyte solutions.

Notice that when compounds containing polyatomic ions such as NO_3^- dissolve, the polyatomic ions dissolve as intact units.

The solubility rules also state that, with some exceptions, compounds containing the CO_3^{2-} ion are insoluble. Therefore, compounds such as $CuCO_3$, $CaCO_3$, $SrCO_3$, and $FeCO_3$ do not dissolve in water. Note that the solubility rules contain many exceptions. For example, compounds containing CO_3^{2-} *are* soluble when paired with Li^+, Na^+, K^+, or NH_4^+. Thus Li_2CO_3, Na_2CO_3, K_2CO_3, and $(NH_4)_2CO_3$ are all soluble.

TABLE 5.1 ■ Solubility Rules for Ionic Compounds in Water

Compounds Containing the Following Ions Are Generally Soluble	**Exceptions**
Li^+, Na^+, K^+, and NH_4^+	None
NO_3^- and $C_2H_3O_2^-$	None
Cl^-, Br^-, and I^-	When these ions pair with Ag^+, Hg_2^{2+}, or Pb^{2+}, the resulting compounds are insoluble.
SO_4^{2-}	When SO_4^{2-} pairs with Sr^{2+}, Ba^{2+}, Pb^{2+}, Ag^+, or Ca^{2+}, the resulting compound is insoluble.
Compounds Containing the Following Ions Are Generally Insoluble	**Exceptions**
OH^- and S^{2-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble. When S^{2-} pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is soluble. When OH^- pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is slightly soluble.
CO_3^{2-} and PO_4^{3-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 5.5

EXAMPLE 5.5 Predicting Whether an Ionic Compound Is Soluble

Predict whether each compound is soluble or insoluble.

(a) $PbCl_2$ **(b)** $CuCl_2$ **(c)** $Ca(NO_3)_2$ **(d)** $BaSO_4$

SOLUTION

(a) Insoluble. Compounds containing Cl^- are normally soluble, but Pb^{2+} is an exception.

(b) Soluble. Compounds containing Cl^- are normally soluble, and Cu^{2+} is not an exception.

(c) Soluble. Compounds containing NO_3^- are always soluble.

(d) Insoluble. Compounds containing SO_4^{2-} are normally soluble, but Ba^{2+} is an exception.

FOR PRACTICE 5.5 Predict whether each compound is soluble or insoluble.

(a) NiS **(b)** $Mg_3(PO_4)_2$ **(c)** Li_2CO_3 **(d)** NH_4Cl

SOLUBILITY OF IONIC COMPOUNDS The presence of one of the following ions within a compound indicates that a compound is soluble with no exceptions. Which ion?

(a) OH^- **(b)** SO_4^{2-} **(c)** NO_3^-

5.6 Cc Conceptual Connection

ANSWER **NOW!**

5.5 Precipitation Reactions

WATCH **NOW!**

KEY CONCEPT VIDEO 5.5
Reactions in Solution

Have you ever taken a bath in hard water? Hard water contains dissolved ions such as Ca^{2+} and Mg^{2+} that diminish the effectiveness of soap. These ions react with soap to form a gray soap scum that may appear as a "bathtub ring" after you drain the tub. Hard water is particularly troublesome when washing clothes. Consider how your white shirt would look covered with the soap scum from the bathtub and you can understand the problem. Consequently, most laundry detergents include substances designed to remove Ca^{2+} and Mg^{2+} from the laundry mixture. The most common substance used for this purpose is sodium carbonate, which dissolves in water to form sodium cations (Na^+) and carbonate (CO_3^{2-}) anions:

$$Na_2CO_3(aq) \longrightarrow 2\,Na^+(aq) + CO_3^{2-}(aq)$$

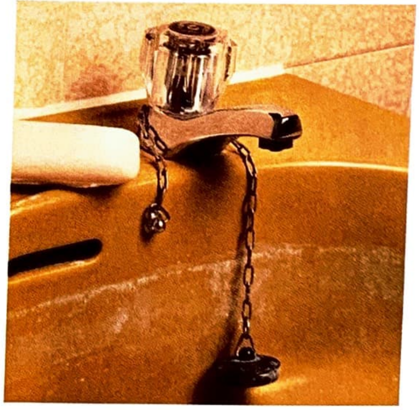
Sodium carbonate is soluble, but calcium carbonate and magnesium carbonate are not (see the solubility rules in Table 5.1). Consequently, the carbonate anions react with dissolved Mg^{2+} and Ca^{2+} ions in hard water to form solids that *precipitate* from (or come out of) solution:

$$Mg^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow MgCO_3(s)$$
$$Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

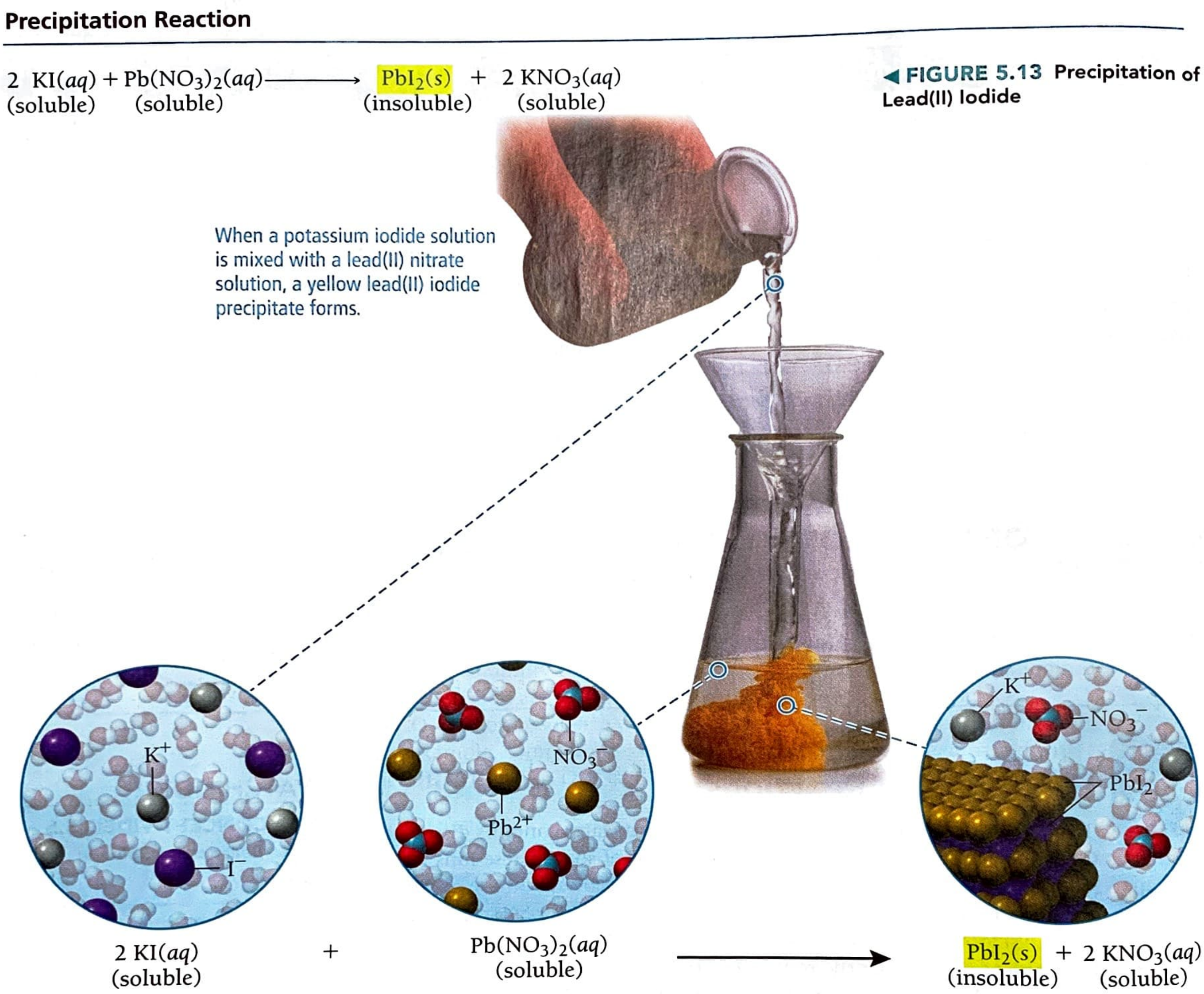
The precipitation of these ions prevents their reaction with the soap, eliminating curd and preventing white shirts from turning gray.

The reactions between CO_3^{2-} and Mg^{2+} and Ca^{2+} are examples of **precipitation reactions**, reactions in which a solid or **precipitate** forms when we mix two solutions. Precipitation reactions are common in chemistry. As another example, consider potassium iodide and lead(II) nitrate, which form colorless, strong electrolyte solutions when dissolved in water separately. When the two solutions are combined, however, a brilliant yellow precipitate forms (Figure 5.13▶). We describe this precipitation reaction with the following chemical equation:

$$2\,KI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2\,KNO_3(aq)$$



▲ The reaction of ions in hard water with soap produces a gray soap scum that is visible after you drain the bathwater.



◀ **FIGURE 5.13 Precipitation of Lead(II) Iodide**

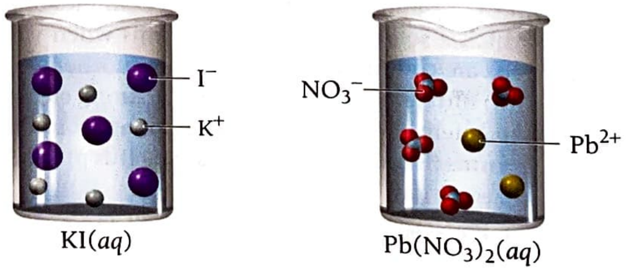
Precipitation reactions do not always occur when two aqueous solutions are mixed. For example, if we combine solutions of KI(*aq*) and NaCl(*aq*), nothing happens (Figure 5.14▶):

$$KI(aq) + NaCl(aq) \not\rightarrow \text{NO REACTION}$$

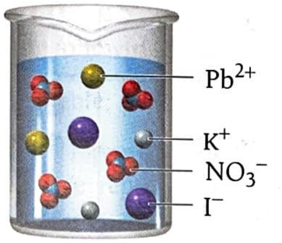
The key to predicting precipitation reactions is to understand that *only insoluble compounds form precipitates.* In a precipitation reaction, two solutions containing soluble compounds combine and an insoluble compound precipitates. Consider the precipitation reaction described previously:

$$\underset{\text{soluble}}{2\,KI(aq)} + \underset{\text{soluble}}{Pb(NO_3)_2(aq)} \longrightarrow \underset{\text{insoluble}}{PbI_2(s)} + \underset{\text{soluble}}{2\,KNO_3(aq)}$$

KI and $Pb(NO_3)_2$ are both soluble, but the precipitate, PbI_2, is insoluble. Before mixing, KI(*aq*) and $Pb(NO_3)_2(aq)$ are both dissociated in their respective solutions:

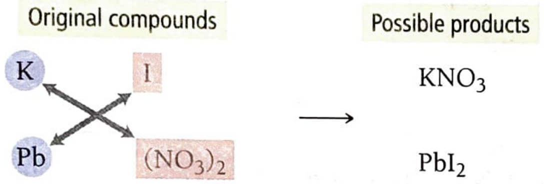


The instant that the solutions come into contact, all four ions are present:



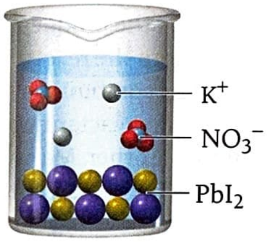
KI(*aq*) and $Pb(NO_3)_2(aq)$

Now, new compounds—one or both of which might be insoluble—are possible. Specifically, the cation from either compound can pair with the anion from the other to form possibly insoluble products:



If the possible products are both soluble, no reaction occurs and no precipitate forms. If one or both of the possible products are insoluble, a precipitation reaction occurs. In this case, KNO_3 is soluble, but PbI_2 is insoluble. Consequently, PbI_2 precipitates.

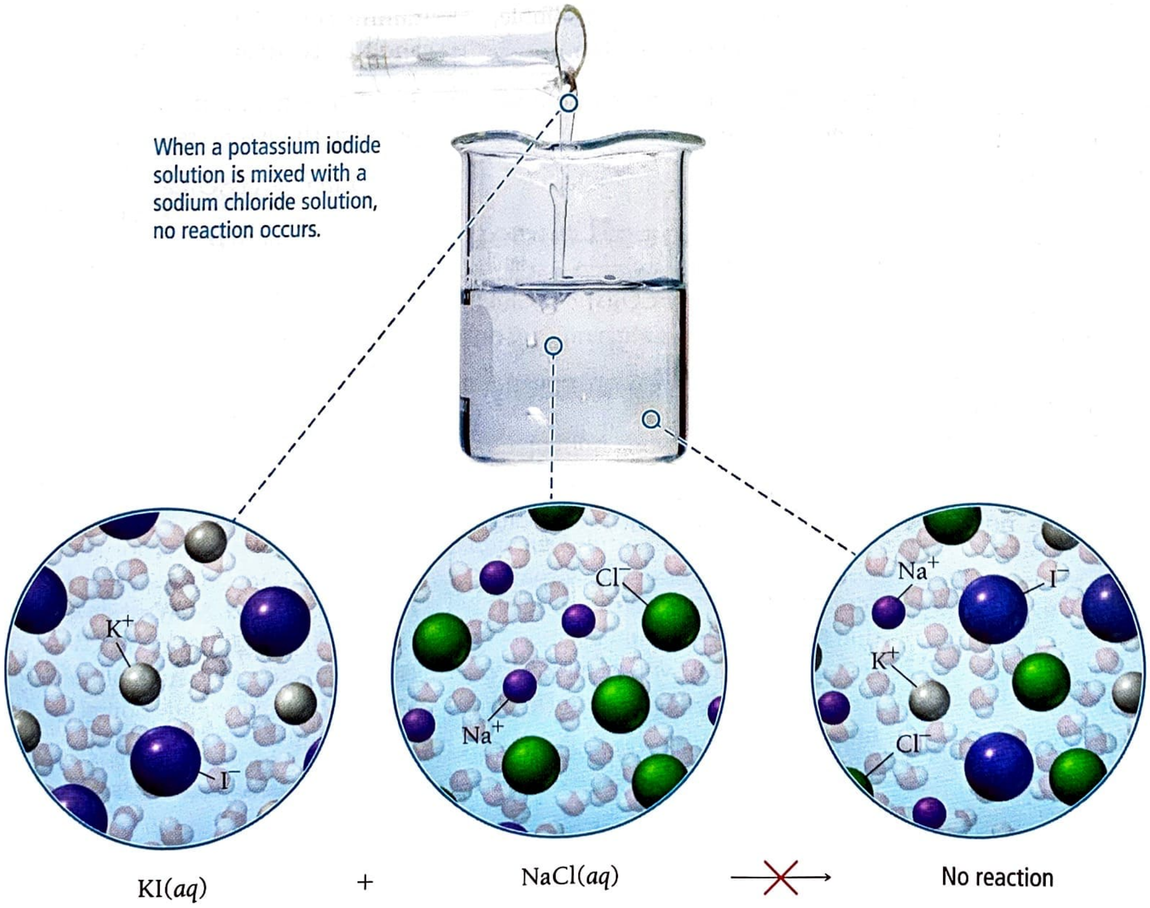
To predict whether a precipitation reaction will occur when two solutions are mixed and to write an equation for the reaction, we use the procedure that follows. The steps are outlined in the left column, and two examples illustrating how to apply the procedure are shown in the center and right columns.



$PbI_2(s)$ and $KNO_3(aq)$

No Reaction

$KI(aq) + NaCl(aq) \not\longrightarrow$ No reaction


When a potassium iodide solution is mixed with a sodium chloride solution, no reaction occurs.

◀ **FIGURE 5.14 No Precipitation**

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE VIDEO 5.6

HOW TO: *Write Equations for Precipitation Reactions*	**EXAMPLE 5.6** **Writing Equations for Precipitation Reactions** Write an equation for the precipitation reaction that occurs (if any) when solutions of potassium carbonate and nickel(II) chloride are mixed.	**EXAMPLE 5.7** **Writing Equations for Precipitation Reactions** Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium nitrate and lithium sulfate are mixed.
1. Write the formulas of the two compounds being mixed as reactants in a chemical equation.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow$	$NaNO_3(aq) + Li_2SO_4(aq) \longrightarrow$
2. Below the equation, write the formulas of the products that could form from the reactants. Obtain these by combining the cation from each reactant with the anion from the other. Make sure to write correct formulas for these ionic compounds using the procedure demonstrated in Section 3.5.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow$ Possible products KCl $NiCO_3$	$NaNO_3(aq) + Li_2SO_4(aq) \longrightarrow$ Possible products $LiNO_3$ Na_2SO_4
3. Refer to the solubility rules to determine whether any of the possible products are insoluble.	KCl is soluble. (Compounds containing Cl^- are usually soluble, and K^+ is not an exception.) $NiCO_3$ is insoluble. (Compounds containing CO_3^{2-} are usually insoluble, and Ni^{2+} is not an exception.)	$LiNO_3$ is soluble. (Compounds containing NO_3^- are soluble, and Li^+ is not an exception.) Na_2SO_4 is soluble. (Compounds containing SO_4^{2-} are generally soluble, and Na^+ is not an exception.)
4. If all of the possible products are soluble, there is no precipitate. Write "NO REACTION" after the arrow.	Since this example has an insoluble product, we proceed to the next step.	Since this example has no insoluble product, there is no reaction. $NaNO_3(aq) + Li_2SO_4(aq) \nrightarrow$ NO REACTION
5. If any of the possible products are insoluble, write their formulas as the products of the reaction, using (*s*) to indicate solid. Include an (*aq*) to indicate aqueous after any soluble products.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow NiCO_3(s) + KCl(aq)$	
6. Balance the equation. Remember to adjust only coefficients, not subscripts.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow NiCO_3(s) + 2\,KCl(aq)$	
	FOR PRACTICE 5.6 Write an equation for the precipitation reaction that occurs (if any) when solutions of ammonium chloride and iron(III) nitrate mix.	**FOR PRACTICE 5.7** Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium hydroxide and copper(II) bromide mix.

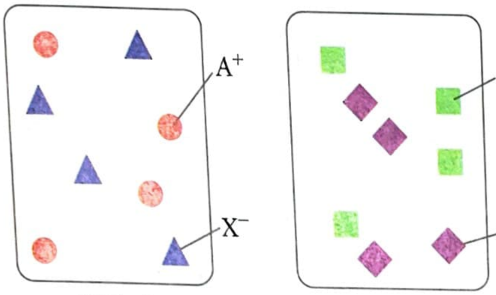
5.7 Conceptual Connection

ANSWER NOW!

PRECIPITATION REACTIONS Consider the generic ionic compounds with the formulas AX and BY and the following solubility rules:

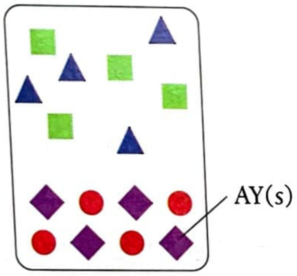
AX soluble; BY soluble; AY soluble; BX insoluble

Let circles represent A^+ ions; squares represent B^+ ions; triangles represent X^- ions; and diamonds represent Y^- ions. We represent solutions of the two compounds (AX and BY) as follows:



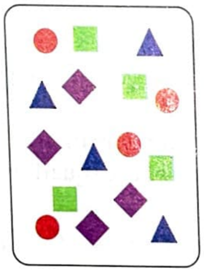
$AX(aq)$ $BY(aq)$

From the answer options, choose the representation that correctly shows the result of mixing the two solutions (AX and BY) and the correct equation to represent the reaction.



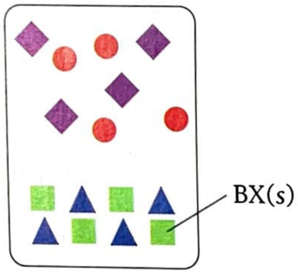
$AX(aq) + BY(aq) \longrightarrow AY(s) + BX(aq)$

(a)



NO REACTION

(b)



$AX(aq) + BY(aq) \longrightarrow BX(s) + AY(aq)$

(c)

5.6 Representing Aqueous Reactions: Molecular, Ionic, and Net Ionic Equations

Consider the following equation for a precipitation reaction:

$$Pb(NO_3)_2(aq) + 2\,KCl(aq) \longrightarrow PbCl_2(s) + 2\,KNO_3(aq)$$

This equation is a **molecular equation**, an equation showing the complete neutral formulas for each compound in the reaction as if they existed as molecules. In actual solutions of soluble ionic compounds, dissolved substances are present as ions. We can write equations for reactions occurring in aqueous solution in a way that better shows the dissociated nature of dissolved ionic compounds. For example, we can rewrite the previous equation as:

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,K^+(aq) + 2\,Cl^-(aq) \longrightarrow PbCl_2(s) + 2\,K^+(aq) + 2\,NO_3^-(aq)$$

Equations such as this, which list all of the ions present as either reactants or products in a chemical reaction, are **complete ionic equations**. Strong electrolytes are always represented as their component ions in ionic equations—weak electrolytes are not.

Notice that in the complete ionic equation, some of the ions in solution appear unchanged on both sides of the equation. These ions are called **spectator ions** because they do not participate in the reaction.

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,K^+(aq) + 2\,Cl^-(aq) \longrightarrow PbCl_2(s) + 2\,K^+(aq) + 2\,NO_3^-(aq)$$

Spectator ions

To simplify the equation and to show more clearly what is happening, we can omit spectator ions:

$$Pb^{2+}(aq) + 2\,Cl^-(aq) \longrightarrow PbCl_2(s)$$

Equations that show only the species that actually change during the reaction are **net ionic equations.**

As another example, consider the reaction between $HCl(aq)$ and $KOH(aq)$:

$$HCl(aq) + KOH(aq) \longrightarrow H_2O(l) + KCl(aq)$$

Since HCl, KOH, and KCl all exist in solution primarily as independent ions, the complete ionic equation is:

$$H^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + K^+(aq) + Cl^-(aq)$$

To write the net ionic equation, we remove the spectator ions, those that are unchanged on both sides of the equation:

$$H^+(aq) + \cancel{Cl^-}(aq) + \cancel{K^+}(aq) + OH^-(aq) \longrightarrow H_2O(l) + \cancel{K^+}(aq) + \cancel{Cl^-}(aq)$$

Spectator ions

The net ionic equation is $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$.

Summarizing Aqueous Equations

- A **molecular equation** is a chemical equation showing the complete, neutral formulas for every compound in a reaction.
- A **complete ionic equation** is a chemical equation showing all of the species as they are actually present in solution: strong electrolytes are therefore represented as their component ions.
- A **net ionic equation** is an equation showing only the species that actually change during the reaction.

ANSWER **NOW!**

5.8
Cc
Conceptual Connection

SPECTATOR IONS Which of the ions listed below is a spectator ion in the complete ionic equation shown here?

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + I^-(aq) \longrightarrow AgI(s) + NO_3^-(aq) + Na^+(aq)$$

(a) $Ag^+(aq)$ **(b)** $NO_3^-(aq)$ **(c)** $I^-(aq)$

EXAMPLE 5.8 Writing Complete Ionic and Net Ionic Equations

Write complete ionic and net ionic equations for each reaction.

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

(b) $HC_2H_3O_2(aq) + KOH(aq) \longrightarrow H_2O(l) + KC_2H_3O_2(aq)$

SOLUTION

(a) Write the complete ionic equation by separating strong electrolytes into their constituent ions. The $Sr_3(PO_4)_2(s)$, precipitating as a solid, remains as one unit.

Complete ionic equation:

$$3\,Sr^{2+}(aq) + 6\,Cl^-(aq) + 6\,Li^+(aq) + 2\,PO_4{}^{3-}(aq) \longrightarrow Sr_3(PO_4)_2(s) + 6\,Li^+(aq) + 6\,Cl^-(aq)$$

Write the net ionic equation by eliminating the spectator ions, those that do not change from one side of the reaction to the other.

Net ionic equation:

$$3\,Sr^{2+}(aq) + 2\,PO_4{}^{3-}(aq) \longrightarrow Sr_3(PO_4)_2(s)$$

(b) Write the complete ionic equation by separating strong electrolytes into their constituent ions. Do not separate $HC_2H_3O_2(aq)$ because it is a weak electrolyte.

Complete ionic equation:

$$HC_2H_3O_2(aq) + K^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + K^+(aq) + C_2H_3O_2{}^-(aq)$$

Write the net ionic equation by eliminating the spectator ions.

Net ionic equation:

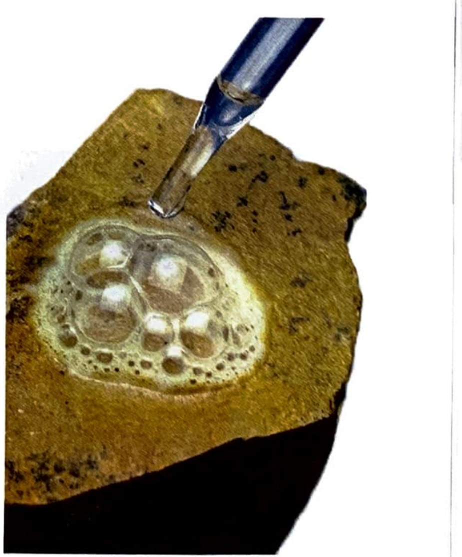
$$HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2{}^-(aq)$$

FOR PRACTICE 5.8 Write the complete ionic equation and net ionic equation for the following reaction:

$$2\,HI(aq) + Ba(OH)_2(aq) \longrightarrow 2\,H_2O(l) + BaI_2(aq)$$

5.7 Acid–Base Reactions

Two other important classes of reactions that occur in aqueous solution are acid–base reactions and gas-evolution reactions. In an **acid–base reaction** (also called a **neutralization reaction**), an acid reacts with a base and the two neutralize each other, producing water (or in some cases a weak electrolyte). In a **gas-evolution reaction**, a gas forms, resulting in bubbling. In both cases, as in precipitation reactions, the reactions occur when the anion from one reactant combines with the cation of the other. Many gas-evolution reactions are also acid–base reactions.



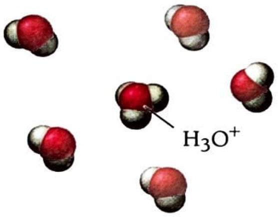
▲ Gas-evolution reactions, such as the reaction of hydrochloric acid (HCl) with limestone ($CaCO_3$), typically produce CO_2; bubbling occurs as the gas is released.

Acid–Base Reactions

Our stomachs contain hydrochloric acid (HCl), which acts in the digestion of food. Certain foods or stress, however, can increase the stomach's acidity to uncomfortable levels, causing acid stomach or heartburn. Antacids are over-the-counter medicines that work by reacting with and neutralizing stomach acid. Antacids employ different *bases*—substances that produce hydroxide (OH^-) ions in water—as neutralizing agents. Milk of magnesia, for example, contains $Mg(OH)_2$ and Mylanta contains $Al(OH)_3$. All antacids, regardless of the base they employ, have the same effect of neutralizing stomach acid and relieving heartburn through *acid–base reactions*.

Recall from Chapter 3 that an acid forms H^+ ions in solution, and we just saw that a base is a substance that produces OH^- ions in solution. More formally:

- Acid: Substance that produces H^+ ions in aqueous solution
- Base: Substance that produces OH^- ions in aqueous solution



▲ **FIGURE 5.15 The Hydronium Ion** Protons normally associate with water molecules in solution to form H_3O^+ ions, which in turn interact with other water molecules.

These definitions of acids and bases, called the **Arrhenius definitions**, are named after Swedish chemist Svante Arrhenius (1859–1927). In Chapter 16, we will learn more general definitions of acids and bases, but these definitions are sufficient to describe neutralization reactions.

According to the Arrhenius definition, HCl is an acid because it produces H^+ ions in solution:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

An H^+ ion is a bare proton. In solution, bare protons normally associate with water molecules to form **hydronium ions** (Figure 5.15◀):

$$H^+(aq) + H_2O(l) \longrightarrow H_3O^+(aq)$$

Chemists use $H^+(aq)$ and $H_3O^+(aq)$ interchangeably to mean the same thing—a hydronium ion. The chemical equation for the ionization of HCl and other acids is often written to show the association of the proton with a water molecule to form the hydronium ion:

$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

As we discussed in Section 5.4, some acids are weak acids—they do not completely ionize in solution. We represent the partial ionization of a weak acid with opposing half arrows.

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

Some acids—called **polyprotic acids**—contain more than one ionizable proton and release them sequentially. For example, sulfuric acid, H_2SO_4, is a **diprotic acid**. It is strong in its first ionizable proton, but weak in its second:

$$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$$

$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$

According to the Arrhenius definition, NaOH is a base because it produces OH^- ions in solution:

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

In analogy to diprotic acids, some bases such as $Sr(OH)_2$ produce two moles of OH^- per mole of the base:

$$Sr(OH)_2(aq) \longrightarrow Sr^{2+}(aq) + 2\,OH^-(aq)$$

Table 5.2 lists common acids and bases. You can find acids and bases in many everyday substances. Foods such as citrus fruits and vinegar contain acids. Soap, baking soda, and milk of magnesia all contain bases.

When we mix an acid and a base, the $H^+(aq)$ from the acid—whether it is weak or strong—combines with the $OH^-(aq)$ from the base to form $H_2O(l)$ (Figure 5.16▶). Consider the reaction between hydrochloric acid and sodium hydroxide:

$$\underset{\text{Acid}}{HCl(aq)} + \underset{\text{Base}}{NaOH(aq)} \longrightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{Salt}}{NaCl(aq)}$$

▲ These household substances all contain acids.

Acid–base reactions generally form water and an ionic compound—called a **salt**—that usually remains dissolved in the solution. The net ionic equation for acid–base reactions involving a strong acid is:

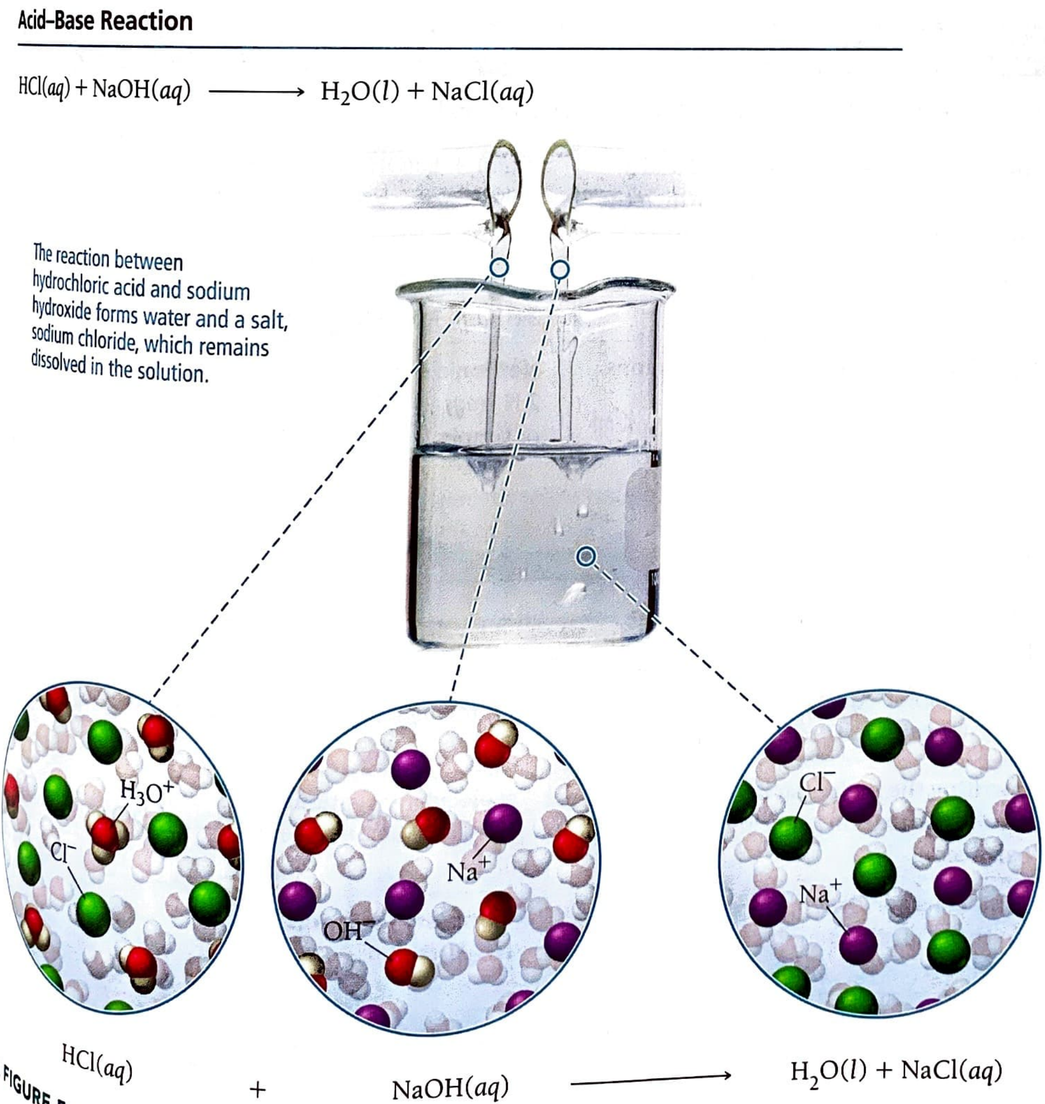
$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

TABLE 5.2 ■ Some Common Acids and Bases

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	$NaOH$
Hydrobromic acid	HBr	Lithium hydroxide	$LiOH$
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Formic acid	$HCHO_2$ (weak acid)		
Acetic acid	$HC_2H_3O_2$ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

*Ammonia does not contain OH^-, but it produces OH^- in a reaction with water that occurs only to a small extent: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

▲ Many common household products contain bases.



Acid–Base Reaction

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

The reaction between hydrochloric acid and sodium hydroxide forms water and a salt, sodium chloride, which remains dissolved in the solution.

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

▲ **FIGURE 5.16 Acid–Base Reaction**

The word *salt* in this sense applies to any ionic compound and is therefore more general than the common usage, which refers only to table salt (NaCl).

However, if the acid is a weak acid, the net ionic equation is slightly different. For example, consider the acid–base equation between hydrofluoric acid and sodium hydroxide:

$$\underset{\text{acid}}{HF(aq)} + \underset{\text{base}}{NaOH(aq)} \longrightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{NaF(aq)}$$

The complete ionic equation and the net ionic equation for this reaction are:

$$HF(aq) + Na^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Na^+(aq) + F^-(aq) \qquad \text{(Complete ionic equation)}$$

$$HF(aq) + OH^-(aq) \longrightarrow H_2O(l) + F^-(aq) \qquad \text{(Net ionic equation)}$$

Notice that, since HF is a weak acid, we do not show it as ionized in the ionic equations.

Another example of an acid–base reaction is the reaction between sulfuric acid and potassium hydroxide:

$$\underset{\text{acid}}{H_2SO_4(aq)} + \underset{\text{base}}{2\,KOH(aq)} \longrightarrow \underset{\text{water}}{2\,H_2O(l)} + \underset{\text{salt}}{K_2SO_4(aq)}$$

Again, notice the pattern of acid and base reacting to form water and a salt.

$$\textbf{Acid} + \textbf{Base} \longrightarrow \textbf{Water} + \textbf{Salt} \qquad \text{(acid–base reactions)}$$

When writing equations for acid–base reactions, write the formula of the salt using the procedure for writing formulas of ionic compounds demonstrated in Section 3.5.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 5.9

EXAMPLE 5.9 Writing Equations for Acid–Base Reactions Involving a Strong Acid

Write a molecular and net ionic equation for the reaction between aqueous HI and aqueous $Ba(OH)_2$.

SOLUTION First identify these substances as an acid and a base. Begin by writing the unbalanced equation in which the acid and the base combine to form water and a salt.	$\underset{\text{acid}}{HI(aq)} + \underset{\text{base}}{Ba(OH)_2(aq)} \longrightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{BaI_2(aq)}$
Next, balance the equation; this is the molecular equation.	**Molecular equation:** $2\,HI(aq) + Ba(OH)_2(aq) \longrightarrow 2\,H_2O(l) + BaI_2(aq)$
Write the net ionic equation by removing the spectator ions.	**Net ionic equation:** $2\,H^+(aq) + 2\,OH^-(aq) \longrightarrow 2\,H_2O(l)$ or simply $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

FOR PRACTICE 5.9 Write a molecular and a net ionic equation for the reaction that occurs between aqueous HBr and aqueous LiOH.

EXAMPLE 5.10 Writing Equations for Acid–Base Reactions Involving a Weak Acid

Write a molecular equation, ionic equation, and net ionic equation for the reaction between aqueous acetic acid ($HC_2H_3O_2$) and aqueous potassium hydroxide (KOH).

SOLUTION Begin by writing the molecular equation in which the acid and the base combine to form water and a salt. (The equation is already balanced.)	**Molecular equation:** $HC_2H_3O_2(aq) + KOH(aq) \longrightarrow H_2O(l) + KC_2H_3O_2(aq)$
Write the complete ionic equation by separating aqueous ionic compounds into their constituent ions. Do not separate $HC_2H_3O_2(aq)$ because it is a weak acid (and a weak electrolyte).	**Complete ionic equation:** $HC_2H_3O_2(aq) + K^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + K^+(aq) + C_2H_3O_2^-(aq)$
Write the net ionic equation by eliminating the spectator ions.	**Net ionic equation:** $HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$

FOR PRACTICE 5.10 Write the net ionic equation for the reaction between $HCHO_2$ (a weak acid) and NaOH.

Acid–Base Titrations

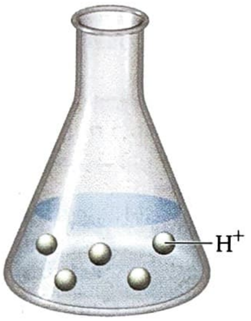
We can apply the principles of acid–base neutralization and stoichiometry to a common laboratory procedure called a *titration*. In a **titration**, a substance in a solution of known concentration is reacted with another substance in a solution of unknown concentration. For example, consider the following acid–base reaction:

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

The net ionic equation for this reaction eliminates the spectator ions:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Suppose we have an HCl solution represented by the molecular diagram shown here (for purposes of clarity, we have omitted the Cl^- ions and the H_2O molecules not involved in the reaction from this representation).



In titrating this sample, we slowly add a solution of known OH^- concentration, as shown in the molecular diagrams in Figure 5.17▼. As we add the OH^-, it reacts with and neutralizes the H^+, forming water. At the **equivalence point**—the point in the titration when the number of moles of OH^- added equals the number of moles of H^+ initially in solution—the titration is complete. The equivalence point is typically signaled by an **indicator**, a dye whose color depends on the acidity or basicity of the solution (Figure 5.18▶).

We cover acid–base titrations and indicators in more detail in Chapter 18. In most laboratory titrations, the concentration of one of the reactant solutions is unknown, and the concentration of the other is precisely known. By carefully measuring the volume of each solution required to reach the equivalence point, we can determine the concentration of the unknown solution, as demonstrated in Example 5.11.

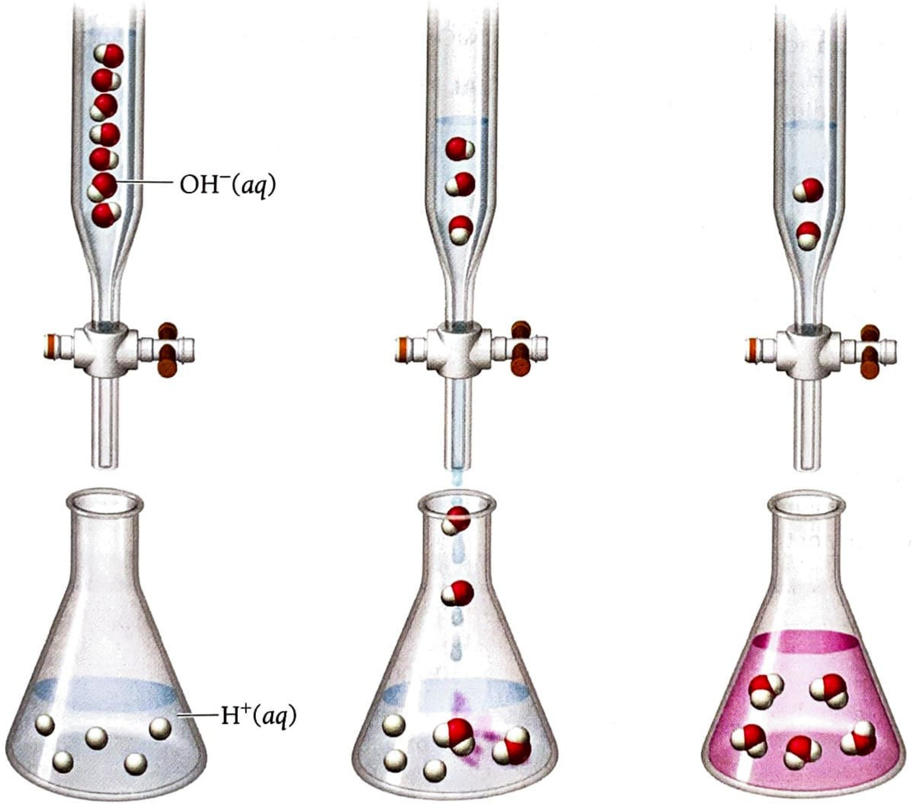
Acid–Base Titration

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Beginning of titration:
Known OH^- concentration
Unknown H^+ concentration

OH^- solution is slowly added to acid solution

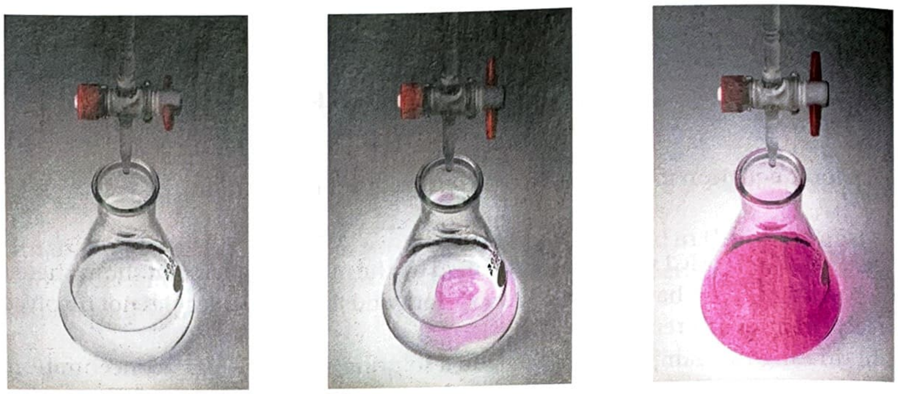
Equivalence point:
Moles of OH^- = moles of H^+



◀ **FIGURE 5.17 Acid–Base Titration**

▶ FIGURE 5.18 Titration In this titration, NaOH is added to a dilute HCl solution. When the NaOH and HCl reach stoichiometric proportions (the equivalence point), the phenolphthalein indicator changes color to pink.

Indicator in Titration



ANSWER **NOW!**

5.9

Cc

Conceptual Connection

ACID–BASE TITRATION A 10.0 mL sample of 0.20 M HBr solution is titrated with 0.10 M NaOH. What volume of NaOH is required to reach the equivalence point?

(a) 10.0 mL **(b)** 20.0 mL **(c)** 40.0 mL

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE **5.11**

EXAMPLE 5.11 Acid–Base Titration

The titration of 10.00 mL of HCl solution of unknown concentration requires 12.54 mL of a 0.100 M NaOH solution to reach the equivalence point. What is the concentration of the unknown HCl solution in M?

SORT You are given the volume and concentration of NaOH solution required to titrate a given volume of HCl solution. You are asked to find the concentration of the HCl solution.

GIVEN: 12.54 mL of NaOH solution, 0.100 M NaOH solution, 10.00 mL of HCl solution

FIND: concentration of HCl solution

STRATEGIZE Since this problem involves an acid–base neutralization reaction between HCl and NaOH, start by writing the balanced equation, using the techniques covered earlier in this section.

The first part of the conceptual plan has the form volume A ⟶ moles A ⟶ moles B. The concentration of the NaOH solution is a conversion factor between moles and volume of NaOH. The balanced equation provides the relationship between number of moles of NaOH and number of moles of HCl.

In the second part of the conceptual plan, use the number of moles of HCl (from the first part) and the volume of HCl solution (given) to calculate the molarity of the HCl solution.

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

CONCEPTUAL PLAN

$$\boxed{\text{mL NaOH}} \xrightarrow[\frac{1\ \text{L}}{1000\ \text{mL}}]{} \boxed{\text{L NaOH}} \xrightarrow[\frac{0.100\ \text{mol NaOH}}{\text{L NaOH}}]{}$$

$$\boxed{\text{mol NaOH}} \xrightarrow[\frac{1\ \text{mol HCl}}{1\ \text{mol NaOH}}]{} \boxed{\text{mol HCl}}$$

$$\boxed{\text{mol HCl, L HCl solution}} \xrightarrow[M = \frac{\text{mol}}{\text{L}}]{} \boxed{\text{Molarity}}$$

RELATIONSHIPS USED

1 L = 1000 mL

$$\text{M (NaOH)} = \frac{0.100\ \text{mol NaOH}}{\text{L NaOH}}$$

1 mol HCl : 1 mol NaOH

$$\text{Molarity (M)} = \frac{\text{moles of solute (mol)}}{\text{volume of solution (L)}}$$

SOLVE In the first part of the solution, determine the number of moles of HCl in the unknown solution.

SOLUTION

$$12.54\ \cancel{\text{mL NaOH}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.100\ \cancel{\text{mol NaOH}}}{\cancel{\text{L}}\ \cancel{\text{NaOH}}} \times \frac{1\ \text{mol HCl}}{1\ \cancel{\text{mol NaOH}}} = 1.25 \times 10^{-3}\ \text{mol HCl}$$

In the second part of the solution, divide the number of moles of HCl by the volume of the HCl solution in L. 10.00 mL is equivalent to 0.01000 L.

$$\text{Molarity} = \frac{1.25 \times 10^{-3}\ \text{mol HCl}}{0.01000\ \text{L}} = 0.125\ \text{M HCl}$$

CHECK The units of the answer (M HCl) are correct. The magnitude of the answer (0.125 M) is reasonable because it is similar to the molarity of the NaOH solution, as expected from the reaction stoichiometry (1 mol HCl reacts with 1 mol NaOH) and the similar volumes of NaOH and HCl.

FOR PRACTICE 5.11 The titration of a 20.0-mL sample of an H_2SO_4 solution of unknown concentration requires 22.87 mL of a 0.158 M KOH solution to reach the equivalence point. What is the concentration of the unknown H_2SO_4 solution?

FOR MORE PRACTICE 5.11 What volume (in mL) of 0.200 M NaOH do we need to titrate 35.00 mL of 0.140 M HBr to the equivalence point?

5.8 Gas-Evolution Reactions

In a *gas-evolution reaction*, two aqueous solutions mix to form a gaseous product that bubbles out of solution. Some gas-evolution reactions form a gaseous product directly when the cation of one reactant combines with the anion of the other. For example, when sulfuric acid reacts with lithium sulfide, dihydrogen sulfide gas forms:

$$H_2SO_4(aq) + Li_2S(aq) \longrightarrow \underset{\text{gas}}{H_2S(g)} + Li_2SO_4(aq)$$

Other gas-evolution reactions often form an intermediate product that then decomposes (breaks down into simpler substances) to form a gas. For example, when aqueous hydrochloric acid is mixed with aqueous sodium bicarbonate, the following reaction occurs (Figure 5.19▶):

$$HCl(aq) + NaHCO_3(aq) \longrightarrow \underset{\text{intermediate product}}{H_2CO_3(aq)} + NaCl(aq) \longrightarrow H_2O(l) + \underset{\text{gas}}{CO_2(g)} + NaCl(aq)$$

Many gas-evolution reactions such as this one are also acid–base reactions. In Chapter 17 we will learn how ions such as CO_3^{2-} act as bases in aqueous solution.

The intermediate product, H_2CO_3, is not stable and decomposes into H_2O and gaseous CO_2. Other important gas-evolution reactions form either H_2SO_3 or NH_4OH as intermediate products:

$$HCl(aq) + NaHSO_3(aq) \longrightarrow \underset{\text{intermediate product}}{H_2SO_3(aq)} + NaCl(aq) \longrightarrow H_2O(l) + \underset{\text{gas}}{SO_2(g)} + NaCl(aq)$$

$$NH_4Cl(aq) + NaOH(aq) \longrightarrow \underset{\text{intermediate product}}{NH_4OH(aq)} + NaCl(aq) \longrightarrow H_2O(l) + \underset{\text{gas}}{NH_3(g)} + NaCl(aq)$$

The intermediate product NH_4OH provides a convenient way to think about this reaction, but the extent to which it actually forms is debatable.

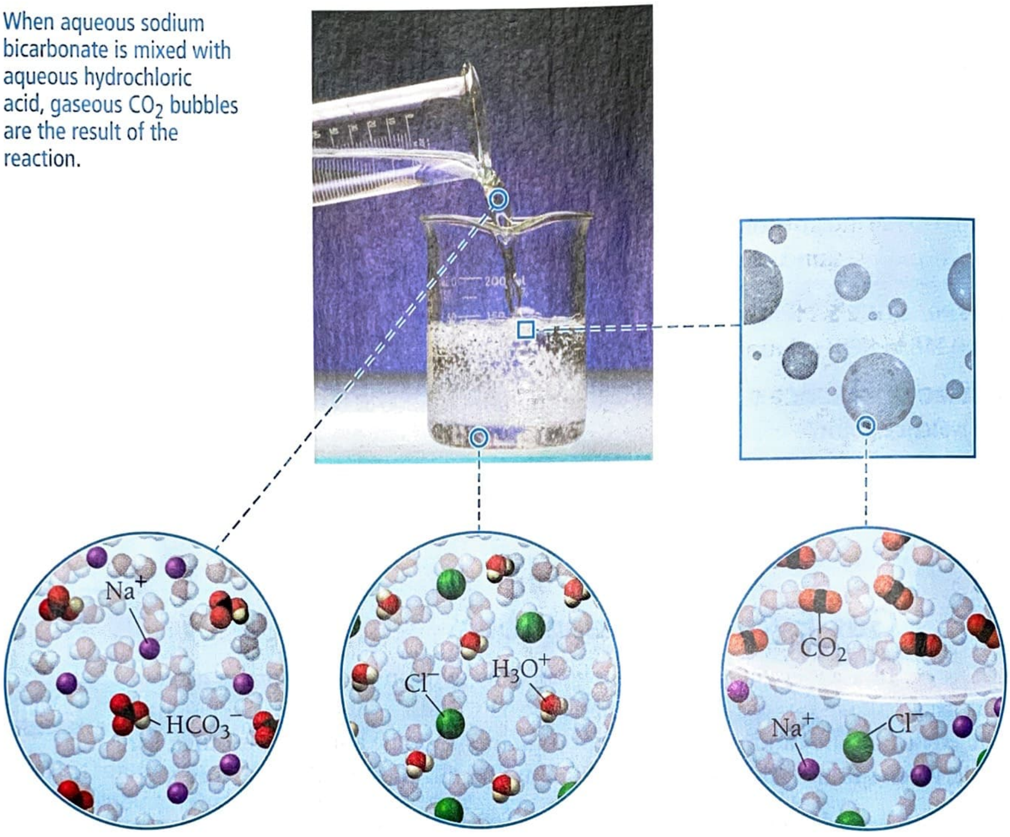
Table 5.3 lists the main types of compounds that form gases in aqueous reactions, as well as the gases formed.

TABLE 5.3 ■ Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H_2S	$2\ HCl(aq) + K_2S(aq) \longrightarrow H_2S(g) + 2\ KCl(aq)$
Carbonates and bicarbonates	H_2CO_3	CO_2	$2\ HCl(aq) + K_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\ KCl(aq)$
Sulfites and bisulfites	H_2SO_3	SO_2	$2\ HCl(aq) + K_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\ KCl(aq)$
Ammonium	NH_4OH	NH_3	$NH_4Cl(aq) + KOH(aq) \longrightarrow H_2O(l) + NH_3(g) + KCl(aq)$

Gas-Evolution Reaction

$$NaHCO_3(aq) + HCl(aq) \longrightarrow H_2O(l) + NaCl(aq) + CO_2(g)$$



When aqueous sodium bicarbonate is mixed with aqueous hydrochloric acid, gaseous CO_2 bubbles are the result of the reaction.

$$NaHCO_3(aq) \quad + \quad HCl(aq) \longrightarrow H_2O(l) + NaCl(aq) + CO_2(g)$$

▲ **FIGURE 5.19** Gas-Evolution Reaction

EXAMPLE 5.12 Writing Equations for Gas-Evolution Reactions

Write a molecular equation for the gas-evolution reaction that occurs when you mix aqueous nitric acid and aqueous sodium carbonate.

Begin by writing an unbalanced equation in which the cation of each reactant combines with the anion of the other.	$HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2CO_3(aq) + NaNO_3(aq)$
You must then recognize that $H_2CO_3(aq)$ decomposes into $H_2O(l)$ and $CO_2(g)$ and write these products into the equation.	$HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + NaNO_3(aq)$
Finally, balance the equation.	$2\,HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\,NaNO_3(aq)$

FOR PRACTICE 5.12 Write a molecular equation for the gas-evolution reaction that occurs when you mix aqueous hydrobromic acid and aqueous potassium sulfite.

FOR MORE PRACTICE 5.12 Write a net ionic equation for the reaction that occurs when you mix hydroiodic acid with calcium sulfide.

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. What is an aqueous solution? What is the difference between the solute and the solvent?

2. What is molarity? How is it useful?

3. Explain how a strong electrolyte, a weak electrolyte, and a nonelectrolyte differ.

4. Explain the difference between a strong acid and a weak acid.

5. What does it mean for a compound to be soluble? Insoluble?

6. What are the solubility rules? How are they useful?

7. What are the cations and anions whose compounds are usually soluble? What are the exceptions? What are the anions whose compounds are mostly insoluble? What are the exceptions?

8. What is a precipitation reaction? Give an example.

9. How can you predict whether a precipitation reaction will occur upon mixing two aqueous solutions?

10. Explain how a molecular equation, a complete ionic equation, and a net ionic equation differ.

11. What is the Arrhenius definition of an acid? A base?

12. What is an acid–base reaction? Give an example.

13. Explain the principles behind an acid–base titration. What is an indicator?

14. What is a gas-evolution reaction? Give an example.

15. What reactant types give rise to gas-evolution reactions?

16. What is an oxidation–reduction reaction? Give an example.

17. What are oxidation states?

18. How can oxidation states be used to identify redox reactions?

19. What happens to a substance when it becomes oxidized? Reduced?

20. In a redox reaction, which reactant is the oxidizing agent? The reducing agent?

PROBLEMS BY TOPIC

Solution Concentration and Solution Stoichiometry

21. Calculate the molarity of each solution.
MISSED THIS? *Read Section 5.2; Watch KCV 5.2, IWE 5.1*

- **a.** 3.25 mol of LiCl in 2.78 L solution
- **b.** 28.33 g $C_6H_{12}O_6$ in 1.28 L of solution
- **c.** 32.4 mg NaCl in 122.4 mL of solution

22. Calculate the molarity of each solution.

- **a.** 0.38 mol of $LiNO_3$ in 6.14 L of solution
- **b.** 72.8 g C_2H_6O in 2.34 L of solution
- **c.** 12.87 mg KI in 112.4 mL of solution

23. What is the molarity of NO_3^- in each solution?
MISSED THIS? *Read Sections 5.2, 5.4; Watch KCV 5.2, IWE 5.1*

- **a.** 0.150 M KNO_3
- **b.** 0.150 M $Ca(NO_3)_2$
- **c.** 0.150 M $Al(NO_3)_3$

24. What is the molarity of Cl^- in each solution?

- **a.** 0.200 M NaCl
- **b.** 0.150 M $SrCl_2$
- **c.** 0.100 M $AlCl_3$

25. How many moles of KCl are contained in each solution?
MISSED THIS? *Read Section 5.2; Watch KCV 5.2, IWE 5.2*

- **a.** 0.556 L of a 2.3 M KCl solution
- **b.** 1.8 L of a 0.85 M KCl solution
- **c.** 114 mL of a 1.85 M KCl solution

26. What volume of 0.200 M ethanol solution contains each amount in moles of ethanol?

- **a.** 0.45 mol ethanol
- **b.** 1.22 mol ethanol
- **c.** 1.2×10^{-2} mol ethanol

27. A laboratory procedure calls for making 400.0 mL of a 1.1 M $NaNO_3$ solution. What mass of $NaNO_3$ (in g) is needed?
MISSED THIS? *Read Section 5.2; Watch KCV 5.2, IWE 5.2*

28. A chemist wants to make 5.5 L of a 0.300 M $CaCl_2$ solution. What mass of $CaCl_2$ (in g) should the chemist use?

29. If 123 mL of a 1.1 M glucose solution is diluted to 500.0 mL, what is the molarity of the diluted solution?
MISSED THIS? *Read Section 5.2; Watch KCV 5.2, IWE 5.3*

30. If 3.5 L of a 4.8 M $SrCl_2$ solution is diluted to 45 L, what is the molarity of the diluted solution?

31. To what volume should you dilute 50.0 mL of a 12 M stock HNO_3 solution to obtain a 0.100 M HNO_3 solution?
MISSED THIS? *Read Section 5.2; Watch KCV 5.2, IWE 5.3*

32. To what volume should you dilute 25 mL of a 10.0 M H_2SO_4 solution to obtain a 0.150 M H_2SO_4 solution?

33. Consider the precipitation reaction:
MISSED THIS? *Read Section 5.3; Watch IWE 5.4*

$$2\,Na_3PO_4(aq) + 3\,CuCl_2(aq) \longrightarrow Cu_3(PO_4)_2(s) + 6\,NaCl(aq)$$

What volume of 0.175 M Na_3PO_4 solution is necessary to completely react with 95.4 mL of 0.102 M $CuCl_2$?

34. Consider the reaction:

$$Li_2S(aq) + Co(NO_3)_2(aq) \longrightarrow 2\,LiNO_3(aq) + CoS(s)$$

What volume of 0.150 M Li_2S solution is required to completely react with 125 mL of 0.150 M $Co(NO_3)_2$?

35. What is the minimum amount of 6.0 M H_2SO_4 necessary to produce 25.0 g of $H_2(g)$ according to the reaction between aluminum and sulfuric acid?
MISSED THIS? *Read Section 5.3; Watch IWE 5.4*

$$2\,Al(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3\,H_2(g)$$

36. What is the molarity of $ZnCl_2$ that forms when 25.0 g of zinc completely reacts with $CuCl_2$ according to the following reaction? Assume a final volume of 275 mL.

$$Zn(s) + CuCl_2(aq) \longrightarrow ZnCl_2(aq) + Cu(s)$$

37. A 25.0-mL sample of a 1.20 M potassium chloride solution is mixed with 15.0 mL of a 0.900 M lead(II) nitrate solution, and this precipitation reaction occurs:
MISSED THIS? *Read Section 4.3, 5.3; Watch KCV 4.3, IWE 4.6, 5.4*

$$2\,KCl(aq) + Pb(NO_3)_2(aq) \longrightarrow PbCl_2(s) + 2\,KNO_3(aq)$$

The solid $PbCl_2$ is collected, dried, and found to have a mass of 2.45 g. Determine the limiting reactant, the theoretical yield, and the percent yield.

38. A 55.0-mL sample of a 0.102 M potassium sulfate solution is mixed with 35.0 mL of a 0.114 M lead(II) acetate solution and this precipitation reaction occurs:

$$K_2SO_4(aq) + Pb(C_2H_3O_2)_2(aq) \longrightarrow 2\,KC_2H_3O_2(aq) + PbSO_4(s)$$

The solid $PbSO_4$ is collected, dried, and found to have a mass of 1.01 g. Determine the limiting reactant, the theoretical yield, and the percent yield.

Types of Aqueous Solutions and Solubility

39. For each compound (all water soluble), would you expect the resulting aqueous solution to conduct electrical current?
MISSED THIS? *Read Section 5.4*

a. $CsCl$ **b.** CH_3OH **c.** $Ca(NO_2)_2$ **d.** $C_6H_{12}O_6$

40. Classify each compound as a strong electrolyte or nonelectrolyte.

a. $MgBr_2$ **b.** $C_{12}H_{22}O_{11}$ **c.** Na_2CO_3 **d.** KOH

41. Determine whether each compound is soluble or insoluble. If the compound is soluble, list the ions present in solution.
MISSED THIS? *Read Section 5.4; Watch IWE 5.5*

a. $AgNO_3$ **b.** $Pb(C_2H_3O_2)_2$
c. KNO_3 **d.** $(NH_4)_2S$

42. Determine whether each compound is soluble or insoluble. If the compound is soluble, list the ions present in solution.

a. AgI **b.** $Cu_3(PO_4)_2$
c. $CoCO_3$ **d.** K_3PO_4

Precipitation Reactions

43. Complete and balance each equation. If no reaction occurs, write “NO REACTION.”
MISSED THIS? *Read Section 5.5; Watch KCV 5.5, IWE 5.6*

a. $LiI(aq) + BaS(aq) \longrightarrow$
b. $KCl(aq) + CaS(aq) \longrightarrow$
c. $CrBr_2(aq) + Na_2CO_3(aq) \longrightarrow$
d. $NaOH(aq) + FeCl_3(aq) \longrightarrow$

44. Complete and balance each equation. If no reaction occurs, write “NO REACTION.”

a. $NaNO_3(aq) + KCl(aq) \longrightarrow$
b. $NaCl(aq) + Hg_2(C_2H_3O_2)_2(aq) \longrightarrow$
c. $(NH_4)_2SO_4(aq) + SrCl_2(aq) \longrightarrow$
d. $NH_4Cl(aq) + AgNO_3(aq) \longrightarrow$

45. Write a molecular equation for the precipitation reaction that occurs (if any) when each pair of aqueous solutions is mixed. If no reaction occurs, write “NO REACTION.”
MISSED THIS? *Read Section 5.5; Watch KCV 5.5, IWE 5.6*

a. potassium carbonate and lead(II) nitrate
b. lithium sulfate and lead(II) acetate
c. copper(II) nitrate and magnesium sulfide
d. strontium nitrate and potassium iodide

46. Write a molecular equation for the precipitation reaction that occurs (if any) when each pair of aqueous solutions is mixed. If no reaction occurs, write “NO REACTION.”

a. sodium chloride and lead(II) acetate
b. potassium sulfate and strontium iodide
c. cesium chloride and calcium sulfide
d. chromium(III) nitrate and sodium phosphate

Ionic and Net Ionic Equations

47. Write balanced complete ionic and net ionic equations for each reaction. **MISSED THIS?** *Read Section 5.6*

a. $HCl(aq) + LiOH(aq) \longrightarrow H_2O(l) + LiCl(aq)$
b. $CaS(aq) + CuCl_2(aq) \longrightarrow CuS(s) + CaCl_2(aq)$
c. $NaOH(aq) + HC_2H_3O_2(aq) \longrightarrow H_2O(l) + NaC_2H_3O_2(aq)$
d. $Na_3PO_4(aq) + NiCl_2(aq) \longrightarrow Ni_3(PO_4)_2(s) + NaCl(aq)$

48. Write balanced complete ionic and net ionic equations for each reaction.

a. $K_2SO_4(aq) + CaI_2(aq) \longrightarrow CaSO_4(s) + KI(aq)$
b. $NH_4Cl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NH_3(g) + NaCl(aq)$
c. $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$
d. $HC_2H_3O_2(aq) + K_2CO_3(aq) \longrightarrow$
$H_2O(l) + CO_2(g) + KC_2H_3O_2(aq)$

49. Mercury(I) ions (Hg_2^{2+}) can be removed from solution by precipitation with Cl^-. Suppose that a solution contains aqueous $Hg_2(NO_3)_2$. Write complete ionic and net ionic equations for the reaction of aqueous $Hg_2(NO_3)_2$ with aqueous sodium chloride to form solid Hg_2Cl_2 and aqueous sodium nitrate.
MISSED THIS? *Read Section 5.6*

50. Lead(II) ions can be removed from solution by precipitation with sulfate ions. Suppose that a solution contains lead(II) nitrate. Write complete ionic and net ionic equations for the reaction of aqueous lead(II) nitrate with aqueous potassium sulfate to form solid lead(II) sulfate and aqueous potassium nitrate.

Acid–Base and Gas-Evolution Reactions

51. Write balanced molecular and net ionic equations for the reaction between hydrobromic acid and potassium hydroxide.
MISSED THIS? *Read Section 5.7; Watch KCV 5.5, IWE 5.9*

52. Write balanced molecular and net ionic equations for the reaction between nitric acid and calcium hydroxide.

53. Complete and balance each acid–base equation.
MISSED THIS? *Read Section 5.7; Watch IWE 5.9*

a. $H_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow$
b. $HClO_4(aq) + KOH(aq) \longrightarrow$
c. $H_2SO_4(aq) + NaOH(aq) \longrightarrow$

54. Complete and balance each acid–base equation.

a. $HI(aq) + LiOH(aq) \longrightarrow$
b. $HC_2H_3O_2(aq) + Ca(OH)_2(aq) \longrightarrow$
c. $HCl(aq) + Ba(OH)_2(aq) \longrightarrow$

55. Write balanced complete ionic and net ionic equations for each acid–base reaction.
MISSED THIS? *Read Section 5.7; Watch KCV 5.5, IWE 5.9*

a. $HBr(aq) + NaOH(aq) \longrightarrow$
b. $HF(aq) + NaOH(aq) \longrightarrow$
c. $HC_2H_3O_2(aq) + RbOH(aq) \longrightarrow$

56. Write balanced complete ionic and net ionic equations for each acid–base reaction.

a. $HI(aq) + RbOH(aq) \longrightarrow$
b. $HCHO_2(aq) + NaOH(aq) \longrightarrow$
c. $HC_2H_3O_2(aq) + LiOH(aq) \longrightarrow$

57. A 25.00-mL sample of an unknown $HClO_4$ solution requires titration with 22.62 mL of 0.2000 M NaOH to reach the equivalence point. What is the concentration of the unknown $HClO_4$ solution? The neutralization reaction is
MISSED THIS? *Read Section 5.7; Watch IWE 5.11*

$$HClO_4(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaClO_4(aq)$$

58. A 30.00-mL sample of an unknown H_3PO_4 solution is titrated with a 0.100 M NaOH solution. The equivalence point is reached when 26.38 mL of NaOH solution is added. What is the concentration of the unknown H_3PO_4 solution? The neutralization reaction is

$$H_3PO_4(aq) + 3\,NaOH(aq) \longrightarrow 3\,H_2O(l) + Na_3PO_4(aq)$$

59. Complete and balance each gas-evolution equation.
MISSED THIS? *Read Section 5.8; Watch KCV 5.5*
a. $HBr(aq) + NiS(s) \longrightarrow$
b. $NH_4I(aq) + NaOH(aq) \longrightarrow$
c. $HBr(aq) + Na_2S(aq) \longrightarrow$
d. $HClO_4(aq) + Li_2CO_3(aq) \longrightarrow$

60. Complete and balance each gas-evolution equation.
a. $HNO_3(aq) + Na_2SO_3(aq) \longrightarrow$
b. $HCl(aq) + KHCO_3(aq) \longrightarrow$
c. $HC_2H_3O_2(aq) + NaHSO_3(aq) \longrightarrow$
d. $(NH_4)_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow$

Oxidation–Reduction

61. Assign oxidation states to each atom in each element, ion, or compound. **MISSED THIS?** *Read Section 5.9; Watch IWE 5.13*
a. Ag **b.** Ag^+ **c.** CaF_2
d. H_2S **e.** CO_3^{2-} **f.** CrO_4^{2-}

62. Assign oxidation states to each atom in each element, ion, or compound.
a. Cl_2 **b.** Fe^{3+} **c.** $CuCl_2$
d. CH_4 **e.** $Cr_2O_7^{2-}$ **f.** HSO_4^-

63. What is the oxidation state of Cr in each compound?
MISSED THIS? *Read Section 5.9; Watch IWE 5.13*
a. CrO **b.** CrO_3 **c.** Cr_2O_3

64. What is the oxidation state of Cl in each ion?
a. ClO^- **b.** ClO_2^- **c.** ClO_3^- **d.** ClO_4^-

65. Determine whether each reaction is a redox reaction. For each redox reaction, identify the oxidizing agent and the reducing agent. **MISSED THIS?** *Read Section 5.9*
a. $4\,Li(s) + O_2(g) \longrightarrow 2\,Li_2O(s)$
b. $Mg(s) + Fe^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Fe(s)$
c. $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow PbSO_4(s) + 2\,NaNO_3(aq)$
d. $HBr(aq) + KOH(aq) \longrightarrow H_2O(l) + KBr(aq)$

66. Determine whether each reaction is a redox reaction. For each redox reaction, identify the oxidizing agent and the reducing agent.
a. $Al(s) + 3\,Ag^+(aq) \longrightarrow Al^{3+}(aq) + 3\,Ag(s)$
b. $SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$
c. $Ba(s) + Cl_2(g) \longrightarrow BaCl_2(s)$
d. $Mg(s) + Br_2(l) \longrightarrow MgBr_2(s)$

67. Determine whether each redox reaction occurs spontaneously in the forward direction. **MISSED THIS?** *Read Section 5.9*
a. $Ni(s) + Zn^{2+}(aq) \longrightarrow Ni^{2+}(aq) + Zn(s)$
b. $Ni(s) + Pb^{2+}(aq) \longrightarrow Ni^{2+}(aq) + Pb(s)$
c. $Al(s) + 3\,Ag^+(aq) \longrightarrow Al^{3+}(aq) + 3\,Ag(s)$
d. $Pb(s) + Mn^{2+}(aq) \longrightarrow Pb^{2+}(aq) + Mn(s)$

68. Determine whether each redox reaction occurs spontaneously in the forward direction.
a. $Ca^{2+}(aq) + Zn(s) \longrightarrow Ca(s) + Zn^{2+}(aq)$
b. $2\,Ag^+(aq) + Ni(s) \longrightarrow 2\,Ag(s) + Ni^{2+}(aq)$
c. $Fe(s) + Mn^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Mn(s)$
d. $2\,Al(s) + 3\,Pb^{2+}(aq) \longrightarrow 2\,Al^{3+}(aq) + 3\,Pb(s)$

69. Suppose you wanted to cause Ni^{2+} ions to come out of solution as solid Ni. Which metal could you use to accomplish this?
MISSED THIS? *Read Section 5.9*

70. Suppose you wanted to cause Pb^{2+} ions to come out of solution as solid Pb. Which metal could you use to accomplish this?

71. Which metal in the activity series reduces Al^{3+} ions but not Na^+ ions? **MISSED THIS?** *Read Section 5.9*

72. Which metal in the activity series is oxidized with a Ni^{2+} solution but not with a Cr^{3+} solution?

CUMULATIVE PROBLEMS

73. The density of a 20.0% by mass ethylene glycol ($C_2H_6O_2$) solution in water is 1.03 g/mL. Find the molarity of the solution.

74. Find the percent by mass of sodium chloride in a 1.35 M NaCl solution. The density of the solution is 1.05 g/mL.

75. People sometimes use sodium bicarbonate as an antacid to neutralize excess hydrochloric acid in an upset stomach. What mass of hydrochloric acid (in grams) can 2.5 g of sodium bicarbonate neutralize? (*Hint*: Begin by writing a balanced equation for the reaction between aqueous sodium bicarbonate and aqueous hydrochloric acid.)

76. Toilet bowl cleaners often contain hydrochloric acid, which dissolves the calcium carbonate deposits that accumulate within a toilet bowl. What mass of calcium carbonate (in grams) can 3.8 g of HCl dissolve? (*Hint*: Begin by writing a balanced equation for the reaction between hydrochloric acid and calcium carbonate.)

77. Predict the products and write a balanced molecular equation for each reaction. If no reaction occurs, write "NO REACTION."
a. $HCl(aq) + Hg_2(NO_3)_2(aq) \longrightarrow$
b. $KHSO_3(aq) + HNO_3(aq) \longrightarrow$
c. aqueous ammonium chloride and aqueous lead(II) nitrate
d. aqueous ammonium chloride and aqueous calcium hydroxide

78. Predict the products and write a balanced molecular equation for each reaction. If no reaction occurs, write "NO REACTION."
a. $H_2SO_4(aq) + HNO_3(aq) \longrightarrow$
b. $Cr(NO_3)_3(aq) + LiOH(aq) \longrightarrow$
c. aqueous strontium sulfide and aqueous copper(II) sulfate

79. Hard water often contains dissolved Ca^{2+} and Mg^{2+} ions. One way to soften water is to add phosphates. The phosphate ion forms insoluble precipitates with calcium and magnesium ions, removing them from solution. A solution is 0.050 M in calcium chloride and 0.085 M in magnesium nitrate. What mass of sodium phosphate would you add to 1.5 L of this solution to completely eliminate the hard water ions? Assume complete reaction.

80. An acid solution is 0.100 M in HCl and 0.200 M in H_2SO_4. What volume of a 0.150 M KOH solution would completely neutralize all the acid in 500.0 mL of this solution?

81. Find the mass of barium metal (in grams) that must react with O_2 to produce enough barium oxide to prepare 1.0 L of a 0.10 M solution of OH^-. (*Hint*: Barium metal reacts with oxygen to form BaO; BaO reacts with water to form $Ba(OH)_2$.)

82. A solution contains Cr^{3+} ions and Mg^{2+} ions. The addition of 1.00 L of 1.51 M NaF solution causes the complete precipitation of these ions as $CrF_3(s)$ and $MgF_2(s)$. The total mass of the precipitate is 49.6 g. Find the mass of Cr^{3+} in the original solution.

83. Treatment of gold metal with BrF_3 and KF produces Br_2 and $KAuF_4$, a salt of gold. Identify the oxidizing agent and the reducing agent in this reaction. Find the mass of the gold salt that forms when a 73.5-g mixture of equal masses of all three reactants is prepared.

84. We prepare a solution by mixing 0.10 L of 0.12 M sodium chloride with 0.23 L of a 0.18 M $MgCl_2$ solution. What volume of a 0.20 M silver nitrate solution do we need to precipitate all the Cl^- ion in the solution as AgCl?

85. A solution contains one or more of the following ions: Ag^+, Ca^{2+}, and Cu^{2+}. When you add sodium chloride to the solution, no precipitate forms. When you add sodium sulfate to the solution, a white precipitate forms. You filter off the precipitate and add sodium carbonate to the remaining solution, producing another precipitate. Which ions were present in the original solution? Write net ionic equations for the formation of each of the precipitates observed.

86. A solution contains one or more of the following ions: Hg_2^{2+}, Ba^{2+}, and Fe^{2+}. When you add potassium chloride to the solution, a precipitate forms. The precipitate is filtered off, and you add potassium sulfate to the remaining solution, producing no precipitate. When you add potassium carbonate to the remaining solution, a precipitate forms. Which ions were present in the original solution? Write net ionic equations for the formation of each of the precipitates observed.

CHALLENGE PROBLEMS

87. A solution contains Ag^+ and Hg^{2+} ions. The addition of 0.100 L of 1.22 M NaI solution is just enough to precipitate all the ions as AgI and HgI_2. The total mass of the precipitate is 28.1 g. Find the mass of AgI in the precipitate.

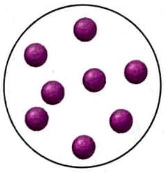
88. The water in lakes that have been acidified by acid rain (HNO_3 and H_2SO_4) can be neutralized by a process called liming, in which limestone ($CaCO_3$) is added to the acidified water. What mass of limestone (in kg) would completely neutralize a 15.2 billion-liter lake that is 1.8×10^{-5} M in H_2SO_4 and 8.7×10^{-6} M in HNO_3?

89. Sodium carbonate is often added to laundry detergents to soften hard water and make the detergent more effective. Suppose that a particular detergent mixture is designed to soften hard water that is 3.5×10^{-3} M in Ca^{2+} and 1.1×10^{-3} M in Mg^{2+} and that the average capacity of a washing machine is 19.5 gallons of water. If the detergent requires using 0.65 kg detergent per load of laundry, what percentage (by mass) of the detergent should be sodium carbonate in order to completely precipitate all of the calcium and magnesium ions in an average load of laundry water?

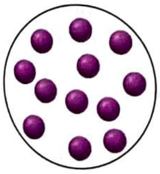
90. Lead poisoning is a serious condition resulting from the ingestion of lead in food, water, or other environmental sources. It affects the central nervous system, leading to a variety of symptoms such as distractibility, lethargy, and loss of motor coordination. Lead poisoning is treated with chelating agents, substances that bind to metal ions, allowing them to be eliminated in the urine. A modern chelating agent used for this purpose is succimer ($C_4H_6O_4S_2$). Suppose you are trying to determine the appropriate dose for succimer treatment of lead poisoning. What minimum mass of succimer (in mg) is needed to bind all of the lead in a patient's bloodstream? Assume that patient blood lead levels are 45 μg/dL, that total blood volume is 5.0 L, and that 1 mol of succimer binds 1 mol of lead.

CONCEPTUAL PROBLEMS

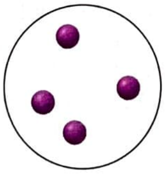
91. The following circle represents 1.0 liter of a solution with a solute concentration of 1 M:



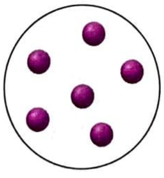
Explain what you would add (the amount of solute or volume of solvent) to the solution to obtain a solution represented by each diagram:



a.



b.

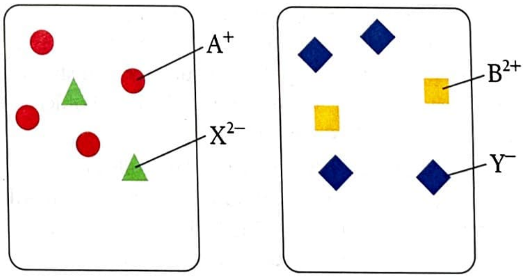


c.

92. Consider the generic ionic compounds with the formulas A_2X and BY_2 and the following solubility rules:

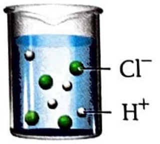
A_2X soluble; BY_2 soluble; AY insoluble; BX soluble.

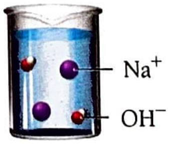
Assume A^+ ions are circles, B^{2+} ions are squares, X^{2-} ions are triangles, and Y^- ions are diamonds. Solutions of the two compounds (A_2X and BY_2) are represented as follows:



Draw a molecular-level representation showing the result of mixing the two given solutions and write an equation to represent the reaction.

93. A hydrochloric acid solution will neutralize a sodium hydroxide solution. Look at the molecular views showing one beaker of HCl and four beakers of NaOH. Which NaOH beaker will just neutralize the HCl beaker? Begin by writing a balanced chemical equation for the neutralization reaction.





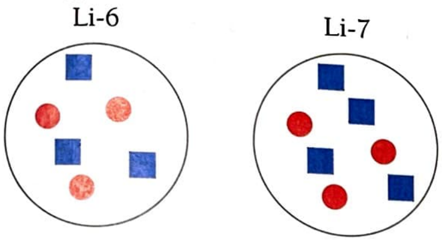
a.

b.

c.

d.

115. 3.56 cm
117. Li-6 = 7.494%, Li-7 = 92.506%
119. 75.0% gold
121. 106.91 amu
123. 1.66×10^{22} gold atoms
125. 1×10^{78} atoms/universe
127. 0.423
129. 63.67 g/mol
131. 25.06 g/mol
133.



135. If the amu and mole were not based on the same isotope, the numerical values obtained for an atom of material and a mole of material would not be the same. If, for example, the mole were based on the number of particles in C-12 but the amu were changed to a fraction of the mass of an atom of Ne-20, the number of particles and the number of amu that make up one mole of material would no longer be the same. We would no longer have the relationship in which the mass of an atom in amu is numerically equal to the mass of a mole of those atoms in grams.

137. The different isotopes of the same element have the same number of protons and electrons, so the attractive forces between the nucleus and the electrons are constant and there is no difference in the radii of the isotopes. Ions, on the other hand, have a different number of electrons than the parent atom from which they are derived. Cations have fewer electrons than the parent atom. The attractive forces are greater because there is a larger positive charge in the nucleus than the negative charge in the electron cloud. So, cations are smaller than the atom they are derived from. Anions have more electrons than the parent. The electron cloud has a greater negative charge than the nucleus, so the anions have larger radii than the parent.

142. **a.** 2000, 0.24 $\mu g/m^3$; 2016, 0.017 $\mu g/m^3$
c. 2.9×10^{14} Pb atoms

Chapter 3

23. **a.** 3 Mg, 2 P, 8 O **b.** 1 Ba, 2 Cl
c. 1 Fe, 2 N, 4 O **d.** 1 Ca, 2 O, 2 H
25. **a.** NH_3 **b.** C_2H_6 **c.** SO_3
27. **a.** atomic **b.** molecular
c. atomic **d.** molecular
29. **a.** molecular **b.** ionic
c. ionic **d.** molecular
31. **a.** molecular element **b.** molecular compound
c. atomic element
33. **a.** CaO **b.** ZnS **c.** RbBr **d.** Al_2O_3
35. **a.** $Ca(OH)_2$ **b.** $CaCrO_4$
c. $Ca_3(PO_4)_2$ **d.** $Ca(CN)_2$
37. **a.** magnesium nitride **b.** potassium fluoride
c. sodium oxide **d.** lithium sulfide
e. cesium fluoride **f.** potassium iodide
39. **a.** tin(II) oxide **b.** chromium(III) sulfide
c. rubidium iodide **d.** barium bromide
41. **a.** copper(I) nitrite **b.** magnesium acetate
c. barium nitrate **d.** lead(II) acetate
43. **a.** $NaHSO_3$ **b.** $LiMnO_4$
c. $AgNO_3$ **d.** K_2SO_4
e. $RbHSO_4$ **f.** $KHCO_3$
45. **a.** cobalt(II) sulfate heptahydrate
b. $IrBr_3 \times 4\,H_2O$
c. magnesium bromate hexahydrate
d. $K_2CO_3 \times 2\,H_2O$
47. **a.** carbon monoxide **b.** nitrogen triiodide
c. silicon tetrachloride
d. tetranitrogen tetraselenide
49. **a.** PCl_3 **b.** ClO
c. S_2F_4 **d.** PF_5
51. **a.** hydroiodic acid **b.** nitric acid
c. carbonic acid
53. **a.** HF **b.** HBr **c.** H_2SO_3
55. **a.** strontium chloride **b.** tin(IV) oxide
c. diphosphorus pentasulfide
d. acetic acid
57. **a.** potassium chlorate **b.** diiodine pentoxide
c. lead(II) sulfate
59. **a.** 46.01 amu **b.** 58.12 amu
c. 180.16 amu **d.** 238.03 amu
61. **a.** 0.471 mol **b.** 0.0362 mol
c. 968 mol **d.** 0.279 mol
63. **a.** 0.554 mol **b.** 28.4 mol
c. 0.378 mol **d.** 1093 mol
65. **a.** 2.2×10^{23} molecules **b.** 7.06×10^{23} molecules
c. 4.16×10^{23} molecules **d.** 1.09×10^{23} molecules
67. **a.** 0.0790 g **b.** 0.84 g **c.** 2.992×10^{-22} g
69. 0.10 mg
71. **a.** 74.87% C **b.** 79.88% C
c. 92.24% C **d.** 37.23% C
73. NH_3: 82.27% N
$CO(NH_2)_2$: 46.65% N
NH_4NO_3: 35.00% N
$(NH_4)_2SO_4$: 21.20% N
NH_3 has the highest N content.
75. 20.8 g F
77. 196 μg KI
79. **a.** 2 : 1 **b.** 4 : 1 **c.** 6 : 2 : 1
81. **a.** 0.885 mol H **b.** 5.2 mol H
c. 29 mol H **d.** 33.7 mol H
83. **a.** 3.3 g Na **b.** 3.6 g Na
c. 1.4 g Na **d.** 1.7 g Na
85. 1.41×10^{23} F atoms
87. **a.** Ag_2O **b.** $Co_3As_2O_8$ **c.** $SeBr_4$
89. **a.** C_5H_7N **b.** $C_4H_5N_2O$
91. $C_{13}H_{18}O_2$
93. NCl_3
95. **a.** $C_{12}H_{14}N_2$ **b.** $C_6H_3Cl_3$ **c.** $C_{10}H_{20}N_2S_4$

97. CH_2

99. C_2H_4O

101. a. inorganic **b.** organic
c. organic **d.** inorganic

103. a. alkene **b.** alkane
c. alkyne **d.** alkane

105. a. $CH_3CH_2CH_3$
b. propane
c. $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
d. pentane

107. a. functionalized hydrocarbon, alcohol
b. hydrocarbon
c. functionalized hydrocarbon, ketone
d. functionalized hydrocarbon, amine

109. 1.50×10^{24} molecules EtOH

111. a. K_2CrO_4, 40.27% K, 26.78% Cr, 32.95% O
b. $Pb_3(PO_4)_2$, 76.60% Pb, 7.63% P, 15.77% O
c. H_2SO_3, 2.46% H, 39.07% S, 58.47% O
d. $CoBr_2$, 26.94% Co, 73.06% Br

113. 1.80×10^2 g Cl_2/yr

115. M = Fe

117. estradiol = $C_{18}H_{24}O_2$

119. $C_{18}H_{20}O_2$

121. 7 H_2O

123. C_6H_9BrO

125. 1.87×10^{21} atoms

127. 92.93 amu

129. $x = 1, y = 2$

131. 41.7 mg

133. 0.224 g

135. 22.0% by mass

137. 1.6×10^7 kg Cl

139. 7.8×10^3 kg rock

141. $C_5H_{10}SI$

143. X_3Y_2

145. The sphere in the molecular models represents the electron cloud of the atom. On this scale, the nucleus would be too small to see.

147. The statement is incorrect because a chemical formula is based on the ratio of atoms combined, not the ratio of grams combined. The statement should read, "The chemical formula for ammonia (NH_3) indicates that ammonia contains three hydrogen atoms to each nitrogen atom."

149. O, S, H

154. a. Yes. **c.** 50.05%

Chapter 4

13. $2\,SO_2(g) + O_2(g) + 2\,H_2O(l) \longrightarrow 2\,H_2SO_4(aq)$

15. $2\,Na(s) + 2\,H_2O(l) \longrightarrow H_2(g) + 2\,NaOH(aq)$

17. $C_{12}H_{22}O_{11}(s) + H_2O(l) \longrightarrow 4\,C_2H_5OH(aq) + 4\,CO_2(g)$

19. a. $PbS(s) + 2\,HBr(aq) \longrightarrow PbBr_2(s) + H_2S(g)$
b. $CO(g) + 3\,H_2(g) \longrightarrow CH_4(g) + H_2O(l)$
c. $4\,HCl(aq) + MnO_2(g) \longrightarrow MnCl_2(aq) + 2\,H_2O(l) + Cl_2(g)$
d. $C_5H_{12}(l) + 8\,O_2(g) \longrightarrow 5\,CO_2(g) + 6\,H_2O(g)$

21. $Na_2CO_3(aq) + CuCl_2(aq) \longrightarrow CuCO_3(s) + 2\,NaCl(aq)$

23. a. $2\,CO_2(g) + CaSiO_3(s) + H_2O(l) \longrightarrow SiO_2(s) + Ca(HCO_3)_2(aq)$
b. $2\,Co(NO_3)_3(aq) + 3\,(NH_4)_2S(aq) \longrightarrow Co_2S_3(s) + 6\,NH_4NO_3(aq)$
c. $Cu_2O(s) + C(s) \longrightarrow 2\,Cu(s) + CO(g)$
d. $H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)$

25. $2\,C_6H_{14}(g) + 19\,O_2(g) \longrightarrow 12\,CO_2(g) + 14\,H_2O(g)$, 68 mol O_2

27. a. 5.0 mol NO_2 **b.** 14. mol NO_2
c. 0.281 mol NO_2 **d.** 53.1 mol NO_2

29.

mol SiO_2	mol C	mol SiC	mol CO
3	9	3	6
2	6	2	4
5	15	5	10
2.8	8.4	2.8	5.6
0.517	1.55	0.517	1.03

31. 9.3 g HBr, 0.12 g H_2

33. a. 5.56 g $BaCl_2$
b. 6.55 g $CaCO_3$
c. 6.09 g Mg O
d. 6.93 g Al_2O_3

35. a. Na **b.** Na **c.** Br_2 **d.** Na

37. 3 molecules Cl_2

39. a. 2 mol **b.** 7 mol **c.** 9.40 mol

41. 0.5 mol O_2

43. a. 2.5 g **b.** 31.1 g **c.** 1.16 g

45. 2.91 grams CO remaining

47. limiting reactant: Pb^{2+} theoretical yield: 34.5 g $PbCl_2$ percent yield: 85.3%

49. limiting reactant: NH_3 theoretical yield: 240.5 kg CH_4N_2O, percent yield: 70.01%

51. a. $S(s) + O_2(g) \longrightarrow SO_2(g)$
b. $2\,C_3H_6(g) + 9\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g)$
c. $2\,Ca(s) + O_2(g) \longrightarrow 2\,CaO(g)$
d. $C_5H_{12}S(l) + 9\,O_2(g) \longrightarrow 5\,CO_2(g) + SO_2(g) + 6\,H_2O(g)$

53. $Sr(s) + I_2(g) \longrightarrow SrI_2(s)$

55. $2\,Li(s) + 2\,H_2O(l) \longrightarrow 2\,Li^+(aq) + 2\,OH^-(aq) + H_2(g)$

57. $H_2(g) + Br_2(g) \longrightarrow 2\,HBr(g)$

59. 3.1 kg

61. limiting reactant: $C_7H_6O_3$, theoretical yield: 1.63 g $C_9H_8O_4$ percent yield: 74.8%

63. b

65. 0.333 g PH_3

67. 30.8 kg CO_2

69. 1.6 g C_2H_2

71. 2.8 mol A

73. 96.6 g Mn

75. d. 1.5 g K, 0.38 g O_2

77. a)

81. a. Experiments 1, 2, and 3
c. 2 A + 1 B **e.** 2 C

Chapter 5

21. **a.** 1.17 M LiCl **b.** 0.123 M $C_6H_{12}O_6$
c. 0.00453 M NaCl

23. **a.** 0.150 M NO_3^- **b.** 0.300 M NO_3^-
c. 0.450 M NO_3^-

25. **a.** 1.3 mol **b.** 1.5 mol
c. 0.211 mol

27. 37 g

29. 0.27 M

31. 6.0 L

33. 37.1 mL

35. 2.1 L

37. lead nitrate, 3.75 g, 65.3%

39. **a.** yes **b.** no **c.** yes **d.** no

41. **a.** soluble Ag^+, NO_3^-
b. soluble Pb^{2+}, $C_2H_3O_2^-$
c. soluble K^+, NO_3^-
d. soluble NH_4^+, S^{2-}

43. **a.** NO REACTION
b. NO REACTION
c. $CrBr_2(aq) + Na_2CO_3(aq) \longrightarrow CrCO_3(s) + 2\,NaBr(aq)$
d. $3\,NaOH(aq) + FeCl_3(aq) \longrightarrow Fe(OH)_3(s) + 3\,NaCl(aq)$

45. **a.** $K_2CO_3(aq) + Pb(NO_3)_2(aq) \longrightarrow PbCO_3(s) + 2\,KNO_3(aq)$
b. $Li_2SO_4(aq) + Pb(C_2H_3O_2)_2(aq) \longrightarrow PbSO_4(s) + 2\,LiC_2H_3O_2(aq)$
c. $Cu(NO_3)_2(aq) + MgS(aq) \longrightarrow CuS(s) + Mg(NO_3)_2(aq)$
d. NO REACTION

47. **a.** Complete:
$H^+(aq) + Cl^-(aq) + Li^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Li^+(aq) + Cl^-(aq)$
Net: $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$
b. Complete:
$Ca^{2+}(aq) + S^{2-}(aq) + Cu^{2+}(aq) + 2\,Cl^-(aq) \longrightarrow CuS(s) + Ca^{2+}(aq) + 2\,Cl^-(aq)$
Net: $Cu^{2+}(aq) + S^{2-}(aq) \longrightarrow CuS(s)$
c. Complete:
$Na^+(aq) + OH^-(aq) + HC_2H_3O_2(aq) \longrightarrow H_2O(l) + Na^+(aq) + C_2H_3O_2^-(aq)$
Net: $OH^-(aq) + HC_2H_3O_2(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$
d. Complete:
$6\,Na^+(aq) + 2\,PO_4^{3-}(aq) + 3\,Ni^{2+}(aq) + 6\,Cl^-(aq) \longrightarrow Ni_3(PO_4)_2(s) + 6\,Na^+(aq) + 6\,Cl^-(aq)$
Net: $3\,Ni^{2+}(aq) + 2\,PO_4^{3-}(aq) \longrightarrow Ni_3(PO_4)_2(s)$

49. Complete:
$Hg_2^{2+}(aq) + 2\,NO_3^-(aq) + 2\,Na^+(aq) + 2\,Cl^-(aq) \longrightarrow Hg_2Cl_2(s) + 2\,Na^+(aq) + 2\,NO_3^-(aq)$
Net: $Hg_2^{2+}(aq) + 2\,Cl^-(aq) \longrightarrow Hg_2Cl_2(s)$

51. Molecular:
$HBr(aq) + KOH(aq) \longrightarrow H_2O(l) + KBr(aq)$
Net ionic: $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

53. **a.** $H_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow 2\,H_2O(l) + CaSO_4(s)$
b. $HClO_4(aq) + KOH(aq) \longrightarrow H_2O(l) + KClO_4(aq)$
c. $H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow 2\,H_2O(l) + Na_2SO_4(aq)$

55. **a.** Complete ionic:
$H^+(aq) + Br^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Na^+(aq) + Br^-(aq)$
Net ionic: $H^+(aq)\ OH^-(aq) \longrightarrow H_2O(l)$
b. Complete ionic:
$HF(aq) + Na^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Na^+(aq) + F^-(aq)$
Net ionic:
$HF(aq) + OH^-(aq) \longrightarrow H_2O(l) + F^-(aq)$
c. Complete ionic:
$HC_2H_3O_2(aq) + Rb^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Rb^+(aq) + C_2H_3O_2^-(aq)$
Net ionic:
$HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$

57. 0.1810 M $HClO_4$

59. **a.** $2\,HBr(aq) + NiS(s) \longrightarrow H_2S(g) + NiBr_2(aq)$
b. $NH_4I(aq) + NaOH(aq) \longrightarrow H_2O(l) + NH_3(g) + NaI(aq)$
c. $2\,HBr(aq) + Na_2S(aq) \longrightarrow H_2S(g) + 2\,NaBr(aq)$
d. $2\,HClO_4(aq) + Li_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\,LiClO_4(aq)$

61. **a.** Ag: 0 **b.** Ag: +1
c. Ca: +2, F: −1 **d.** H: +1, S: −2
e. C: +4, O: −2 **f.** Cr: +6, O: −2

63. **a.** +2 **b.** +6 **c.** +3

65. **a.** redox reaction, oxidizing agent: O_2 reducing agent: Li
b. redox reaction, oxidizing agent: Fe^{2+} reducing agent: Mg
c. not a redox reaction **d.** not a redox reaction

67. b and c occur spontaneously in the forward direction.

69. Fe, Cr, Zn, Mn, Al, Mg, Na, Ca, K, Li

71. Mg

73. 3.32 M

75. 1.1 g

77. **a.** $2\,HCl(aq) + Hg_2(NO_3)_2(aq) \longrightarrow Hg_2Cl_2(s) + 2\,HNO_3(aq)$
b. $KHSO_3(aq) + HNO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + KNO_3(aq)$
c. $2\,NH_4Cl(aq) + Pb(NO_3)_2(aq) \longrightarrow PbCl_2(s) + 2\,NH_4NO_3(aq)$
d. $2\,NH_4Cl(aq) + Ca(OH)_2(aq) \longrightarrow 2\,NH_3(g) + 2\,H_2O(g) + CaCl_2(aq)$

79. 22 g

81. 6.9 g

83. Br is the oxidizing agent, Au is the reducing agent, 38.8 g $KAuF_4$.

85. Ca^{2+} and Cu^{2+} present in the original solution.
Net ionic for first precipitate:
$Ca^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow CaSO_4(s)$
Net ionic for second precipitate:
$Cu^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CuCO_3(s)$

87. 11.8 g AgI

89. 5.5% by mass