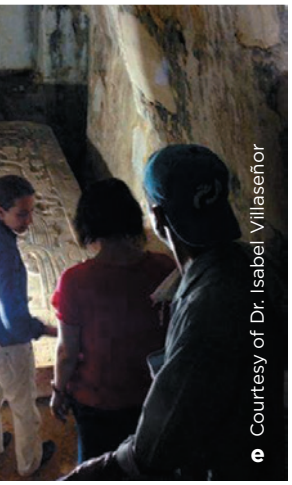
Chemical Reactions

b John Mitchell/Alamy

c De Agostini/G. Dagli Orti/Getty Images

d Courtesy of Dr. Isabel Villaseñor

Lost Cities of the Maya



For thousands of years, the Mayan civilization dominated the Yucatán region in southeast Mexico, Guatemala, and Belize. Massive cities like Tikal and Palenque glistened with temples, palaces, and stadiums. Their culture flourished, with remarkable achievements in architecture and mathematics. Mayan sculpture and writing, which used intricate figures called *hieroglyphs*, left a stunning record of their history, culture, and beliefs (**Figure 6.1**).

But something happened. Around 900 C.E., the cities of the inner Yucatán began to decline. Their populations quickly diminished. By the time the Spanish arrived in the early 1500s, most of the Maya had migrated north and east to the coastal areas. The great cities were abandoned, lost to the jungle.

What caused the decline?

For years, archaeological conservationist Isabel Villaseñor studied the ruins of Mayan cities like Palenque. Her work focused on a specific feature—the white plaster covering the great temples and palaces. In many places, ancient artisans even molded this plaster to create intricate artwork and hieroglyphs.

Over the past several decades, Dr. Villaseñor and other archaeologists have learned how the Mayans produced plaster for their buildings. The story involves chemical reactions that are still used today to produce plaster and cement.

The Maya began with limestone, a type of rock composed of calcium carbonate ($CaCO_3$), which is abundant in the Yucatán Peninsula. They crushed the limestone into small pieces and then heated it over huge fires. Above about 900 °C, calcium carbonate decomposes to form two new compounds—carbon dioxide (CO_2) and calcium oxide (CaO). The powdery calcium oxide was then mixed with water to produce a new compound, called *slaked lime*. Combined with volcanic ash, this material formed a soft putty that was easily molded. As the lime mixture slowly dried, a third chemical reaction took place, producing a hard and beautiful white surface.

But the plaster came at a price: Producing lime created huge demands on the local resources. The Maya used freshly cut trees to generate the hot, slow-burning fires needed to convert limestone to calcium oxide. Quarrying the limestone, cutting down the trees, and maintaining the fires required immense human effort—but the environmental cost may have been greater still.

Figure 6.1 (a) This image shows the palace ruins at Palenque. (b) These plaster sculptures are from the Mayan city Ek'Balam. (c) Mayan writing used hieroglyphs like these, which were often made from molded plaster. (d) Isabel Villaseñor overlooks the Temple of the Foliated Cross, Palenque. (e) Villaseñor and coworkers worked inside the crypt of the Mayan King Tikal, located inside the temple. The sarcophagus lid in front of her is carved limestone.

CHAPTER SIX

The story of Mayan plaster is a story of chemical changes. In this chapter, we will explore common chemical changes. We'll see how these changes are expressed as chemical equations and how atomic structure determines the way different elements and compounds react. We'll explore patterns of chemical change and use them to predict the outcomes of common reactions. And, at chapter's end, we will return to Palenque, to revisit the reactions that produced the great plastered surfaces, and examine clues that suggest how cities like Palenque were transformed from thriving cultural centers into abandoned ruins hidden for centuries in the Yucatán jungles.

Intended Learning Outcomes

After completing this chapter and working the practice problems, you should be able to:

6.1 Chemical Equations

- Write chemical equations to express the identity and ratio of species in a chemical change.
- Use a balanced equation to describe the ratio in which atoms or compounds react.
- Correctly balance an equation.

6.2 Classifying Reactions

- Classify synthesis, decomposition, single-displacement, and double-displacement reactions.

6.3 Reactions between Metals and Nonmetals

- Predict the products formed from the reaction of metals and nonmetals.
- Identify the species that are oxidized and reduced in a metal-nonmetal combination reaction.

6.4 Combustion Reactions

- Predict the products formed from the combustion of metals and from the combustion of hydrocarbons.

6.5 Reactions in Aqueous Solution

- Apply the solubility rules to determine whether common ionic compounds are water soluble and predict the products of precipitation reactions.
- Predict the products of acid-base neutralization reactions.
- Describe precipitation and neutralization reactions using molecular, complete ionic, and net ionic equations.

6.1 Chemical Equations

Scientists use **chemical equations** to describe chemical changes. For example, we just saw how the Mayans heated limestone (calcium carbonate) to convert it to lime (calcium oxide) and carbon dioxide gas. We could represent this change with a chemical equation:

$$CaCO_3 \rightarrow CaO + CO_2$$

In this reaction, $CaCO_3$ is the **reactant** (sometimes called the *reagent* or the *starting material*). The reactants are shown on the left side of the arrow. The **products** that form in the reaction (CaO and CO_2) are shown on the right side of the arrow.

Chemical equations also convey the *ratios* of reactants and products. For example, look carefully at the molecular representation of the reaction in **Figure 6.2**. In this reaction, two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water. Notice that this reaction follows the law of conservation of mass: Atoms are not created or destroyed in this chemical change, and the total mass of the starting materials is equal to the mass of the products (**Figure 6.3**).

We describe the reaction of hydrogen and oxygen using this chemical equation:

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

Recall from Chapter 5 that the **subscripts** (for example, H_2, O_2, H_2O) show the number of atoms in each molecule or formula unit. The numbers that precede the chemical formulas are called **coefficients**. These numbers show the ratio in

Infrequent_Flyer/Alamy

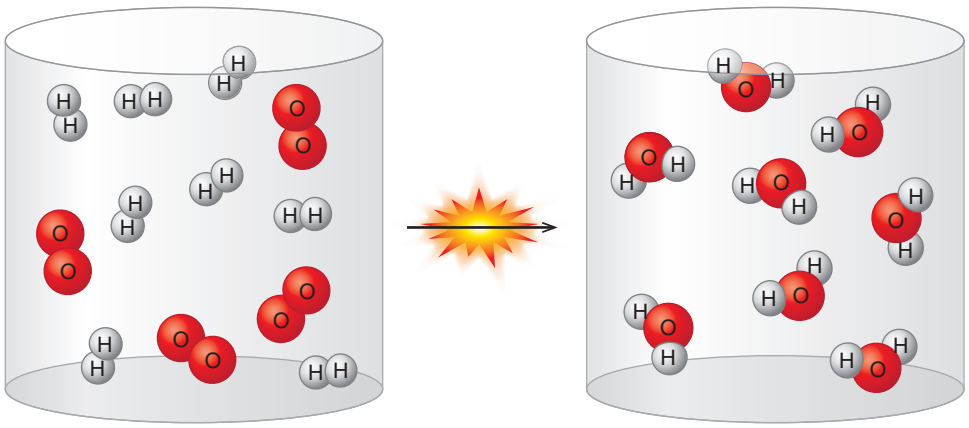


Figure 6.2 (Top) When ignited with a flame, a balloon filled with hydrogen reacts in a dramatic fireball. (Bottom) The reaction involves the combination of hydrogen with oxygen, as shown. Each oxygen molecule combines with two hydrogen molecules to produce two new water molecules.

Explore
Figure 6.2

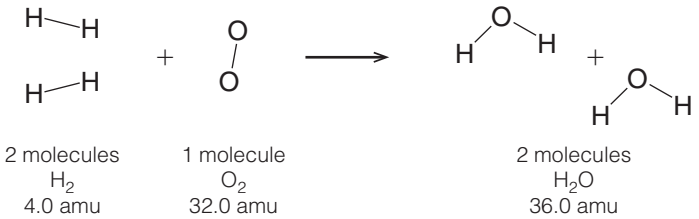


Figure 6.3 The reaction of hydrogen with oxygen to form water follows the conservation of mass. The total mass of the reactants is equal to the total mass of the products.

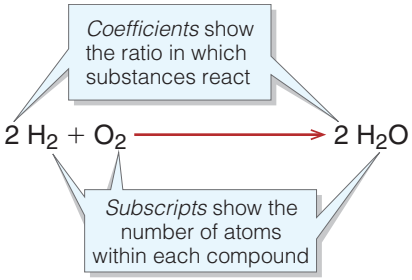


Figure 6.4 It is important to understand the difference between coefficients and subscripts.

A balanced equation contains the same number and type of atom on each side of the arrow.

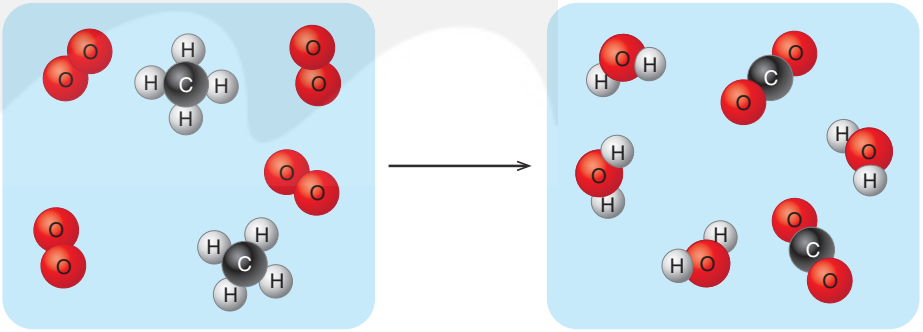
which compounds react or are formed (**Figure 6.4**). The coefficients are written to show the smallest whole-number ratio in which molecules react. If a reactant or product does not have a coefficient, we assume that coefficient to be one.

Chemical equations describe what happens at the atomic level, but they also provide a ratio for larger reactions. For example, what if 2 million molecules of hydrogen reacted with 1 million molecules of oxygen? Based on the ratio in the equation, 2 million molecules of water would form.

If we have correctly described the ratios of reactants and products in a chemical reaction, the equation is *balanced*. In a **balanced equation**, the number and type of each atom are the same in the reactants as in the products. A properly balanced equation shows the smallest whole-number ratio of reactants to products.

Example 6.1 Writing Chemical Equations

When methane gas (CH_4) burns, it reacts with molecular oxygen to form carbon dioxide and water, as represented in the image below. Write a balanced equation to describe this reaction.



In this chemical reaction, the reactants are CH_4 and O_2. The products for this reaction are H_2O and CO_2. Notice that in this picture, the number of carbon, hydrogen, and oxygen atoms is the same in both the starting materials and the products. Two CH_4 molecules react with four O_2 molecules to produce two CO_2 molecules and four H_2O molecules. We therefore write this as

$$2\ CH_4 + 4\ O_2 \rightarrow 2\ CO_2 + 4\ H_2O$$

While this is a correct representation, it is not the smallest whole-number ratio. After dividing each coefficient by two, we get the properly balanced equation. ●

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

Example 6.2 Predicting Ratios from Balanced Chemical Equations

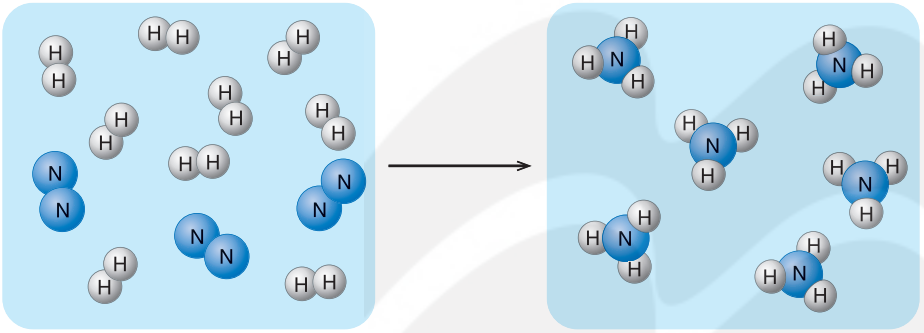
When heated, CH_2N_2 undergoes an explosive chemical reaction to produce two new compounds, N_2 and C_2H_4, as shown in the balanced equation below. If 20 molecules of CH_2N_2 react in this way, how many molecules of C_2H_4 form? How many molecules of N_2 and C_2H_4 form if 2 million molecules of C_2H_4 react?

$$2\ CH_2N_2 \rightarrow 2\ N_2 + C_2H_4$$

The balanced equation tells us that 2 molecules of CH_2N_2 produce 2 molecules of N_2 and 1 molecule of C_2H_4. Following this ratio, 20 molecules of CH_2N_2 react to produce 20 molecules of N_2 and 10 molecules of C_2H_4. Similarly, 2 million molecules of CH_3N_2 react to produce 2 million molecules of N_2 and 1 million molecules of C_2H_4. ●

TRY IT

1. Write a balanced equation to describe the reaction of nitrogen and hydrogen shown here:



Check it Watch explanation

2. Based on the balanced equation in Question 1, how many molecules of NH_3 could be made from the reaction of 3,000 molecules of hydrogen gas with 1,000 molecules of nitrogen gas?

Check it Watch explanation

Balancing Equations

Sometimes, when describing a reaction, we begin with the identities of the starting materials and products without knowing the ratio in which the reaction takes place. For example, iron can react with oxygen to form iron(III) oxide—the reddish-brown compound we know as rust. To describe this process, we could write the equation as

$$Fe + O_2 \rightarrow Fe_2O_3$$

However, notice that something is funny with this equation: We start with two oxygen atoms on the left-hand side, but we have three oxygen atoms on the right-hand side. Did we conjure up another oxygen atom? Of course not! Although we have identified the compounds correctly, we have not accounted for the *ratio* in which the atoms react.

To more accurately describe this reaction, we must balance the equation to account for each atom in the starting materials and products. To do this, we add coefficients to the elements and compounds in the equation until the number and type of atom on each side are the same. This process is shown in **Figure 6.5**.

KENNETH GARRETT/National Geographic Creative

The Maya used iron oxide to produce the red-brown colors in their artwork.

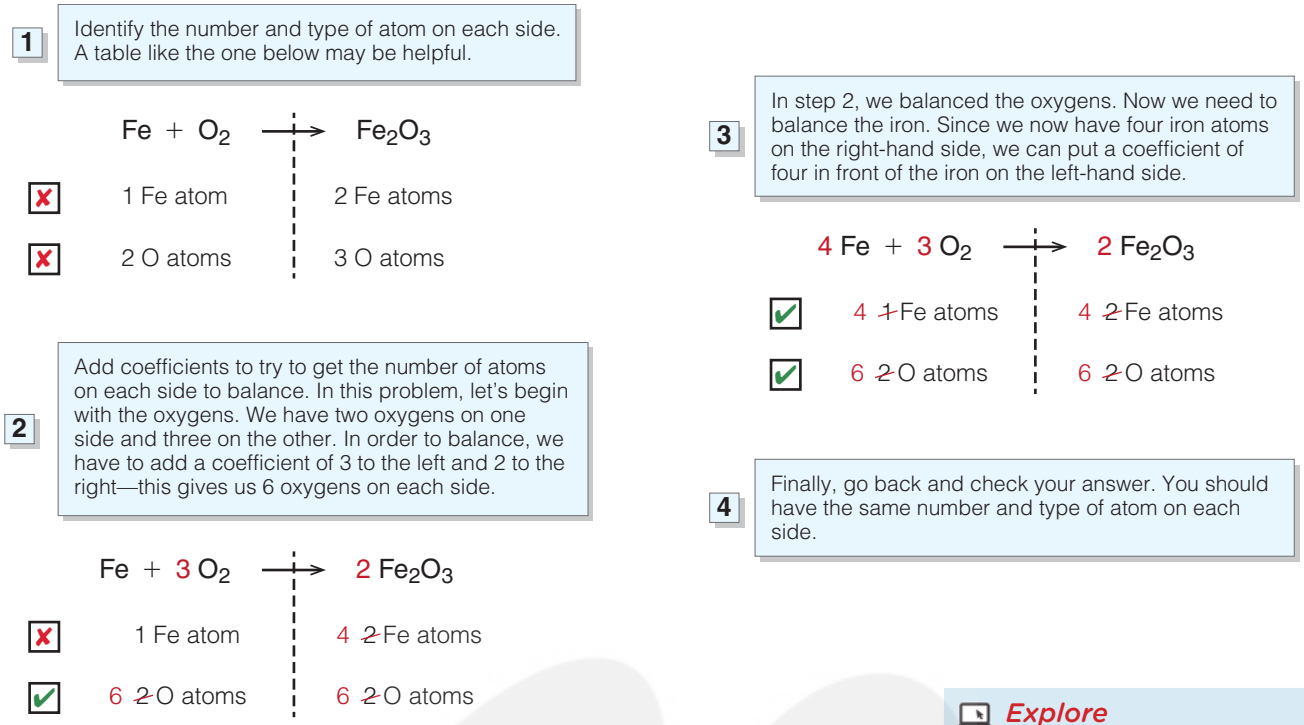


Figure 6.5 Balancing equations is a step-by-step process.

Explore Figure 6.5

Balance equations by adjusting the coefficients, not the subscripts.

Notice that when balancing this equation, *we never touched the subscripts*. The subscripts show the chemical identity of each molecule or formula unit present. When we balance equations, we are simply adjusting the ratio of starting materials and products—not altering the identity of the substances. Examples 6.3 and 6.4 describe how to balance two more equations.

Example 6.3 Balancing an Equation

Balance the equation below to show the proper ratio for the reaction of potassium metal with copper(II) chloride to produce potassium chloride and copper metal:

$$K + CuCl_2 \rightarrow Cu + KCl \qquad \text{not balanced}$$

Let's begin by balancing the chlorine atoms. To do this, we put the coefficient 2 in front of KCl. This balances the chlorine atoms, but not the potassium atoms. However, we can balance the potassium atoms by putting a 2 in front of the K. •

$$2\ K + CuCl_2 \rightarrow Cu + 2\ KCl \qquad \text{balanced}$$

Example 6.4 Balancing an Equation

Balance the following equation:

$$Al_2O_3 + C + Cl_2 \rightarrow AlCl_3 + CO \qquad \text{not balanced}$$

Notice that in this equation, aluminum and oxygen always appear as part of a compound while carbon and chlorine appear as elemental forms. This means that we can add coefficients to the carbon or chlorine without disrupting the balance of any

of the other compounds. *In general, it is easier to balance elemental forms last.* So in this example, we would balance aluminum and oxygen first because they appear only in compounds:

$$Al_2O_3 + C + Cl_2 \rightarrow 2\ AlCl_3 + 3\ CO \qquad \text{C, Cl not balanced}$$

Now that aluminum and oxygen are balanced, we can balance the elemental forms. Adding coefficients to the carbon and chlorine gives us the balanced equation:

$$Al_2O_3 + 3\ C + 3\ Cl_2 \rightarrow 2\ AlCl_3 + 3\ CO \qquad \text{balanced}$$

The stepwise strategy for solving this equation is shown in **Figure 6.6.**

It is helpful to balance elemental forms last.

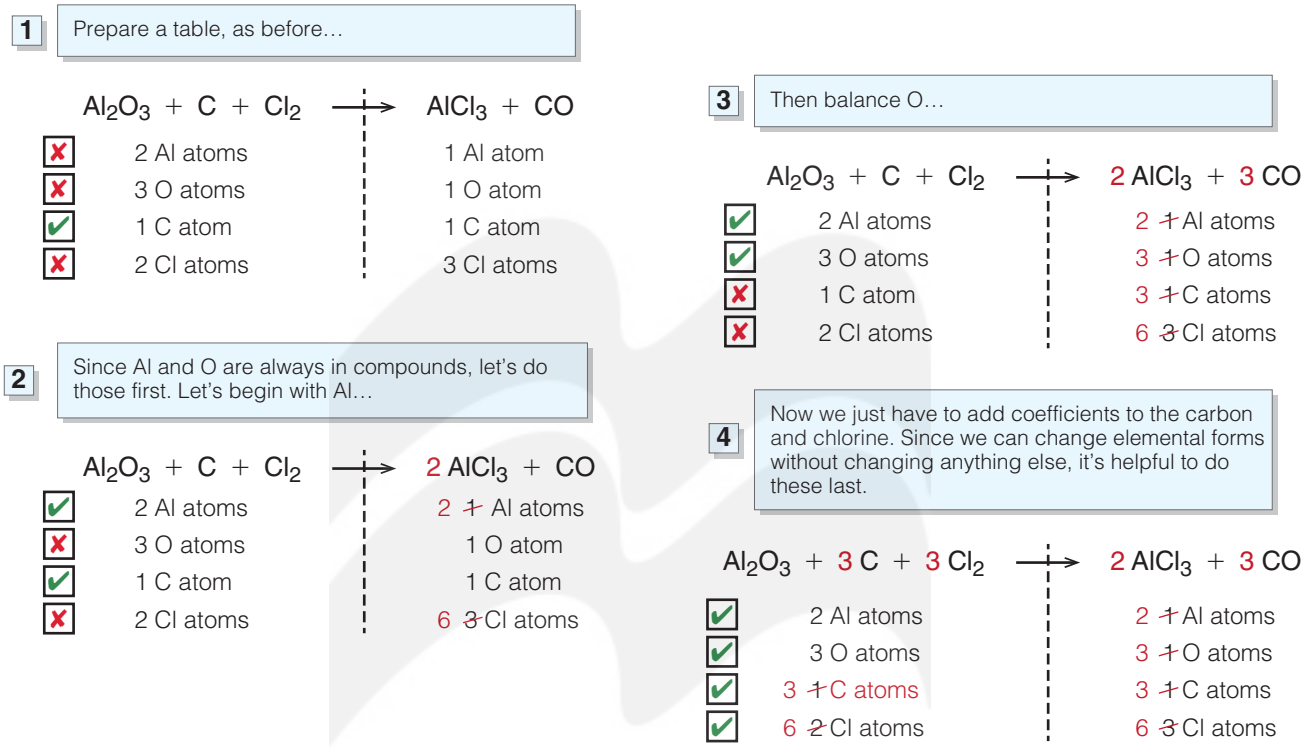


Figure 6.6 When balancing an equation, it is helpful to balance the elemental forms last.

Explore Figure 6.6

TRY **IT**

3. Balance each of these equations:

$$Mg + O_2 \rightarrow MgO \qquad HCl + Co \rightarrow CoCl_2 + H_2$$

$$Br_2 + Al \rightarrow AlBr_3 \qquad C_4H_8 + O_2 \rightarrow CO_2 + H_2O$$

Check it *Watch explanation*

Strategies for Balancing Equations

We can use a few other techniques to balance complicated equations more easily. In some equations, it is possible to balance polyatomic ions rather than atoms. In other equations, using a fractional coefficient may be helpful as an intermediate step. Examples 6.5 and 6.6 illustrate these techniques.

Example 6.5 Balancing Equations with Polyatomic Ions

Balance the following equation:

$$Ni(NO_3)_2 + NaOH \rightarrow Ni(OH)_2 + NaNO_3 \qquad \text{not balanced}$$

We could try to balance this equation by making a table of all the elements present, as we did before. However, notice that this equation contains two common polyatomic ions—nitrate (NO_3^-) and hydroxide (OH^-). Although it is possible to balance the atoms individually, it is much easier to simply balance the nitrate and hydroxide ions. We have two hydroxide ions on the right, so we add the coefficient 2 in front of NaOH to balance them on the left. Similarly, we place the coefficient 2 in front of $NaNO_3$, to balance the nitrate ions on the left and right:

$$Ni(NO_3)_2 + 2\ NaOH \rightarrow Ni(OH)_2 + 2\ NaNO_3 \qquad \text{balanced}$$

This approach works as long as the ions in the reaction do not change. We will see several examples of this type of reaction in the sections that follow. ●

Example 6.6 Balancing with Fractional Coefficients

Balance the following equation:

$$C_2H_6 + O_2 \rightarrow CO_2 + H_2O \qquad \text{not balanced}$$

Because oxygen appears in elemental form in this equation, we should balance it last. We therefore begin by adding coefficients to the CO_2 and the H_2O to balance the carbon and hydrogen:

$$C_2H_6 + O_2 \rightarrow 2\ CO_2 + 3\ H_2O \qquad \text{oxygen not balanced}$$

Now we need to balance the oxygen atoms. Notice that on the right-hand side of the equation, seven oxygen atoms are present. For the equation to balance, we need seven oxygen atoms on the left-hand side. But oxygen atoms come in groups of two. So how do we get seven?

The easiest way to solve this problem is to recognize that we need three and one-half oxygen molecules to give us seven oxygen atoms. We write 3½ as the improper fraction 7/2, as follows:

$$C_2H_6 + \frac{7}{2}\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O \qquad \text{balanced, but not whole numbers}$$

Now the equation is balanced, but it is not properly written as the lowest whole-number ratio. To correct this, we multiply each coefficient by two, resulting in the lowest whole-number balanced equation:

$$2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O \qquad \text{balanced}$$

Finally, the key to being able to balance equations consistently is *practice*. Be sure to try each of the problems below, and do the additional problems at the end of the chapter. ●

Photo Researchers, Inc.

Here's a mnemonic to keep in mind when writing and balancing equations: Oxygen atoms are like peanut butter cups, they come in packs of two.

TRY IT

4. Balance each of these equations, using the techniques described in this section:

$$Pb(ClO_3)_2 + NaI \rightarrow NaClO_3 + PbI_2$$

$$Al(NO_2)_3 + Ca \rightarrow Al + Ca(NO_2)_2$$

$$C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$$

Check it *Watch explanation*

Equations with Phase Notations

For many chemical processes, it is important to know not only the identity of the components but also the *phase* or *state* of the components. Because of this, *phase notations* are sometimes used in chemical equations (**Table 6.1**). If we wish to indicate that a compound is a solid, we write (*s*) after the chemical formula. Similarly, we represent liquids by the symbol (*l*) and gases by the symbol (*g*).

Many reactions involve substances that are dissolved in water. A substance that is dissolved in water is said to be in an *aqueous solution*. We represent this by using the symbol (*aq*) in the balanced equation. For example, consider the change that occurs when you pour a soft drink (**Figure 6.7**): The dissolved carbonic acid (H_2CO_3) reacts to produce two new compounds, water and carbon dioxide gas. We can write this equation using phase notation as follows:

$$H_2CO_3\ (aq) \rightarrow H_2O\ (l) + CO_2\ (g)$$

We sometimes use equations with phase symbols to represent physical changes, such as changes of state. For example, we could use the following equation to show water freezing:

$$H_2O\ (l) \rightarrow H_2O\ (s)$$

In the sections that follow, we will use the phase symbols extensively to describe reactions.

TABLE 6.1 Phase Symbols

Symbol	Meaning
(*s*)	Solid
(*l*)	Liquid
(*g*)	Gas
(*aq*)	Aqueous solution (dissolved in water)


studiovin/Shutterstock

Figure 6.7 The bubbles that form when a soft drink is poured arise from the reaction of carbonic acid to produce carbon dioxide and water.

6.2 Classifying Reactions

As you begin to study chemical reactions, you'll encounter a large number of examples. You may find it overwhelming—I certainly did when I started learning chemistry, and sometimes I still do. So how can you begin to make sense of it all? One of the keys is to organize reactions into similar types. This makes it easier to recognize patterns in chemical behavior and to predict similar reactions that may take place.

Let's illustrate this idea with an example from the grocery. To make shopping easier, the store organizes food by groups: meats, fresh vegetables, breads. It also arranges food by origin: Italian, Mexican, Thai. These simple organizational systems help us remember, understand, and communicate with one another.

In this section and those that follow, we'll explore broad patterns of chemical behavior. We'll classify reactions by the types of products, or by the way the reaction takes place. No single classification system is universal, but they all serve one purpose—to organize and simplify the world around us. As we begin to explore chemical reactions, let's consider four broad types of reactions: decomposition, synthesis, single displacement, and double displacement (**Figure 6.8**).

Decomposition:
One forms two or more

$$2H_2O \longrightarrow 2H_2 + O_2$$
$$CaCO_3 \longrightarrow CaO + CO_2$$

Single Displacement:
One element replaces another

$$Zn + CuCl_2 \longrightarrow ZnCl_2 + Cu$$
$$Ca + 2HBr \longrightarrow CaBr_2 + H_2$$

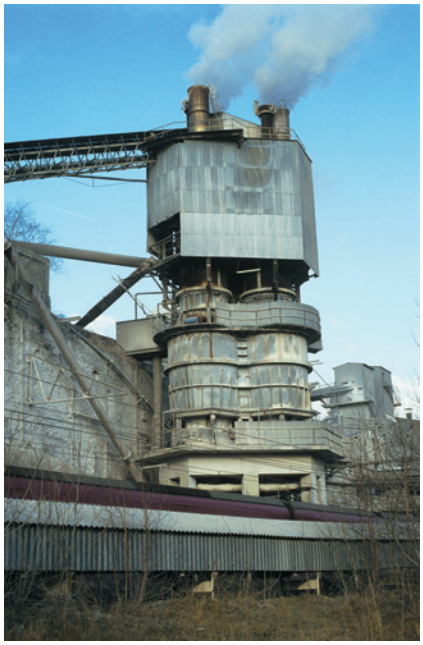
Synthesis (Combination):
Two forms one

$$H_2 + Cl_2 \longrightarrow 2HCl$$
$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Double Displacement:
Two ions replace each other

$$NaI + AgNO_3 \longrightarrow AgI + NaNO_3$$
$$MgBr_2 + Pb(ClO_4)_2 \longrightarrow PbBr_2 + Mg(ClO_4)_2$$

Figure 6.8 Many reactions fall under the four categories shown here.



Dirk Wiersma/Science Source

Figure 6.9 Lime kilns like this one convert limestone into calcium oxide.

In **decomposition** reactions, a single reactant forms two or more products. As an example of this type of reaction, let's return to the Mayan production of lime plaster. The first step in this process was the decomposition of calcium carbonate into two simpler compounds, calcium oxide and carbon dioxide:

$$CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$$

Although it is an ancient technology, this decomposition reaction is still vitally important today. Calcium oxide is the key ingredient in cement and mortar, and manufacturers use modern kilns to convert limestone (calcium carbonate) into calcium oxide (**Figure 6.9**).

Synthesis (or *combination*) reactions occur when two reactants join together to form a single product. The second step in the production of lime plaster is an example of a synthesis reaction: In this reaction, calcium oxide combines with water to produce a single new compound, calcium hydroxide:

$$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (s)$$

Another example of a synthesis reaction is the combination of hydrogen gas and bromine gas to form hydrogen bromide:

$$H_2\ (g) + Br_2\ (g) \rightarrow 2\ HBr\ (g)$$

We will look extensively at synthesis reactions in Section 6.3.

In a **single-displacement** reaction, one element replaces another element in a compound. For example, when zinc is combined with aqueous copper sulfate, elemental zinc displaces copper to form a new compound. This process converts copper into its elemental form:

$$Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq)\ Cu\ (s)$$

Similarly, the compound HCl reacts with elemental tin to produce a new compound, tin(II) chloride, and elemental hydrogen:

$$Sn\ (s) + 2\ HCl\ (aq) \rightarrow SnCl_2\ (aq) + H_2\ (g)$$

In a **double-displacement** reaction, two compounds rearrange to form two new compounds. These reactions involve a "swap" of cation-anion pairs. For example, the reaction of KCl with $AgNO_3$ in aqueous solution produces two new compounds, KNO_3 and AgCl. This reaction can be thought of as simply a swap of anions:

$$KCl + AgNO_3 \rightarrow KNO_3 + AgCl$$

The anions "swap" positions.

A simple way to think about single- and double-displacement reactions is to consider the cations and anions as dance partners (**Figure 6.10**). In a single displacement, an element "cuts in" on the dance—producing a new element and a new compound (new dance partners). In a double displacement, the two dancing couples switch partners, producing two new compounds.

Example 6.7 Classifying a Chemical Reaction

An aqueous solution of hydrochloric acid reacts with solid zinc metal to produce hydrogen gas and a new compound, zinc chloride, that dissolves in water. Write a balanced equation to show this reaction, then classify it as a synthesis, decomposition, single-displacement, or double-displacement reaction. Finally, add phase symbols to describe the phase of each reactant and starting material.

Our two reactants are hydrochloric acid and elemental zinc. The two products are zinc chloride and hydrogen gas. From Chapter 5, we know that the formula for hydrochloric

acid is HCl. The formula for zinc chloride is $ZnCl_2$, and hydrogen gas is H_2. We begin with this unbalanced equation:

$$HCl + Zn \rightarrow ZnCl_2 + H_2 \qquad \text{not balanced}$$

Next, we balance this equation. In this case, addition of the coefficient 2 in front of HCl is all that is needed:

$$2\ HCl + Zn \rightarrow ZnCl_2 + H_2 \qquad \text{balanced}$$

Notice that this reaction contains one element and one compound in the reactants, and a different element and compound in the products. This is an example of a single-displacement reaction.

From the question, we know that HCl and $ZnCl_2$ are both dissolved in water, zinc is a solid, and H_2 is a gas. In the final step, we represent this using phase symbols. •

$$2\ HCl\ (aq) + Zn\ (s) \rightarrow ZnCl_2\ (aq) + H_2\ (s) \qquad \text{balanced, with phase symbols}$$

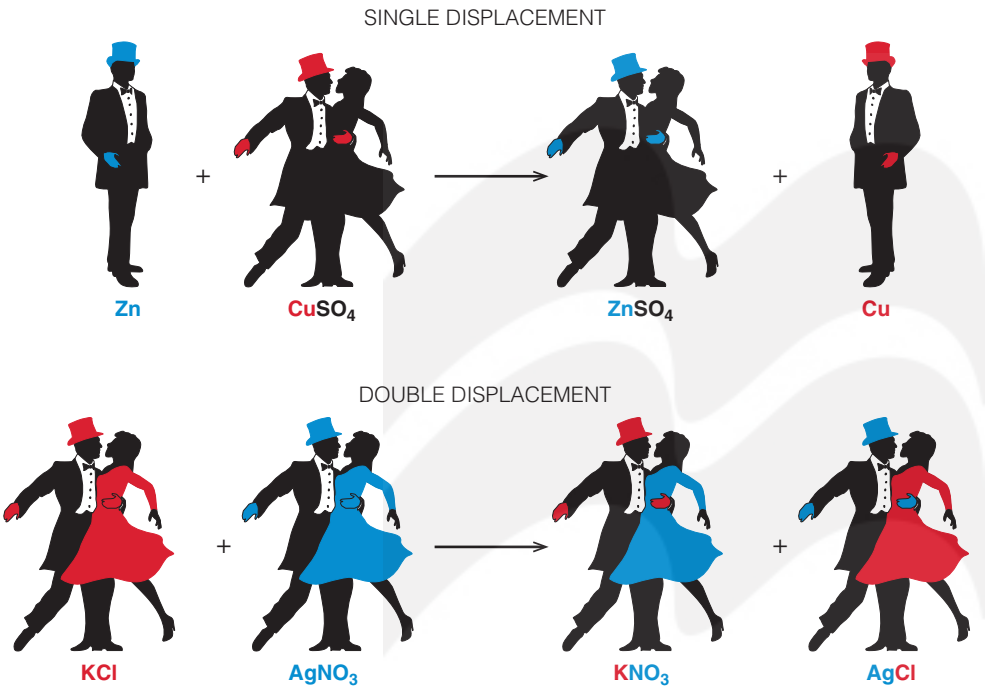


Figure 6.10 Single- and double-displacement reactions are similar to changes in dance partners.

TRY IT

5. Classify the following changes as decomposition, synthesis, single-displacement, or double-displacement reactions:

$$Pb(NO_3)_2 + 2\ KI \rightarrow 2\ KNO_3 + PbI_2$$

$$4\ Al + 3\ O_2 \rightarrow 2\ Al_2O_3$$

$$2\ H_2O_2 \rightarrow 2\ H_2O + O_2$$

$$Fe + SnCl_2 \rightarrow Sn + FeCl_2$$

Check it *Watch explanation*

6. When solid lead(II) sulfite is heated, it forms two new compounds, solid lead(II) oxide and sulfur dioxide gas. What type of reaction is this? Write a balanced equation for this reaction, including phase symbols.

Check it *Watch explanation*

6.3 Reactions between Metals and Nonmetals

Now that we have a method for classifying reactions, let's look at some ways that specific substances react. We'll begin with one of the most common types of synthesis reactions—the combination of a metal and a nonmetal.

Metals and nonmetals react to form ionic compounds. Many of these reactions are intense, producing bright flames and tremendous heat (**Figure 6.11**).

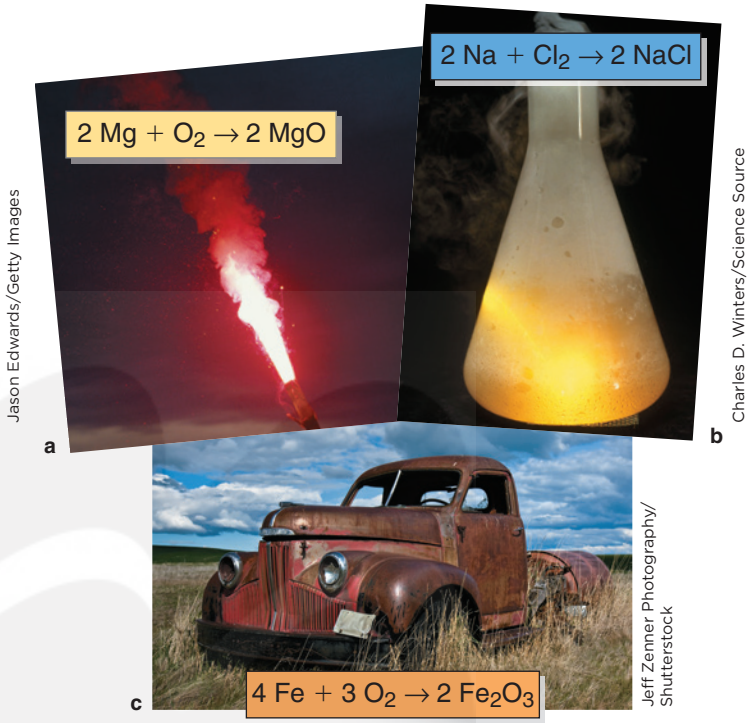


Figure 6.11 Metals and nonmetals react to form ionic compounds. (a) The bright light of a signal flare results from the reaction of metallic magnesium with oxygen gas. (b) Sodium and chlorine gas react violently to produce sodium chloride. (c) Rust forms by the slow reaction of iron with oxygen gas.

Explore Figure 6.11

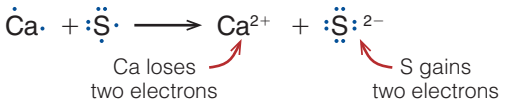
Reacting with O_2
Most metals react with oxygen to form new compounds. In these reactions, the metal loses electrons. The term *oxidized* arises from these reactions, but its meaning is broader. In any reaction where an atom loses electrons, we say the atom is oxidized.

In these reactions, the metal loses electrons to form a positively charged ion (cation) while the nonmetal gains electrons to form a negatively charged ion (anion). The result is an ionic compound:

$$\text{Metal} + \text{Nonmetal} \rightarrow \text{Ionic compound}$$

The reactions of metals and nonmetals are an example of an **oxidation-reduction reaction**. When an atom is **oxidized**, it loses electrons. When an atom is **reduced**, it gains electrons. We will explore oxidation-reduction reactions in more detail in Chapter 14.

Let's look at some examples of this type of reaction. We'll begin with the reaction of elemental calcium with sulfur to form calcium sulfide. The calcium loses two electrons to form Ca^{2+} while the sulfur gains two electrons to form the sulfide ion, S^{2-}. We can illustrate this reaction using Lewis symbols, or we can represent it using a chemical equation, as shown here:



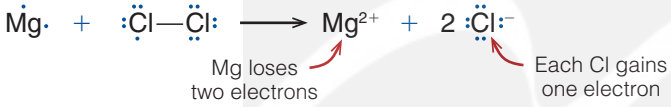
$$Ca + S \rightarrow CaS$$

Charles D. Winters/Science Source

Figure 6.12 Magnesium reacts violently with chlorine gas.

Consider another example: Magnesium metal reacts with chlorine gas to produce magnesium chloride (**Figure 6.12**).

The metal atom (magnesium) loses two electrons while each chlorine atom gains one electron. In this reaction, a covalent bond breaks as the ions form:



$$Mg + Cl_2 \rightarrow MgCl_2$$

So, how can we predict the compounds that will form in metal-nonmetal reactions? Recall from Chapter 5 that metals and nonmetals form specific, stable ions. By knowing the charges of the common ions, we can often predict the ionic compounds that result from these reactions. This process is illustrated in the following two examples.

Many metal-nonmetal reactions produce common ions.

Example 6.8 Predicting the Products when Metals and Nonmetals React

What compound is formed when aluminum metal reacts with chlorine gas? Write a balanced equation for this reaction.

We know that aluminum forms a +3 ion while a chloride ion has a charge of −1. Therefore, the compound formed must have the formula $AlCl_3$. We can write an unbalanced equation as shown:

$$Al + Cl_2 \rightarrow AlCl_3 \qquad \text{not balanced}$$

To balance the equation, we add coefficients to make sure the number and type of atom on each side of the equation are the same, as shown. ●

$$2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3 \qquad \text{balanced}$$

Example 6.9 Predicting the Products when Metals and Nonmetals React

When tin metal reacts with bromine, it is oxidized to the tin(IV) ion while bromine is reduced to form bromide ions. Write a balanced equation for this reaction.

The question indicates that the tin(IV) ion is formed. Therefore, the neutral compound produced will have the formula $SnBr_4$. We can write an unbalanced equation as shown:

$$Sn + Br_2 \rightarrow SnBr_4 \qquad \text{not balanced}$$

To balance the equation, we only need to add a coefficient in front of the bromine. This gives us the final balanced equation. ●

$$Sn + 2\ Br_2 \rightarrow SnBr_4 \qquad \text{balanced}$$

TRY IT

7. If zinc metal reacts with bromine gas, which species is oxidized? Which species is reduced?

Check it *Watch explanation*

8. Predict the products from these reactions:

a. the reaction of potassium and sulfur

b. the reaction of silver and oxygen

Check it *Watch explanation*

6.4 Combustion Reactions

Many common reactions involve elemental oxygen. Oxygen is a major component of Earth's atmosphere, and it reacts with nearly every other element and with many compounds. Both metals and nonmetals react with oxygen. For example, consider the way that oxygen reacts with tin (a metal) and also with carbon and with sulfur (both nonmetals):

$$Sn + O_2 \rightarrow SnO_2$$

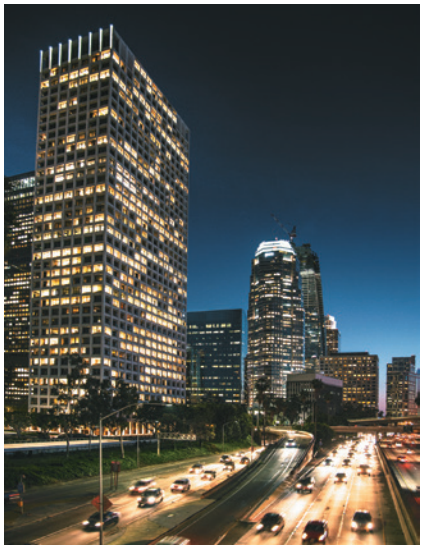
$$C + O_2 \rightarrow CO_2$$

$$S + O_2 \rightarrow SO_2$$

The first product, tin(IV) oxide, is an ionic compound. The other two products, carbon dioxide and sulfur dioxide, are molecular compounds with covalent bonds. And yet all three elements react with oxygen in similar ways. Each of these reactions releases a large amount of heat and forms an oxide compound. These rapid and exothermic reactions with oxygen are called **combustion** reactions.

Some of the most important combustion reactions involve compounds composed of hydrogen and carbon, called *hydrocarbons*. These reactions are important because fossil fuels, such as coal, oil, and natural gas, are composed primarily of hydrocarbons (**Table 6.2**). The combustion reactions of fossil fuels produce most of the energy used worldwide (**Figure 6.13**).

When hydrocarbons react with oxygen, they form two main products, carbon dioxide and water. For example, methane (CH_4) is the main component of natural gas (**Figure 6.14**).



franckreporter/Getty Images

Figure 6.13 The combustion of coal and natural gas produces most of our electricity, and the combustion of gasoline powers most of our transportation.

TABLE 6.2 Common Hydrocarbons

Formula	Name	Use
CH_4	Methane	Natural gas
C_2H_2	Acetylene	Torches for cutting and welding
C_2H_4	Ethylene	Manufacture of plastic
C_3H_8	Propane	Natural gas component; used for heating and power
C_4H_{10}	Butane	Lighter fluid
C_6H_6	Benzene	Solvent; precursor for many pharmaceutical compounds
C_8H_{18}	Octane	Component of gasoline

Chepko Danil Vitalevich/ Shutterstock

Figure 6.14 Gas stoves use the combustion of methane gas to produce heat.

The combustion reaction of methane follows this balanced equation:

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

Similarly, octane, C_8H_{18}, is a key component of gasoline. This compound also reacts with oxygen to produce carbon dioxide and water:

$$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$$

Nonmetal elements often react with oxygen to form many different compounds. For example, the combustion of elemental sulfur produces several different oxides, including sulfur dioxide (SO_2) and sulfur trioxide (SO_3). These compounds, commonly referred to as *sulfur oxides* (SO_x), are toxic gases that contribute to environmental issues such as acid rain. Because coal contains a small amount of sulfur (in addition to hydrocarbon compounds), the combustion of coal leads to the release of sulfur oxides into the atmosphere (**Figure 6.15**).

The combustion of hydrocarbons produces carbon dioxide and water.

Example 6.10 Predicting the Products of Combustion Reactions

Write a balanced equation for the combustion of propane gas, a common fuel used for home heating, cooking, and so on. The formula for propane is C_3H_8.

Hydrocarbons react with oxygen to produce carbon dioxide and water. We can write an unbalanced equation for this reaction first:

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

We then balance the equation to give us the final answer:

$$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$

For another example of balancing a combustion reaction, see Example 6.6 earlier in this chapter. ●

TRY IT

9. Write a balanced equation that describes the combustion of magnesium.

Check it *Watch explanation*

10. Write a balanced equation that describes the combustion of C_5H_{12}.

Check it *Watch explanation*

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Figure 6.15 When oxygen reacts with sulfur, it produces a set of compounds known as sulfur oxides. Because coal contains small amounts of sulfur, sulfur oxides are produced when coal is burned for fuel.

6.5 Reactions in Aqueous Solution

Representing Dissociation: Molecular and Ionic Equations

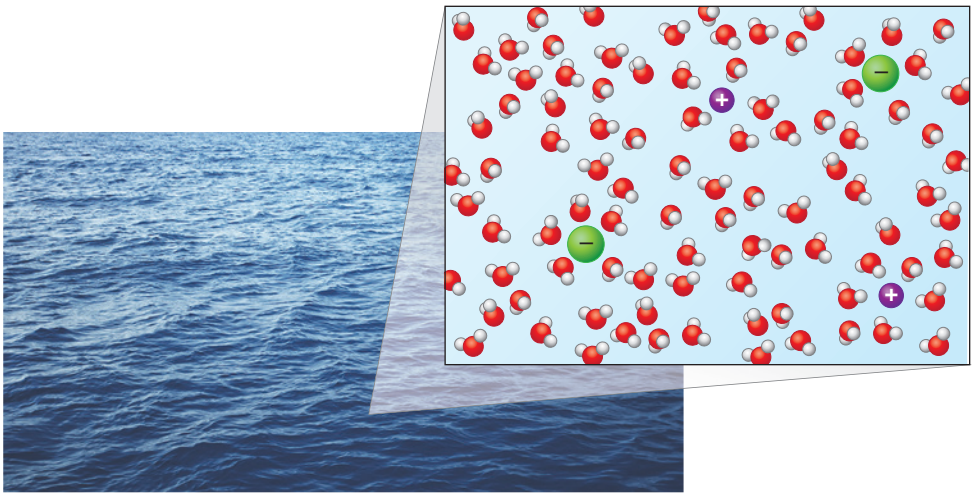
Ionic compounds dissociate when they dissolve in water.

In Chapter 5 we described the behavior of different types of compounds when dissolved in water (Section 5.6). Recall that when an ionic compound dissolves in water, it dissociates into cations and anions. That is, the ions that make up the compound are pulled apart and surrounded by water molecules (**Figure 6.16**). This behavior enables different ions to react in different ways.

When describing the chemical processes that take place in solution, we use two different equation styles. In a **molecular equation**, we write ions together as compounds. In an **ionic equation**, we show dissociated ions as separate species. For example, potassium bromide dissociates in water. We can write this as a molecular equation (shown in blue) or as an ionic equation (shown in red):

Molecular equation: $KBr\ (s) \rightarrow KBr\ (aq)$

Ionic equation: $KBr\ (s) \rightarrow K^+\ (aq) + Br^-\ (aq)$



Vitaly Korovin/Shutterstock

Figure 6.16 When an ionic compound dissolves in water, the cations and anions dissociate.

Similarly, we could show the dissociation of magnesium nitrate in solution. In this example, the formula unit $Mg(NO_3)_2$ dissociates to form one magnesium ion and two nitrate ions:

Molecular equation: $Mg(NO_3)_2\ (s) \rightarrow Mg(NO_3)_2\ (aq)$

Ionic equation: $MgNO_3\ (s) \rightarrow Mg^{2+}\ (aq) + 2\ NO_3^-\ (aq)$

Notice that the ionic equation shows only the dissociated (aqueous) ions separately; we show solid ionic compounds as a complete unit.

Example 6.11 Writing an Ionic Equation

Write an ionic equation showing how solid potassium sulfate dissociates when dissolved in water. Include phase symbols.

Potassium sulfate has the formula K_2SO_4. When dissolved in water, this compound dissociates into the component potassium and sulfate ions. We write this as follows:

$$K_2SO_4\ (s) \rightarrow 2\ K^+\ (aq) + SO_4^{2-}\ (aq)$$

In the reactant, we show the two potassium ions with a subscript because they are part of a compound. However, when the compound dissociates, the potassium ions also separate. Because the two aqueous ions are no longer together, they are shown with a coefficient in the products. ●

TRY IT

11. Write ionic equations to show the following two changes:

a. Solid ammonium bicarbonate dissolves in water.

b. Solid aluminum nitrate dissolves in water.

Check it *Watch explanation*

Solubility Rules and Precipitation Reactions

Not all ionic solids dissolve in water. Many ionic compounds are insoluble. What determines whether a compound is water soluble?

To answer this question, let's think about the energy changes that occur when an ionic solid dissociates in water (**Figure 6.17**). In an ionic lattice, ions are stabilized by the oppositely charged ions that surround them. When they dissociate, the ions are stabilized by water molecules. Which situation is more energetically favorable? The answer to this question determines whether a compound will dissolve.

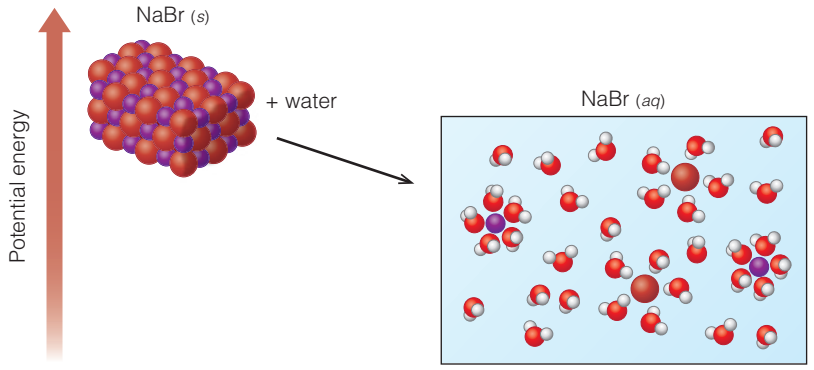


Figure 6.17 Sodium bromide dissolves in water because the energy gain from the interaction between the ions and water is greater than the energy gain between the ions in the lattice.

Practice
Ion charges

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Do you remember the names, symbols, and charges for the common ions covered in Chapter 5? If not, you should go back and relearn the list. You will use them extensively in the sections ahead. Try the Ion Charges Quick Recall to test your knowledge!

Ions with a charge of +1 or −1 are usually soluble.

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Have you ever thought about leaving one job to take another? There are many factors to consider. How does the pay compare? Are the hours better? Is the work more enjoyable or rewarding? In the same way, solubility is determined by competing factors. Which is more stable—the ionic lattice, or the dissociated ions surrounded by water molecules? The answer to this question varies for different compounds.

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Magnesium hydroxide is a common treatment for acid indigestion. Because magnesium hydroxide is insoluble in water, the mixture has a white, milky appearance, leading to the name *milk of magnesia*.

The solubility of ionic compounds depends on many factors, including the charge of the ions, the size of the ions, and the way they pack together. The more strongly the ions bind to each other, the less likely they are to dissolve in water. For example, the greater the charge on an ion, the more tightly it sticks to oppositely charged ions. As a result, compounds composed of ions with a charge of two or three tend to be insoluble in water. Iron(III) oxide (Fe_2O_3), lead(II) sulfide (PbS), and barium carbonate ($BaCO_3$) are all insoluble in water.

The way ions fit together can be hard to predict, and sometimes ionic compounds are surprisingly stable. For example, ionic compounds containing the halogen ions (chloride, bromide, and iodide) are nearly always water soluble—unless they are bonded to silver (Ag^+) or lead(II) (Pb^{2+}).

The solubilities of common ionic compounds are summarized in **Figure 6.18** and in **Table 6.3**. Note some of the key trends:

1. Compounds containing alkali metals (Li^+, Na^+, K^+), ammonium (NH_4^+), or the large oxyanions nitrate (NO_3^-), chlorate (ClO_3^-), perchlorate (ClO_4^-), or acetate ($C_2H_3O_2^-$), are always soluble.
2. Compounds containing chloride (Cl^-), bromide (Br^-), and iodide (I^-) are soluble, except when they are bonded to silver (Ag^+) or lead(II) (Pb^{2+}).
3. Compounds containing sulfate (SO_4^{2-}) are usually soluble, except when they are bonded to Ba^{2+}, Ca^{2+}, Pb^{2+}, or Ag^+.
4. Most other ionic compounds are insoluble.

These guidelines can help to determine whether compounds are water soluble. However, many compounds are slightly soluble—meaning that a small number of ions will dissolve, but most of the ionic solid will not. In some applications, you may need to know the solubilities more precisely than is given here. Solubility also depends on temperature: Many ionic compounds that are insoluble at room temperature become quite soluble at higher temperatures.

Example 6.12 Predicting the Solubility of an Ionic Compound

Using the information in Figure 6.18 and Table 6.3, determine whether these compounds are soluble or insoluble in water:

a. Na_3PO_4 *b.* $AlCl_3$ *c.* $CaCO_3$

Let's begin with sodium phosphate, Na_3PO_4. Notice that Na^+ is on the list of ions that are always soluble. Even though phosphate is not typically a soluble ion, this compound is soluble because it contains sodium.

The situation with aluminum chloride is similar: Chlorides are nearly always soluble, and so $AlCl_3$ is water soluble.

Finally, let's look at calcium carbonate, $CaCO_3$. Neither of these ions is listed as soluble or as usually soluble. This compound is insoluble. •

TRY IT

12. Use Table 6.3 to determine if these compounds are soluble or insoluble in water:

a. $NaC_2H_3O_2$ b. $FeCl_3$ c. $CaSO_4$ d. $AlPO_4$

Check it *Watch explanation*

13. Which of these compounds is more likely to be water soluble: cesium fluoride (CsF) or barium chromate ($BaCrO_4$)? Why?

Check it *Watch explanation*

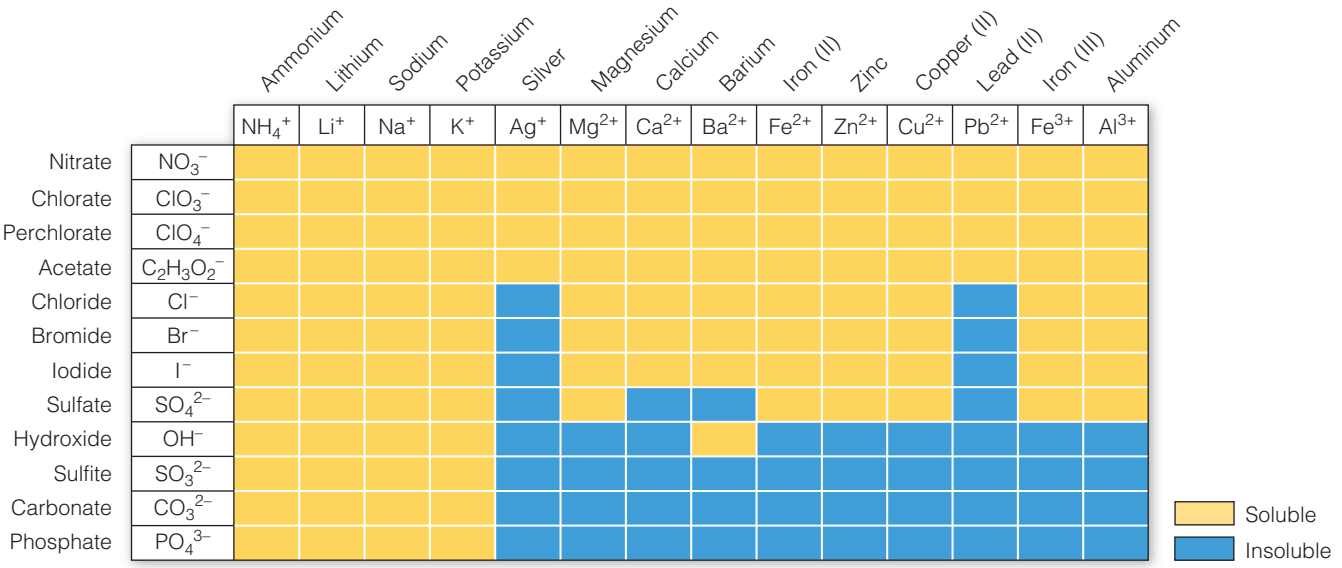
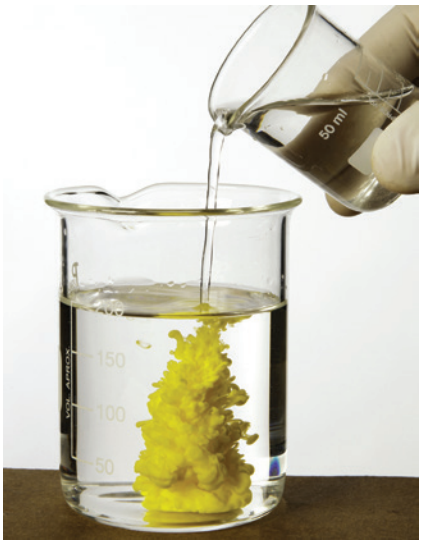


Figure 6.18 A graphical approach to the solubility rules is presented here. Notice that compounds containing the alkali metals, NH_4^+, NO_3^-, ClO_3^-, ClO_4^-, and $C_2H_3O_2^-$ are always soluble.

TABLE 6.3 Solubility Rules

Compounds Containing These Ions Are Always Soluble	
Alkali metals	Li^+, Na^+, K^+, Rb^+
Ammonium	NH_4^+
Large −1 oxyanions	NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$
Compounds Containing These Ions Are Usually Soluble	
Halides (except Pb^{2+}, Ag^+)	F^-, Cl^-, Br^-, I^-
Sulfate (except Ba^{2+}, Ca^{2+}, Pb^{2+}, Ag^+)	SO_4^{2-}
Not Soluble	
Most other ions	



Alexandre Dotta/Science Source

Figure 6.19 Lead(II) iodide is insoluble in water. When a solution containing dissolved Pb^{2+} is combined with a solution containing dissolved I^-, PbI_2 immediately forms as a yellow solid.

Explore Figure 6.19

The different solubilities of ionic compounds cause some dramatic reactions. For example, aqueous lead(II) nitrate and aqueous sodium iodide are both colorless solutions. However, if we combine these solutions, a yellow solid immediately forms (**Figure 6.19**). The yellow solid is lead(II) iodide.

The other product from this reaction, sodium nitrate, remains in solution. We describe this reaction using the following equation:

$$Pb(NO_3)_2\ (aq) + 2\ NaI\ (aq) \rightarrow 2\ NaNO_3\ (aq) + PbI_2\ (s)$$

This type of reaction, in which two aqueous solutions combine to produce a solid (insoluble) product, is called a **precipitation reaction**. The solid product formed in the reaction is the **precipitate**. In the equation above, the precipitate is indicated by the (*s*) notation after the chemical formula. Compounds that remain in solution are shown using the (*aq*) notation. Precipitation reactions are examples of double-displacement reactions.

Precipitation reactions are double-displacement reactions.

It is helpful to think of this reaction in terms of ionic equations. Before the solutions are mixed, each contains dissolved ions:

$$\text{Lead(II) nitrate solution: } Pb^{2+}\ (aq) + 2\ NO_3^-\ (aq)$$

$$\text{Sodium iodide solution: } Na^+\ (aq) + I^-\ (aq)$$

GIPhotoStock/Science Source

Sometimes precipitates form distinctive solids, but at other times they appear as "cloudy" mixtures.

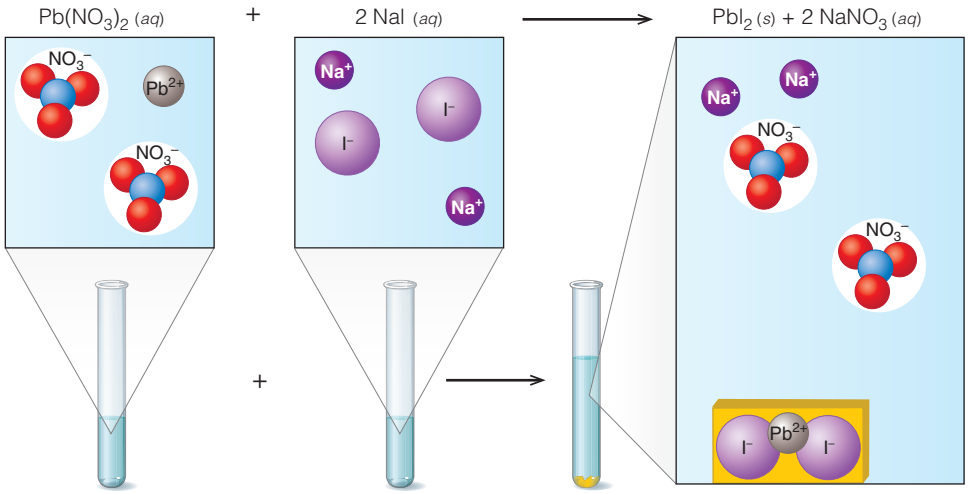


Figure 6.20 The driving force for a precipitation reaction is the formation of the solid. If a cation and anion precipitate out of solution, the other two ions are left in solution together.

When these solutions are combined, all four ions are present in solution. However, as the lead and iodide ions encounter each other (**Figure 6.20**), they bind together and drop out of solution as a solid (precipitate).

We can write this as an ionic equation. Because all of the ions are shown, this is sometimes called a **complete ionic equation**:

$$Pb^{2+}\ (aq) + 2\ NO_3^-\ (aq) + 2\ Na^+\ (aq) + 2\ I^-\ (aq) \rightarrow 2\ Na^+\ (aq) + 2\ NO_3^-\ (aq) + PbI_2\ (s)$$

The driving force for a precipitation reaction is the formation of the solid. Notice in the ionic equation that the nitrate and the sodium ions don't actually react: They are present in the starting materials and are unchanged in the products. The sodium and nitrate ions are paired together only because the other ions (Pb^{2+} and I^-, shown in red in the equation above) have dropped out of the solution. The sodium and nitrate ions are **spectator ions** because they are not changed in the reaction.

When ions combine to form an insoluble compound, a precipitation reaction occurs.

Sometimes it is helpful to show precipitation reactions as a **net ionic equation**. In this type of equation, we omit the spectator ions and include only the ions directly involved in the precipitation (**Figure 6.21**). For the reaction above, the net ionic equation is

$$Pb^{2+}\ (aq) + 2\ I^-\ (aq) \rightarrow PbI_2\ (s)$$

Molecular Equation
Shows neutral compounds

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

Complete Ionic Equation
Shows all ions present

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

Net Ionic Equation
Omits spectator ions; only shows ions that react

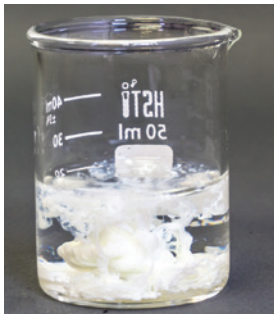
$$Pb^{2+}\ (aq) + 2\ Cl^-\ (aq) \longrightarrow PbCl_2\ (s)$$

Figure 6.21 We use three types of equations to describe aqueous reactions.

The solubility rules are the key to predicting precipitation reactions. If a cation and an anion can combine to form an insoluble product, a precipitation reaction will occur. **Figure 6.22** contains several additional examples of precipitation reactions. Can you identify the products of these reactions?

We use the solubility rules to predict precipitation reactions.

$AgNO_3\ (aq) + KBr\ (aq) \longrightarrow$



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$Na_2CO_3\ (aq) + Fe(NO_3)_2\ (aq) \longrightarrow$

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$Pb(NO_3)_2\ (aq) + Na_2SO_4\ (aq) \longrightarrow$

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Figure 6.22 Each of the reactions shown produces a precipitate. Can you identify the solid formed in each example?

Explore Figure 6.22

Example 6.13 Predicting Precipitation Reactions

Write a balanced equation to show the reaction of aqueous silver acetate with aqueous barium chloride. Include phase symbols.

To solve this problem, we first need to identify the ions that are present:

Silver acetate solution: $Ag^+\ (aq) + C_2H_3O_2^-\ (aq)$

Barium chloride solution: $Ba^{2+}\ (aq) + 2\ Cl^-\ (aq)$

According to the solubility chart (Table 6.3), the compound formed from silver and chloride ions is insoluble. Therefore, one of the products will be AgCl (*s*). We can write this as a balanced, complete ionic equation:

$$2\ Ag^+\ (aq) + 2\ C_2H_3O_2^-\ (aq) + Ba^{2+}\ (aq) + 2\ Cl^-\ (aq) \rightarrow Ba^{2+}\ (aq) + 2\ C_2H_3O_2^-\ (aq) + 2\ AgCl\ (s)$$

By removing the spectator ions, we can also represent this as a net ionic equation:

$$Ag^+\ (aq) + Cl^-\ (aq) \rightarrow AgCl\ (s)$$

Or as a balanced molecular equation:

$$2\ AgC_2H_3O_2\ (aq) + BaCl_2\ (aq) \rightarrow Ba(C_2H_3O_2)_2\ (aq) + 2\ AgCl\ (s)$$

If we tried this reaction in the lab, silver chloride would form as a white precipitate (**Figure 6.23**). •

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Figure 6.23 Silver and chloride ions combine to form a white precipitate.

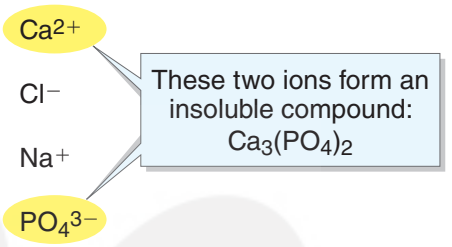
Example 6.14 Predicting and Balancing Precipitation Reactions

Write a balanced molecular equation to show the precipitation reaction of calcium chloride with sodium phosphate. Include phase symbols.

To solve this problem, we first write out the empirical formulas for the two starting materials:

$$CaCl_2 + Na_3PO_4$$

These ionic compounds are both water soluble, so they exist in solution as dissociated ions. Based on Table 6.3, two of these ions, calcium and phosphate, form an insoluble solid. The formula for this compound is $Ca_3(PO_4)_2$:



The other two ions, sodium and chloride, remain in solution. We can describe this reaction with an unbalanced equation, showing the precipitate as a solid:

$$CaCl_2\ (aq) + Na_3PO_4\ (aq) \rightarrow Ca_3(PO_4)_2\ (s) + NaCl\ (aq) \qquad \text{unbalanced}$$

Finally, we balance this equation by balancing the ions on each side:

$$3\ CaCl_2\ (aq) + 2\ Na_3PO_4\ (aq) \rightarrow Ca_3(PO_4)_2\ (s) + 6\ NaCl\ (aq) \qquad \text{balanced}$$

Another approach to solving this problem is to recognize that precipitation reactions are double-displacement reactions. For a precipitation reaction to occur, the cation-anion pairs must swap. We must keep in mind the charges on the ions and write empirical formulas that result in an overall charge of zero. In this instance, the anions swap positions, producing calcium phosphate and sodium chloride:

$$3\ CaCl_2\ (aq) + 2\ Na_3PO_4\ (aq) \rightarrow 3\ Ca_3(PO_4)_2\ (s) + 6\ NaCl\ (aq)$$

The anions "swap" positions.

Referring again to Table 6.3, we see that one of these two products, calcium phosphate, forms an insoluble solid.

While this approach may make the problem simpler, remember that a precipitation reaction cannot occur unless one of the two product compounds is insoluble. If neither product is insoluble, all of the ions remain dissolved in water, and no reaction takes place. ●

TRY **IT**

14. Write a complete ionic equation showing the reaction of aqueous barium nitrate with aqueous sodium phosphate to form barium phosphate and sodium nitrate. Include phase symbols.

Check it *Watch explanation*

15. Refer to Table 6.3 to identify the precipitate formed in the following reactions. Balance the reactions.

$$KOH\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow$$

$$FeCl_2\ (aq) + Li_2S\ (aq) \rightarrow$$

$$MgBr_2\ (aq) + K_3PO_4\ (aq) \rightarrow$$

Check it *Watch explanation*

16. In which of these combinations would no precipitation reaction occur?

a. Ammonium iodide is combined with silver acetate.

b. Potassium chloride is combined with sodium hydroxide.

c. Zinc sulfate is combined with calcium chloride.

Check it *Watch explanation*

Acid-Base Neutralization Reactions

Acids and Bases

In Chapter 5 we introduced the very important class of compounds called *acids*. Recall that acids are covalent compounds that ionize to produce H^+ ions and a stable anion when dissolved in water. A list of common acids is given in **Table 6.4**.

We often use ionic equations to show how acids ionize in water. For example, aqueous hydrochloric acid (HCl) and aqueous nitric acid (HNO_3) ionize as follows:

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

$$HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$$

Bases are compounds that produce hydroxide (OH^-) ions in aqueous solution. Many common bases are soluble metal hydroxides (**Table 6.5**). In water, these compounds dissociate to give metal cations and hydroxide anions. For example, sodium hydroxide dissociates in water to produce sodium and hydroxide ions:

$$NaOH\ (s) \rightarrow Na^+\ (aq) + OH^-\ (aq)$$

TABLE 6.4 Common Acids

Formula	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
H_2CO_3	Carbonic acid
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic acid

TABLE 6.5 Common Hydroxide Bases

Formula	Name
LiOH	Lithium hydroxide
NaOH	Sodium hydroxide
KOH	Potassium hydroxide
$Ba(OH)_2$	Barium hydroxide

In aqueous solution, acids produce H^+ ions and bases produce OH^- ions.

Neutralization Reactions

Acids and bases undergo **neutralization reactions.** In these reactions, the H^+ from the acid combines with the OH^- from the base to form water. We can show this as a net ionic equation:

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

For example, hydrochloric acid reacts with sodium hydroxide. We can write this as a molecular equation or as a complete ionic equation:

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

$$H^+ (aq) + Cl^- (aq) + Na^+ (aq) + OH^- (aq) \rightarrow H_2O (l) + Na^+ (aq) + Cl^- (aq)$$

The spectator ions, Na^+ and Cl^-, combine to form an ionic compound. Similarly, nitric acid reacts with lithium hydroxide:

$$HNO_3 (aq) + LiOH (aq) \rightarrow H_2O (l) + LiNO_3 (aq)$$

Chemists sometimes refer to ionic compounds as *salts*.

Again, the H^+ from the acid reacts with the OH^- from the base to form water. The remaining ions end up together as $LiNO_3$. The ionic compound that is formed from the spectator ions is often called a *salt*.

In general, the following expression describes the neutralization of an acid with a hydroxide base:

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

The formation of water is the driving force for a neutralization reaction.

Like precipitation reactions, acid-base neutralizations are double-displacement reactions: The H^+ and OH^- combine to form water while the other two ions combine to form the salt. The driving force of the reaction is the formation of water; the salt forms from the other two ions.

Acid-base neutralizations are double-displacement reactions. The H^+ and OH^- combine to form water (HOH), and the two spectator ions combine to form a salt:

$$HCl + NaOH \rightarrow HOH + NaCl$$

Example 6.15 Predicting the Products from Acid-Base Neutralization Reactions

Write a balanced equation to show the reaction of sulfuric acid with potassium hydroxide. Include phase symbols.

We begin by writing an unbalanced equation to describe this reaction. One product is water, and the other product is the ionic compound produced by the spectator ions:

$$H_2SO_4 + NaOH \rightarrow H_2O + Na_2SO_4 \qquad \text{unbalanced}$$

A simple way to balance neutralization reactions is to balance the number of H^+ and OH^- ions that combine. Because H_2SO_4 produces two H^+ ions, there must be two OH^- ions to neutralize them, and two water molecules will be produced. Adding phase symbols gives us the balanced form shown here. ●

$$H_2SO_4 (aq) + 2\ NaOH (aq) \rightarrow 2\ H_2O (l) + Na_2SO_4 (aq) \qquad \text{balanced}$$

TRY IT

17. Predict the products for the following neutralization reactions. Make sure the equations are balanced.

$$HBr + NaOH \rightarrow$$

$$HClO_4 (aq) + Ba(OH)_2 (aq) \rightarrow$$

$$H_3PO_4 (aq) + KOH (aq) \rightarrow$$

Check it *Watch explanation*

18. Write a complete ionic equation and a net ionic equation for the following neutralization reaction:

$$LiOH\ (aq) + HNO_3\ (aq) \rightarrow LiNO_3\ (aq) + H_2O\ (l)$$

Check it *Watch explanation*

SUMMARY

Scientists describe chemical changes, or reactions, using chemical equations. Chemical equations show the substances present before the reaction (called the reactants) on the left side of the arrow, and the substances present after the reaction (called the products) on the right side of the arrow.

Atoms are not created or destroyed in chemical reactions. To reflect this, we write balanced equations. An equation is balanced if it contains the same number and type of atom in the reactants and in the products. Balancing equations often involves trial and error and requires practice to do it efficiently. Sometimes chemical equations also show the phase of the reactants and products in parentheses after each element or compound.

There are many types of chemical reactions and different ways of organizing reactions. Many reactions can be classified as decomposition, synthesis, single-displacement, and double-displacement reactions. These classifications are based on the number and type of reactant and product formed.

We also classify reactions by patterns of chemical behavior. In this chapter, we examined four broad classes of reactions: metal-nonmetal reactions, combustion reactions, precipitation reactions, and acid-base neutralization reactions (**Table 6.6**).

Oxidation-reduction reactions involve the gain and loss of electrons. An atom or ion that loses electrons is oxidized; a species that gains electrons is reduced. Metals and nonmetals combine in these reactions to produce ionic compounds.

Many elements react with oxygen to produce oxide compounds. Similarly, hydrocarbons (compounds composed of hydrogen and carbon) react with oxygen to form carbon dioxide and water. Collectively, reactions with oxygen are called *combustion reactions*.

When an ionic compound dissolves in water, the ions separate in a process called *dissociation*. Not all ionic compounds dissolve in water. Solubility rules provide broad guidelines for determining which compounds do and do not dissolve in water.

Precipitation reactions occur when two aqueous solutions combine to produce an insoluble product. The insoluble product drops out of solution (precipitates), leaving the other ions—called spectator ions—remaining in

TABLE 6.6 Patterns of Chemical Behavior

Reaction Type	Classification	Starting Materials	Products
Metal-Nonmetal	Combination	Metal + Nonmetal	Ionic compound
Combustion		Element + Oxygen Hydrocarbon + Oxygen	Oxide compound $CO_2 + H_2O$
Precipitation reaction	Double displacement	Two soluble compounds	Insoluble compound (precipitate)
Acid-Base neutralization	Double displacement	Acid + base	Water + salt

solution. We commonly describe precipitation reactions using ionic equations. Complete ionic equations show all ions present while net ionic equations show only those directly involved in the central reaction.

Acids are covalent compounds that ionize in water to produce hydrogen cations (H^+) and a corresponding anion. Bases are compounds that produce hydroxide anions (OH^-) when dissolved in water. The most common bases are metal hydroxides. When acids and bases combine, they undergo a neutralization reaction in which the H^+ from the acid reacts with the OH^- from the base to form water. Like precipitation reactions, acid-base reactions are examples of double-displacement reactions.

The Mayan Lime Cycle

We began this chapter with the great Mayan ruins of Palenque. Now that we have a broader understanding of chemical reactions, let's look at the specific process the Mayans used to make plaster for their temples and palaces (**Figure 6.24**):

The first step in lime production is heating. Above about 900 °C, calcium carbonate decomposes to produce calcium oxide and carbon dioxide:

$$CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$$

Calcium oxide is a fine white powder, commonly called *quicklime*. The calcium oxide is then mixed with water to produce calcium hydroxide, also called *slaked lime*, in a synthesis reaction:

$$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (s)$$

The slaked lime is often mixed with some other material, such as volcanic ash, that strengthens the overall structure. As this lime plaster dries, it reacts with carbon dioxide in the air to return to calcium carbonate:

$$Ca(OH)_2\ (s) + CO_2\ (g) \rightarrow CaCO_3\ (s) + H_2O\ (l)$$

This three-step process is called the *lime cycle*. But what does this have to do with the decline of Mayan cities?

Evidence from the hieroglyphs at Palenque suggests that a major building project was completed just before the population began to decline. To produce enough lime for this project, the Maya needed massive amounts of firewood. To obtain the firewood, they cleared huge swaths of the Yucatán jungles. Many archaeologists believe that this deforestation harmed the water supplies and

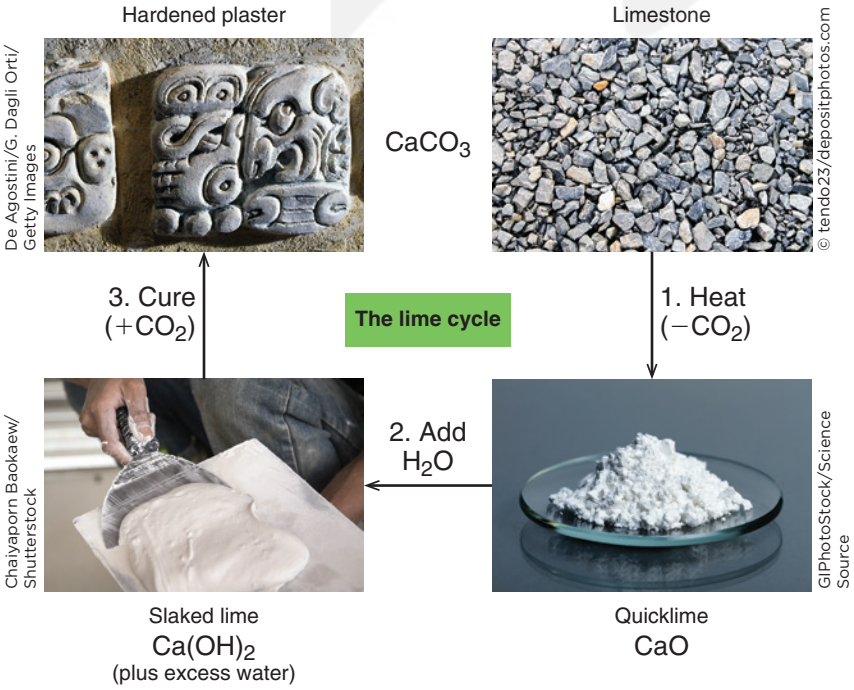
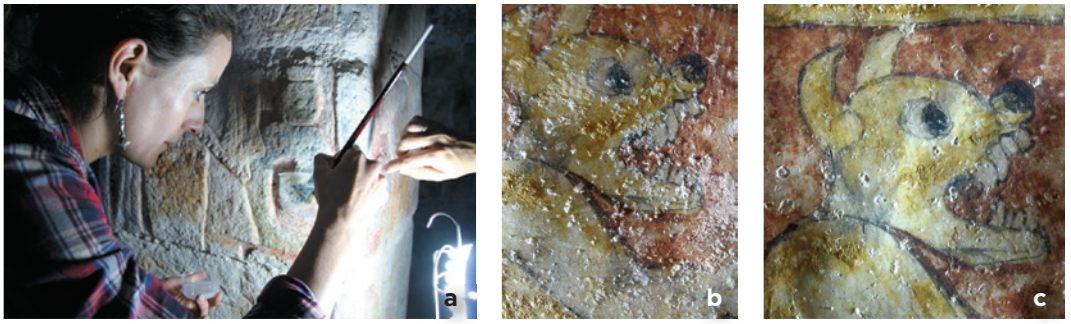


Figure 6.24 Through the lime cycle, natural limestone ($CaCO_3$) is converted into CaO, then into $Ca(OH)_2$. $Ca(OH)_2$ can be molded or shaped before it cures, reforming the hard $CaCO_3$.



Photograph by Yareli Jáidar, CNCPC-INAH

Figure 6.25 (a) Dr. Yareli Jáidar applies a coat of nanolime to a Mayan structure. (b) This image shows a jaguar painting that has degraded over time. (c) This is the same painting after restoration.

promoted widespread drought. This drought may have led to the decline in population and ultimately to the abandonment of cities. Years later, the jungles slowly grew back and overtook the ancient structures.

Today, archaeological conservators like Yareli Jáidar work to preserve archaeological and artistic treasures (**Figure 6.25**). She often uses a modern material called *nanolime*. Nanolime is composed of very tiny particles of calcium hydroxide, often just 100 nanometers across (from which the name originates). When conservationists coat a plaster structure with nanolime, the tiny lime particles penetrate the painting surface to fill pores and cracks. As the lime dries, it fuses with the ancient lime, sealing the structures against further damage.

Key Terms

6.1 Chemical Equations

chemical equation A symbolic representation of a chemical change. Such equations consist of reactants and products separated by an arrow.

reactant The starting material in a chemical change, shown on the left-hand side of a chemical equation.

product The compounds produced in a chemical change, shown on the right-hand side of a chemical equation.

subscript In a chemical formula, subscript numbers show the number of each atom or ion present.

coefficient In a chemical formula, the numbers written before each reactant or product to indicate the ratios in which components of the reaction are consumed or produced.

balanced equation A chemical equation in which the number and type of atom are the same for the reactants and the products.

6.2 Classifying Reactions

decomposition A reaction in which a single reactant forms two or more products.

synthesis A reaction in which two reactants join together to form a single product; also called a *combination reaction*.

single displacement A reaction in which one element replaces another element in a compound.

double displacement A reaction in which two compounds swap cation-anion pairs to form two new compounds.

6.3 Reactions between Metals and Nonmetals

oxidation-reduction reaction A chemical change in which one species loses electrons (oxidation) while another gains electrons (reduction).

oxidation The loss of electrons.

reduction The gain of electrons.

6.4 Combustion Reactions

combustion A reaction in which oxygen gas combines with elements or compounds to produce oxide compounds.

6.5 Reactions in Aqueous Solution

molecular equation A chemical equation in which all species are written as neutral compounds.

ionic equation A chemical equation that shows dissociated ions as separate species.

precipitation reaction A type of chemical change in which two aqueous solutions combine to produce an insoluble product.

precipitate A solid product formed from the combination of two solutions.

complete ionic equation An equation that shows all ions present in a solution.

spectator ion An ion that is present in a solution but not directly involved in a chemical change.

net ionic equation An equation in which the only ions shown are those directly involved in the chemical change, and spectator ions are omitted.

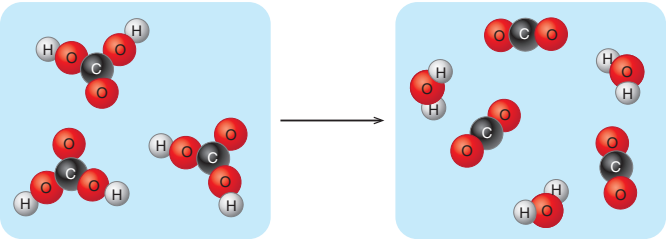
base A compound that produces hydroxide ions in aqueous solutions. In later chapters, bases will be further defined as proton (H^+) acceptors.

neutralization reaction A reaction in which an acid and a base combine to produce water and an ionic compound (called a *salt*).

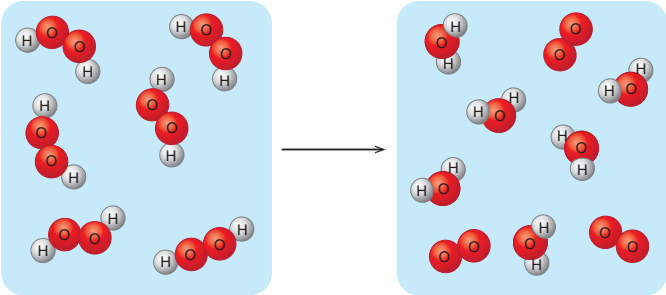
Additional Problems

6.1 Chemical Equations

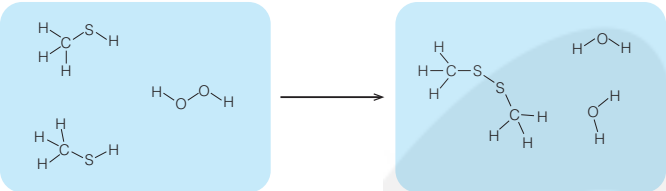
19. Write a balanced equation to represent the reaction shown. Write the coefficients as the lowest whole-number ratio.



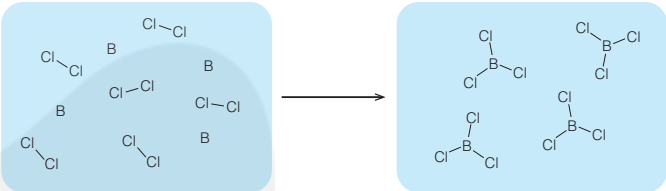
20. Write a balanced equation to represent the reaction shown. Write the coefficients as the lowest whole-number ratio.



21. Write a balanced equation to represent the reaction shown.



22. Write a balanced equation to represent the reaction shown. Write the coefficients as the lowest whole-number ratio.



23. Describe each of the following changes using chemical equations:

a. One molecule of hydrogen reacts with one molecule of bromine to produce two molecules of hydrogen bromide.
b. Two atoms of sodium and one molecule of fluorine react to produce two molecules of sodium fluoride.

24. Describe each of the following changes using chemical equations:

a. One unit of calcium hydroxide reacts to form one unit of calcium oxide and one molecule of water.
b. Four iron atoms react with three oxygen molecules to form two units of iron(III) oxide.

25. In this reaction, how many units of C form when 30 units of A react?

$$2\ A + B \rightarrow C$$

26. In this reaction, how many units of B are required to form 40 units of D?

$$A + 2\ B \rightarrow C + D$$

27. The balanced equation here shows the reaction of hydrogen gas with chlorine gas. Based on this equation,

$$H_2 + Cl_2 \rightarrow 2\ HCl$$

a. How many molecules of HCl form from 30 molecules of H_2?
b. How many molecules of Cl_2 are required to produce 12 molecules of HCl?

28. The balanced equation here shows the reaction of calcium with oxygen to produce calcium oxide. Based on this equation,

$$2\ Ca + O_2 \rightarrow 2\ CaO$$

a. How many molecules of oxygen react with 18 calcium atoms?
b. If 30 calcium atoms react with 15 oxygen molecules, how many units of calcium oxide form?

29. The reaction of aluminum with chlorine gas is shown here. Based on this equation,

$$2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$$

a. How many molecules of chlorine gas are needed to react with 10 aluminum atoms?
b. How many units of $AlCl_3$ can be produced from 10 aluminum atoms?

30. Hydrobromic acid (HBr) can react with barium hydroxide as shown here. Based on this equation,

$$2\ HBr + Ba(OH)_2 \rightarrow BaBr_2 + 2\ H_2O$$

a. If 5 units of barium hydroxide react in this way, how many water molecules form?
b. How many HBr molecules are needed to produce 40 units of $BaBr_2$?

31. Balance each of the following equations:

a. $PCl_3 + F_2 \rightarrow PF_3 + Cl_2$
b. $SO_2 + O_2 \rightarrow SO_3$
c. $B + F_2 \rightarrow BF_3$

32. Balance each of the following equations:

a. $Zn + HCl \rightarrow ZnCl_2 + H_2$
b. $Li + CuCl_2 \rightarrow LiCl + Cu$
c. $C_3H_4O + H_2 \rightarrow C_3H_8O$

33. Balance the following equations:

a. $PCl_3 + H_2O \rightarrow H_3PO_3 + HCl$
b. $O_2 + H_2S \rightarrow SO_2 + H_2O$
c. $HCl + MnO_2 \rightarrow MnCl_2 + 2H_2O + Cl_2$

34. Balance the following equations:

a. $TiCl_4 + H_2O \rightarrow TiO_2 + HCl$
b. $VCl_3 \rightarrow VCl_2 + VCl_4$
c. $Cr_2O_3 + Al \rightarrow Al_2O_3 + Cr$

35. Balance the following equations by balancing the ions:

a. $Hg(NO_3)_2 + KCl \rightarrow KNO_3 + HgCl_2$
b. $MgBr_2 + NaOH \rightarrow NaBr + Mg(OH)_2$
c. $AgC_2H_3O_2 + BaCl_2 \rightarrow AgCl + Ba(C_2H_3O_2)_2$

36. Balance the following equations by balancing the ions:

a. $Cr(NO_3)_2 + KOH \rightarrow KNO_3 + Cr(OH)_2$
b. $(NH_4)_2CO_3 + Ba(C_2H_3O_2)_2 \rightarrow NH_4C_2H_3O_2 + BaCO_3$
c. $AgC_2H_3O_2 + BaCl_2 \rightarrow AgCl + Ba(C_2H_3O_2)_2$

37. Balance the following equations. For these equations, it may be helpful to write a fractional coefficient for the diatomic element before converting to whole numbers.

a. $N_2 + H_2 \rightarrow NH_3$
b. $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
c. $HCl + Al \rightarrow AlCl_3 + H_2$

38. Balance the following equations. For these equations, it may be helpful to write a fractional coefficient for the diatomic element before converting to whole numbers.

a. $P + Cl_2 \rightarrow PCl_5$
b. $IrF_5 + H_2 \rightarrow IrF_4 + HF$
c. $C_5H_{10} + O_2 \rightarrow CO_2 + H_2O$

39. Balance the following equations:

a. $Na + O_2 \rightarrow Na_2O$
b. $Pb(NO_3)_2 + KCl \rightarrow PbCl_2 + KNO_3$
c. $C_2H_2 + O_2 \rightarrow CO_2 + H_2O$

40. Balance the following equations:

a. $NaOCl + HCl \rightarrow Cl_2 + NaCl + H_2O$
b. $AgC_2H_3O_2 + MgCl_2 \rightarrow AgCl + Mg(C_2H_3O_2)_2$
c. $N_2H_6 + O_2 \rightarrow NO_2 + H_2O$

41. Write a balanced equation to describe each of these chemical changes. Include phase symbols.

a. Zinc metal reacts with aqueous copper(II) chloride to produce copper metal and aqueous zinc chloride.
b. Propane gas, C_3H_8, reacts with molecular oxygen to produce carbon dioxide gas and water vapor.
c. Solid iron reacts with chlorine gas to produce solid iron(II) chloride.

42. Write a balanced equation to describe each of these chemical changes. Include phase symbols.

a. When calcium metal is heated in water, aqueous calcium hydroxide and hydrogen gas form.
b. When solid magnesium burns (reacts with oxygen gas), it forms solid magnesium oxide.
c. Aqueous silver nitrate reacts with aqueous sodium chloride to produce solid silver chloride and aqueous sodium nitrate.

43. Provide the name and the phase of the reactants for each reaction:

a. $2\,HCl\,(aq) + Na_2CO_3\,(s) \rightarrow 2\,NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)$
b. $Fe\,(s) + 2\,HNO_3\,(aq) \rightarrow Fe(NO_3)_2\,(aq) + H_2\,(g)$
c. $ZnCl_2\,(aq) + Pb(ClO_4)_2\,(aq) \rightarrow PbCl_2\,(s) + Zn(ClO_4)_2\,(aq)$

44. Provide the name and phase of the products for each reaction:

a. $P\,(s) + Cl_2\,(g) \rightarrow PCl_3\,(g)$
b. $Na_2SO_4\,(aq) + BaBr_2\,(aq) \rightarrow BaSO_4\,(s) + 2\,NaBr\,(aq)$
c. $Cr\,(s) + 2\,HCl\,(aq) \rightarrow CrCl_2\,(aq) + H_2\,(g)$

6.2 Classifying Reactions

45. Classify each of these chemical reactions as a synthesis, decomposition, single-displacement, or double-displacement reaction:

a. $3\,NH_3 \rightarrow 3\,N_2 + 3\,H_2$
b. $CaO + H_2O \rightarrow Ca(OH)_2$
c. $AgNO_3\,(aq) + KBr\,(aq) \rightarrow KNO_3\,(aq) + AgBr\,(s)$

46. Classify each of these chemical reactions as a synthesis, decomposition, single-displacement, or double-displacement reaction:

a. $Sn + 2\,Br_2 \rightarrow SnBr_4$
b. $2\,HBr + Sn \rightarrow SnBr_2 + H_2$
c. $HBr + NaOH \rightarrow NaBr + H_2O$

47. Classify each of these chemical reactions as a synthesis, decomposition, single-displacement, or double-displacement reaction:

a. $Ca + O_2 \rightarrow 2\,CaO$
b. $Co\,(s) + 2\,HCl\,(aq) \rightarrow CoCl_2\,(aq) + H_2\,(g)$
c. $PbNO_3\,(aq) + 2\,KBr\,(aq) \rightarrow 2\,KNO_3\,(aq) + PbBr_2\,(s)$

48. Classify each of these chemical reactions as a synthesis, decomposition, single-displacement, or double-displacement reaction:

a. $K_2SO_4\,(aq) + Pb(C_2H_3O_2)_2\,(aq) \rightarrow 2\,KC_2H_3O_2\,(aq) + PbSO_4\,(s)$
b. $2\,Zn + O_2 \rightarrow 2\,ZnO$
c. $PCl_5 \rightarrow PCl_3 + Cl_2$

6.3 Reactions between Metals and Nonmetals

49. What types of compounds form when metals and nonmetals react?

50. What is oxidation? Do metals or nonmetals oxidize more easily?

51. In these reactions, identify the element that is oxidized and the element that is reduced:

a. $2\ Mg + O_2 \rightarrow 2\ MgO$
b. $Fe + S \rightarrow FeS$
c. $Br_2 + Ca \rightarrow CaBr_2$

52. In these reactions, identify the element that is oxidized and the element that is reduced:

a. $2\ Cu + O_2 \rightarrow 2\ CuO$
b. $S + Hg \rightarrow HgS$
c. $Sn + 2\ Cl_2 \rightarrow SnCl_4$

53. Predict the products from these reactions, and balance the equations:

a. $Be + Cl_2 \rightarrow$
b. $K + Cl_2 \rightarrow$
c. $Co\ (s) + Cl_2\ (g) \rightarrow$
d. $Cu\ (s) + O_2\ (g) \rightarrow$

54. Predict the products from these reactions, and balance the equations:

a. $Ca + O_2 \rightarrow$
b. $Sr + Br_2 \rightarrow$
c. $Pb\ (s) + O_2\ (g) \rightarrow$
d. $Cr\ (s) + O_2\ (g) \rightarrow$

55. Predict the products from these combustion reactions, and balance the equations:

a. $Mg + O_2 \rightarrow$
b. $C_2H_4 + O_2 \rightarrow$
c. $C_4H_{10}\ (g) + O_2\ (g) \rightarrow$

56. Predict the products from these combustion reactions, and balance the equations:

a. $Zn + O_2 \rightarrow$
b. $C_5H_{12} + O_2 \rightarrow$
c. $C_8H_{18}\ (l) + O_2\ (g) \rightarrow$

6.4 Combustion Reactions

57. What element is always involved in combustion reactions?

58. Each of the following elements reacts with oxygen in combustion reactions. Identify whether the products of these reactions would be ionic compounds or covalent compounds.

a. magnesium b. calcium c. carbon d. nitrogen

59. What are hydrocarbons? What products form from the combustion of hydrocarbons?

60. Fossil fuels like coal contain primarily hydrocarbons, but they also contain a small amount of other elements, like sulfur. What products result from the combustion of hydrocarbons? What products result from the combustion of sulfur?

61. Write a balanced equation for each of these reactions:

a. combustion of C_2H_2
b. combustion of C_4H_8
c. combustion of C_9H_{18}

62. Write a balanced equation for each of these reactions:

a. combustion of CH_4
b. combustion of C_3H_8
c. combustion of C_5H_{10}

6.5 Reactions in Aqueous Solution

63. The following molecular equations each show an ionic compound dissolving in water. Show the ionization process by rewriting each one as an ionic equation:

a. $KCl\ (s) \rightarrow KCl\ (aq)$
b. $Li_2SO_4\ (s) \rightarrow Li_2SO_4\ (aq)$
c. $(NH_4)_2CO_3\ (s) \rightarrow (NH_4)_2CO_3\ (aq)$

64. The following molecular equations each show an ionic compound dissolving in water. Show the ionization process by rewriting each one as an ionic equation:

a. $CaBr_2\ (aq) \rightarrow CaBr_2\ (aq)$
b. $LiNO_3\ (s) \rightarrow LiNO_3\ (aq)$
c. $Cu(ClO_4)_2\ (s) \rightarrow Cu(ClO_4)_2\ (aq)$

65. Write ionic equations to show the dissociation of each solid as it is dissolved in water:

a. sodium bicarbonate
b. aluminum chloride
c. chromium(III) chlorite

66. Write ionic equations to show the dissociation of each solid as it is dissolved in water:

a. ammonium cyanide
b. tin(IV) bromide
c. cobalt(II) chlorate

67. Using the general solubility rules in Section 6.5, state whether each of these compounds is likely to be soluble or insoluble:

a. an ionic compound containing the ammonium ion
b. an ionic compound containing a +2 cation and a −3 anion
c. an ionic compound containing a nitrate ion

68. Using the general solubility rules in Section 6.5, state whether each of these compounds is likely to be soluble or insoluble:

a. an ionic compound composed of an alkaline earth metal and the oxide ion
b. an ionic compound containing a transition metal ion and an acetate ion
c. an ionic compound containing a silver ion and a halogen ion

69. Using the general solubility rules in Section 6.5, state whether each of these compounds is likely to be soluble or insoluble:

a. an ionic compound composed of ammonium and a chlorate ion
b. an ionic compound containing a transition metal ion and a phosphate ion
c. an ionic compound composed of a transition metal cation and a perchlorate anion

70. Using the general solubility rules in Section 6.5, state whether each of these compounds is likely to be soluble or insoluble:

a. an ionic compound composed of an alkaline earth metal and the sulfide ion
b. an ionic compound containing an aluminum cation and a halogen anion
c. an ionic compound containing a lead(II) cation and a halogen anion

71. Refer to the general solubility rules or Table 6.3 to determine whether each of these compounds is soluble or insoluble in water:

a. LiOH
b. K_2S
c. $CaSO_4$
d. zinc sulfate
e. copper(II) nitrate
f. $Fe(C_2H_3O_2)_3$

72. Refer to the general solubility rules or Table 6.3 to determine whether each of these compounds is soluble or insoluble in water:

a. NH_4OH
b. $BaCl_2$
c. $FePO_4$
d. iron(II) chloride
e. lead(II) chloride
f. aluminum sulfide

73. In general, which would you expect to be more water soluble: a salt with a +1 cation and a –2 anion, or a salt with a +2 cation and a –2 anion?

74. Which would you expect to be more soluble, potassium nitrite (KNO_2) or aluminum oxide (Al_2O_3)? Why?

75. These equations represent precipitation reactions. Rewrite them as complete ionic equations.

a. $AgNO_3\ (aq) + KCl\ (aq) \rightarrow KNO_3\ (aq) + AgCl\ (s)$
b. $Ba(ClO_4)_2\ (aq) + K_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + 2\ KClO_4\ (aq)$

76. These equations represent precipitation reactions. Rewrite them as complete ionic equations.

a. $MgBr_2\ (aq) + Pb(ClO_3)_2\ (aq) \rightarrow PbBr_2\ (s) + Mg(ClO_3)_2\ (aq)$
b. $K_3PO_4\ (aq) + AlCl_3\ (aq) \rightarrow 3\ KCl\ (aq) + AlPO_4\ (s)$

77. Identify the spectator ions in this ionic equation:

$$Ca^{2+}\ (aq) + 2\ NO_3^-\ (aq) + 2\ Cs^+\ (aq) + CO_3^{2-}\ (aq) \rightarrow CaCO_3\ (s) + 2\ Cs^+\ (aq) + 2\ NO_3^-\ (aq)$$

78. Identify the spectator ions in this ionic equation:

$$2\ NH_4^+\ (aq) + S^{2-}\ (aq) + Zn^{2+}\ (aq) + 2\ ClO_3^-\ (aq) \rightarrow 2\ NH_4^+\ (aq) + 2\ ClO_3^-\ (aq) + ZnS\ (s)$$

79. Identify the spectator ions in this equation, and rewrite it as a net ionic equation:

$$Na^+\ (aq) + Br^-\ (aq) + Ag^+\ (aq) + NO_3^-(aq) \rightarrow AgBr\ (s) + Na^+\ (aq) + NO_3^-\ (aq)$$

80. Identify the spectator ions in this equation, and rewrite it as a net ionic equation:

$$Fe^{2+}\ (aq) + 2\ NO_3^-\ (aq) + 2\ K^+\ (aq) + 2\ OH^-\ (aq) \rightarrow Fe(OH)_2\ (s) + 2\ K^+\ (aq) + 2\ NO_3^-\ (aq)$$

81. Aqueous solutions of potassium hydroxide and iron chloride react in a precipitation reaction to form insoluble iron(II) hydroxide, as shown here:

$$2\ KOH\ (aq) + FeCl_2\ (aq) \rightarrow 2\ KCl\ (aq) + Fe(OH)_2\ (s)$$

a. Rewrite this equation as an ionic equation.
b. What are the spectator ions in this equation?
c. Remove the spectator ions, and rewrite this as a net ionic equation.
d. What is the driving force for this reaction?

82. Aqueous solutions of silver nitrate and potassium bromide react in a precipitation reaction to form insoluble silver bromide, as shown here:

$$AgNO_3\ (aq) + KBr\ (aq) \rightarrow KNO_3\ (aq) + AgBr\ (s)$$

a. Rewrite this equation as an ionic equation.
b. What are the spectator ions in this equation?
c. Remove the spectator ions, and rewrite this as a net ionic equation.
d. What is the driving force for this reaction?

83. Each of the following reactions results in one water-soluble product and one precipitate. Complete and balance each reaction, and show phases to indicate whether the products are aqueous or solid.

a. $KCl\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow$
b. $KOH\ (aq) + FeCl_3\ (aq) \rightarrow$
c. $BaCl_2\ (aq) + K_3PO_4\ (aq) \rightarrow$

84. Each of the following reactions results in one water-soluble product and one precipitate. Complete and balance each reaction, and show phases to indicate whether the products are aqueous or solid.

a. $K_2SO_4\ (aq) + FeBr_3\ (aq) \rightarrow$
b. $ZnCl_2\ (aq) + AgC_2H_3O_2\ (aq) \rightarrow$
c. $Pb(C_2H_3O_2)_2\ (aq) + MgBr_2\ (aq) \rightarrow$

85. Write balanced equations for the following precipitation reactions, including phase symbols to indicate the insoluble product. If no precipitate forms, indicate "No reaction."

a. Iron(III) acetate reacts with barium sulfide.
b. Lithium phosphate reacts with copper(II) sulfate.
c. Calcium iodide reacts with ammonium acetate.
d. Iron(II) sulfate reacts with barium hydroxide.

86. Write balanced equations for the following precipitation reactions, including phase symbols to indicate the insoluble product. If no precipitate forms, indicate "No reaction."

a. Sodium chloride reacts with silver nitrate.
b. Lead(II) nitrate reacts with sodium iodide.
c. Ammonium phosphate reacts with aluminum chloride.
d. Barium hydroxide reacts with iron(III) chloride.

87. Write molecular equations, complete ionic equations, and net ionic equations to describe the following precipitation reactions. Include phase symbols.

a. reaction of aqueous lithium iodide with aqueous silver nitrate
b. reaction of aqueous lead(II) acetate with aqueous zinc sulfate
c. reaction of aqueous calcium iodide with aqueous lead(II) nitrate

88. Write molecular equations, complete ionic equations, and net ionic equations to describe the following precipitation reactions. Include phase symbols.

a. reaction of aqueous lead(II) nitrate with aqueous sodium bromide
b. reaction of aqueous iron(II) perchlorate with aqueous sodium carbonate
c. reaction of aqueous aluminum sulfate with aqueous lead(III) nitrate

89. What is the difference between an acid and a base?

90. What are the two common products of acid-base neutralization reactions?

91. Write ionic equations to show the dissociation of each of these acids:

a. HCl b. HNO_3

92. Write ionic equations to show the dissociation of each of these acids:

a. HBr b. HNO_2

93. Balance these neutralization reactions:

a. $H_2SO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$
b. $HNO_3 + Al(OH)_3 \rightarrow Al(NO_3)_3 + H_2O$
c. $H_2SO_4 + Cr(OH)_3 \rightarrow Cr_2(SO_4)_3 + H_2O$

94. Balance these neutralization reactions:

a. $HNO_3 + Ba(OH)_2 \rightarrow Ba(NO_3)_2 + H_2O$
b. $HF + Ba(OH)_2 \rightarrow BaF_2 + H_2O$
c. $NaOH + H_3PO_4 \rightarrow Na_3PO_4 + H_2O$

95. Show the products from these neutralization reactions. Balance the equations if necessary.

a. $HCl + KOH \rightarrow$
b. $H_2SO_4 + 2\ LiOH \rightarrow$
c. $Ba(OH)_2\ (aq) + 2\ HNO_2\ (aq) \rightarrow$

96. Show the products from these neutralization reactions. Balance the equations if necessary.

a. $HBr + NaOH \rightarrow$
b. $H_3PO_4\ (aq) + 3\ KOH\ (aq) \rightarrow$
c. $Ba(OH)_2 + H_2SO_4 \rightarrow$

97. Write a balanced, complete ionic equation showing the reaction of sodium hydroxide with hydroiodic acid.

98. Write a balanced, complete ionic equation showing the neutralization of sulfuric acid with potassium hydroxide.

99. Write a net ionic equation for a double-displacement reaction in which lead(II) iodide is a precipitate.

100. Write a net ionic equation for the neutralization of any acid with a hydroxide base.

Cumulative Questions

101. Classify the reactions below as synthesis, decomposition, single displacement, double displacement, precipitation, acid-base neutralization, oxidation-reduction, or combustion. Some reactions can be classified with more than one of these terms.

a. $Hg\ (l) + S\ (s) \rightarrow HgS\ (s)$
b. $Fe\ (s) + Cu(C_2H_3O_2)_2\ (aq) \rightarrow Cu\ (s) + Fe(C_2H_3O_2)_2\ (aq)$
c. $Pb(NO_3)_2\ (aq) + 2\ KI\ (aq) \rightarrow PbI_2\ (s) + 2\ KNO_3\ (aq)$
d. $2\ K\ (s) + 2\ H_2O\ (l) \rightarrow 2\ KOH\ (aq) + H_2\ (g)$
e. $H_2SO_4\ (aq) + Ba(OH)_2\ (aq) \rightarrow 2\ H_2O\ (l) + BaSO_4\ (s)$
f. $C_8H_{16}\ (l) + 12\ O_2\ (g) \rightarrow 8\ CO_2\ (g) + 16\ H_2O\ (g)$

102. Classify the reactions below as synthesis, decomposition, single displacement, double displacement, precipitation, acid-base neutralization, oxidation-reduction, or combustion. Some reactions can be classified with more than one of these terms.

a. $Fe\ (s) + 2\ AgC_2H_3O_2\ (aq) \rightarrow Fe(C_2H_3O_2)_2\ (aq) + 2\ Ag\ (s)$
b. $AgNO_3\ (aq) + KCl\ (aq) \rightarrow AgCl\ (s) + KNO_3\ (aq)$
c. $HBr\ (aq) + KOH\ (aq) \rightarrow KBr\ (aq) + H_2O\ (l)$
d. $2\ C_2H_2\ (g) + 5\ O_2\ (g) \rightarrow 4\ CO_2\ (g) + 2\ H_2O\ (g)$
e. $Fe\ (s) + 2\ HCl\ (aq) \rightarrow FeCl_2\ (aq) + H_2\ (g)$
f. $2\ K\ (s) + Br_2\ (g) \rightarrow 2\ KBr\ (s)$

103. The reactions here draw from all the reaction types introduced in this chapter. Predict the products, and balance each equation.

a. $AgC_2H_3O_2\ (aq) + NaCl\ (aq) \rightarrow$
b. $Na\ (s) + Cl_2\ (g) \rightarrow$
c. $HCl\ (aq) + Ba(OH)_2\ (aq) \rightarrow$
d. $C_{10}H_{20}\ (l) + O_2\ (g) \rightarrow$

104. The reactions here draw from all the reaction types introduced in this chapter. Predict the products, and balance each equation.

a. $Co\ (s) + Br_2\ (l) \rightarrow$
b. $MgBr_2\ (aq) + K_3PO_4\ (aq) \rightarrow$
c. $Mg\ (s) + O_2\ (g) \rightarrow$
d. $LiOH\ (aq) + H_3PO_4\ (aq) \rightarrow$