



# Chemical Reactions

**BIG Idea** Millions of chemical reactions in and around you transform reactants into products, resulting in the absorption or release of energy.

## 9.1 Reactions and Equations

**MAIN Idea** Chemical reactions are represented by balanced chemical equations.

## 9.2 Classifying Chemical Reactions

**MAIN Idea** There are four types of chemical reactions: synthesis, combustion, decomposition, and replacement reactions.

## 9.3 Reactions in Aqueous Solutions

**MAIN Idea** Double-replacement reactions occur between substances in aqueous solutions and produce precipitates, water, or gases.

## ChemFacts

- Wood has to be heated to  $260^{\circ}\text{C}$  before it bursts into flames.
- Before wood burns, the water in it boils off. This produces sizzling sounds.
- The smoke produced when wood burns contains more than 100 substances.



Before fire

After fire

# Start-Up Activities

## LAUNCH Lab

### How do you know when a chemical change has occurred?

An indicator is a chemical that is added to the substances in a chemical reaction to show when change occurs.



#### Procedure



1. Read and complete the lab safety form.
2. Measure 10.0 mL of **distilled water** in a **25-mL graduated cylinder**, and pour it into a **100-mL beaker**. Using a **pipette**, add one drop of **0.1M ammonia** to the water.  
**WARNING: Ammonia vapors are extremely irritating.**
3. Stir 15 drops of **universal indicator** into the solution with a **stirring rod**. Observe the solution's color. Measure its temperature with a **thermometer**.
4. Drop an **effervescent tablet** into the solution. Observe what happens. Record your observations, including any temperature change.

#### Analysis

1. **Describe** any changes in the color or temperature of the solution.
2. **Explain** Was a gas produced? If so, what did you observe to support this conclusion?
3. **Analyze** Did a physical change or a chemical change occur? Explain.

**Inquiry** What does the universal indicator tell you about the solution? Design an experiment to support your prediction.

### FOLDABLES™ Study Organizer

**Chemical Reactions** Make the following Foldable to help you organize information about how chemical reactions are classified.

- ▶ **STEP 1** Fold a sheet of paper lengthwise, keeping the margin visible on the left side.



- ▶ **STEP 2** Cut the top flap into five tabs.



- ▶ **STEP 3** Label as follows: *Chemical Reactions*, *Synthesis*, *Combustion*, *Decomposition*, *Single-Replacement*, and *Double-Replacement*.



- FOLDABLES** Use this Foldable with Section 9.2. As you read the section, summarize each type of chemical reaction and provide examples.

### Chemistry Online

Visit [glencoe.com](http://glencoe.com) to:

- ▶ study the entire chapter online
- ▶ explore **Concepts in Motion**
- ▶ take Self-Check Quizzes
- ▶ use the Personal Tutor to work Example Problems step-by-step
- ▶ access Web Links for more information, projects, and activities
- ▶ find the Try at Home Lab, Preventing a Chemical Reaction

## Section 9.1

### Objectives

- ▶ **Recognize** evidence of chemical change.
- ▶ **Represent** chemical reactions with equations.
- ▶ **Balance** chemical equations.

### Review Vocabulary

**chemical change:** a process involving one or more substances changing into a new substance

### New Vocabulary

chemical reaction  
reactant  
product  
chemical equation  
coefficient

# Reactions and Equations

**MAIN Idea** Chemical reactions are represented by balanced chemical equations.

**Real-World Reading Link** When you purchase bananas from a grocery store, they might be green. Within a few days, the bananas turn yellow. This color change is one of the ways you can tell a chemical reaction occurs.

## Chemical Reactions

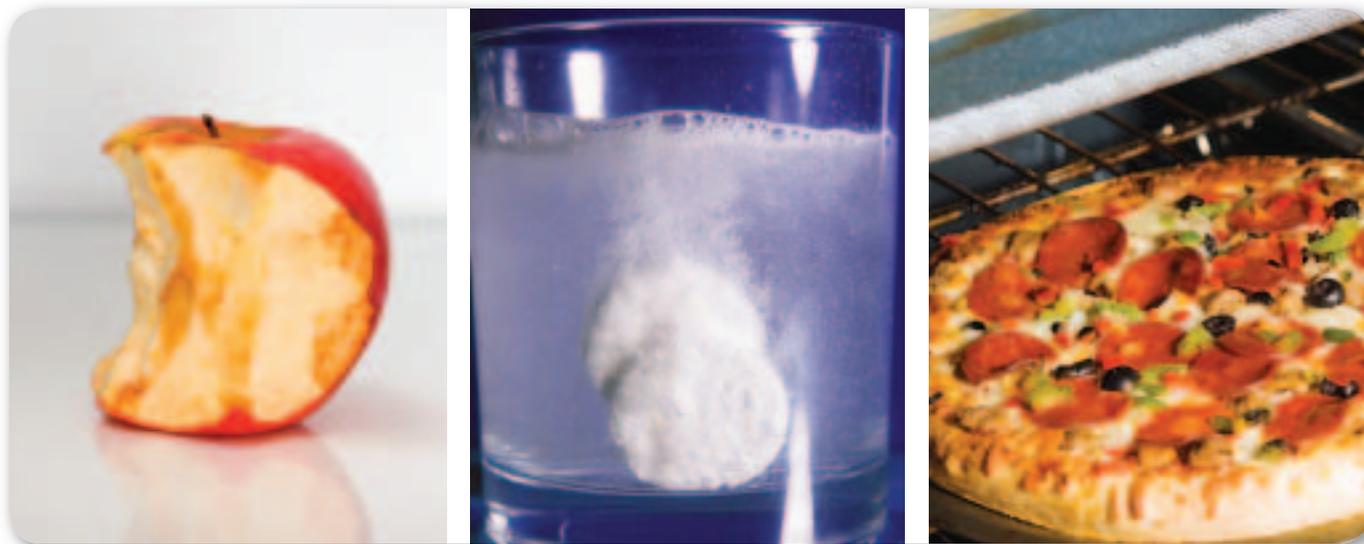
Do you know that the foods you eat, the fibers in your clothes, and the plastic in your CDs have something in common? Foods, fibers, and plastics are produced when the atoms in substances are rearranged to form different substances. Atoms are rearranged during the forest fire shown in the photo at the beginning of the chapter. They were also rearranged when you dropped the effervescent tablet into the beaker of water and indicator during the Launch Lab.

The process by which the atoms of one or more substances are rearranged to form different substances is called a **chemical reaction**. A chemical reaction is another name for a chemical change, which you read about in Chapter 3. Chemical reactions affect every part of your life. They break down your food, producing the energy you need to live. Chemical reactions in the engines of cars and buses provide the energy to power the vehicles. They produce natural fibers, such as cotton and wool, in plants and animals. In factories, they produce synthetic fibers such as nylon, shown in **Figure 9.1**.

**Evidence of a chemical reaction** How can you tell when a chemical reaction has taken place? Although some chemical reactions are hard to detect, many reactions provide physical evidence that they have occurred. A temperature change can indicate a chemical reaction. Many reactions, such as those that occur during the burning of wood, release energy in the form of heat and light. Other chemical reactions absorb heat.

■ **Figure 9.1** When adipoyl chloride in dichloromethane reacts with hexanediamine, nylon is formed. Nylon is used in many products, including carpeting, clothing, sports equipment, and tires.

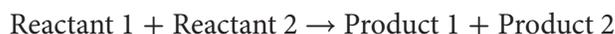




In addition to a temperature change, other types of evidence might indicate that a chemical reaction has occurred. One indication of a chemical reaction is a color change. For example, you might have noticed that the color of some nails that are left outside changes from silver to orange-brown in a short time. The color change is evidence that a chemical reaction occurred between the iron in the nail and the oxygen in air. A banana changing from green to yellow is another example of a color change indicating that a chemical reaction has occurred. Odor, gas bubbles, and the appearance of a solid are also indications of chemical change. Each of the photographs in **Figure 9.2** shows evidence of a chemical reaction.

## Representing Chemical Reactions

Chemists use statements called equations to represent chemical reactions. Equations show a reaction's **reactants**, which are the starting substances, and **products**, which are the substances formed during the reaction. Chemical equations do not express numerical equalities as mathematical equations do because during chemical reactions the reactants are used up as the products form. Instead, the equations used by chemists show the direction in which the reaction progresses. Therefore, an arrow rather than an equal sign is used to separate the reactants from the products. You read the arrow as *react to produce* or *yield*. The reactants are written to the left of the arrow, and the products are written to right of the arrow. When there are two or more reactants, or when there are two or more products, a plus sign separates each reactant or each product. These elements of equation notation are shown below.



In equations, symbols are used to show the physical states of the reactants and products. Reactants and products can exist as solids, liquids, and gases. When they are dissolved in water, they are said to be aqueous. It is important to show the physical states of a reaction's reactants and products in an equation because the physical states provide clues about how the reaction occurs. Some basic symbols used in equations are shown in **Table 9.1**.

■ **Figure 9.2** Each of these photos illustrates evidence of a chemical reaction.

**Describe** the evidence in each photo that tells you a chemical reaction has occurred.

<b>Table 9.1</b>		<b>Symbols Used in Equations</b>
Symbol	Purpose	
+	separates two or more reactants or products	
→	separates reactants from products	
⇌	separates reactants from products and indicates a reversible reaction	
(s)	identifies a solid state	
(l)	identifies a liquid state	
(g)	identifies a gaseous state	
(aq)	identifies a water solution	



■ **Figure 9.3** Science, like all other disciplines, has a specialized language that allows specific information to be communicated in a uniform manner. This reaction between aluminum and bromine can be described by a word equation, a skeleton equation, or a balanced chemical equation.

## VOCABULARY

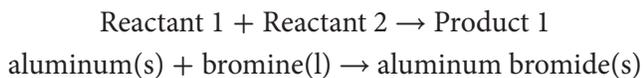
### ACADEMIC VOCABULARY

#### Formula

an expression using chemical symbols to represent a chemical reaction

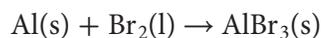
*The chemical formula for water is  $H_2O$ .*

**Word equations** You can use statements called word equations to indicate the reactants and products of chemical reactions. The word equation below describes the reaction between aluminum (Al) and bromine (Br), which is shown in **Figure 9.3**. Aluminum is a solid, and bromine is a liquid. The brownish-red cloud in the photograph is excess bromine. The reaction's product, which is solid particles of aluminum bromide ( $AlBr_3$ ), settles on the bottom of the beaker.



This word equation reads, "Aluminum and bromine react to produce aluminum bromide."

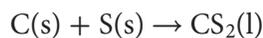
**Skeleton equations** Although word equations help to describe chemical reactions, they lack important information. A skeleton equation uses chemical formulas rather than words to identify the reactants and the products. For example, the skeleton equation for the reaction between aluminum and bromine uses the formulas for aluminum, bromine, and aluminum bromide in place of words.



How would you write the skeleton equation that describes the reaction between carbon and sulfur to form carbon disulfide? Carbon and sulfur are solids. First, write the chemical formulas for the reactants to the left of the arrow. Then, separate the reactants with a plus sign and indicate their physical states.



Finally, write the chemical formula for the product, liquid carbon disulfide, to the right of the arrow and indicate its physical state. The result is the skeleton equation for the reaction.



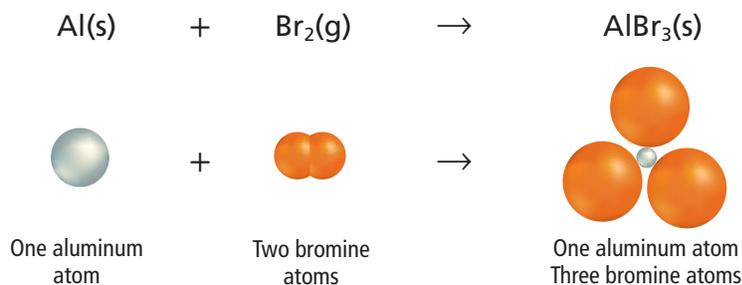
This skeleton equation tells us that carbon in the solid state reacts with sulfur in the solid state to produce carbon disulfide in the liquid state.

## PRACTICE Problems

Extra Practice Page 980 and [glencoe.com](http://glencoe.com)

**Write skeleton equations for the following word equations.**

1. Hydrogen and bromine gases react to yield hydrogen bromide.  
$$\text{hydrogen(g)} + \text{bromine(g)} \rightarrow \text{hydrogen bromide(g)}$$
2. When carbon monoxide and oxygen react, carbon dioxide forms.  
$$\text{carbon monoxide(g)} + \text{oxygen(g)} \rightarrow \text{carbon dioxide(g)}$$
3. **Challenge** Write the word equation and the skeleton equation for the following reaction: when heated, solid potassium chlorate yields solid potassium chloride and oxygen gas.



■ **Figure 9.4** The information conveyed by skeleton equations is limited. In this case, the skeleton equation is correct, but it does not show the exact number of atoms that interact. Refer to **Table R-1** on page 968 for a key to atom color conventions.

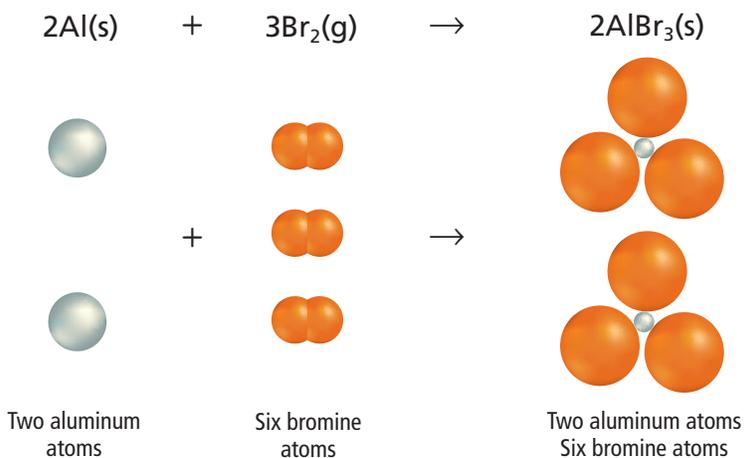
**Chemical equations** Like word equations, skeleton equations lack some information about reactions. Recall from Chapter 3 that the law of conservation of mass states that in a chemical change, matter is neither created nor destroyed. Chemical equations must show that matter is conserved during a reaction. Skeleton equations lack that information.

Look at **Figure 9.4**. The skeleton equation for the reaction between aluminum and bromine shows that one aluminum atom and two bromine atoms react to produce a substance containing one aluminum atom and three bromine atoms. Was a bromine atom created in the reaction? Atoms are not created in chemical reactions, and to accurately show what happened, more information is needed.

To accurately represent a chemical reaction by an equation, the equation must show equal numbers of atoms of each reactant and each product on both sides of the arrow. Such an equation is called a balanced chemical equation. A **chemical equation** is a statement that uses chemical formulas to show the identities and relative amounts of the substances involved in a chemical reaction.

## Balancing Chemical Equations

The balanced equation for the reaction between aluminum and bromine, shown in **Figure 9.5**, reflects the law of conservation of mass. To balance an equation, you must find the correct coefficients for the chemical formulas in the skeleton equation. A **coefficient** in a chemical equation is the number written in front of a reactant or product. Coefficients are usually whole numbers and are not usually written if the value is one. The coefficients in a balanced equation describe the lowest whole-number ratio of the amounts of all of the reactants and products.



■ **Figure 9.5** In a balanced chemical equation, the number of particles on the reactant side of the equation equals the number of particles on the product side of the equation. In this case, two aluminum atoms and six bromine atoms are needed on both sides of the equation.

**Chemistry**  **Online**

**Personal Tutor** For an online tutorial on balancing chemical equations, visit [glencoe.com](http://glencoe.com).

**Steps for balancing equations** Most chemical equations can be balanced by following the steps given in **Table 9.2**. For example, you can use these steps to write the chemical equation for the reaction between hydrogen ( $\text{H}_2$ ) and chlorine ( $\text{Cl}_2$ ) that produces hydrogen chloride ( $\text{HCl}$ ).

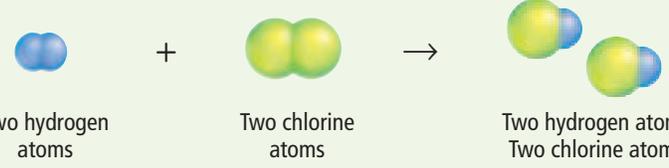
Concepts in Motion

Interactive Table Explore balancing chemical equations at [glencoe.com](http://glencoe.com).



**Table 9.2**

**Steps for Balancing Equations**

Step	Process	Example
1	Write the skeleton equation for the reaction. Make sure that the chemical formulas correctly represent the substances. An arrow separates the reactants from the products, and a plus sign separates multiple reactants and products. Show the physical states of all reactants and products.	$\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{HCl}(\text{g})$  <p>Two hydrogen atoms      Two chlorine atoms      One hydrogen atom One chlorine atom</p>
2	Count the atoms of the elements in the reactants. If a reaction involves identical polyatomic ions in the reactants and products, count each polyatomic ion as a single element. This reaction does not involve any polyatomic ions. Two atoms of hydrogen and two atoms of chlorine are reacting.	$\text{H}_2 + \text{Cl}_2 \rightarrow$ <p>2 atoms H      2 atoms Cl</p>
3	Count the atoms of the elements in the products. One atom of hydrogen and one atom of chlorine are produced.	<p>HCl 1 atom H + 1 atom Cl</p>
4	Change the coefficients to make the number of atoms of each element equal on both sides of the equation. Never change a subscript in a chemical formula to balance an equation because doing so changes the identity of the substance.	$2\text{H}_2 + 2\text{Cl}_2 \rightarrow 2\text{HCl}$  <p>Two hydrogen atoms      Two chlorine atoms      Two hydrogen atoms Two chlorine atoms</p>
5	Write the coefficients in their lowest possible ratio. The coefficients should be the smallest possible whole numbers. The ratio 1 hydrogen to 1 chlorine to 2 hydrogen chloride (1:1:2) is the lowest-possible ratio because the coefficients cannot be reduced further and still remain whole numbers.	$\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$ <p>1:1:2 1 <math>\text{H}_2</math> to 1 <math>\text{Cl}_2</math> to 2 HCl</p>
6	Check your work. Make sure that the chemical formulas are written correctly. Then, check that the number of atoms of each element is equal on both sides of the equation.	$\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ <p>2 atoms H      2 atoms Cl      2 atoms H + 2 atoms Cl There are two hydrogen atoms and two chlorine atoms on both sides of the equation.</p>

## EXAMPLE Problem 9.1

**Writing a Balanced Chemical Equation** Write the balanced chemical equation for the reaction in which aqueous sodium hydroxide and aqueous calcium bromide react to produce solid calcium hydroxide and aqueous sodium bromide.

### 1 Analyze the Problem

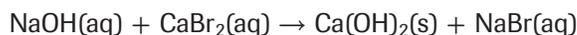
You are given the reactants and products in a chemical reaction. Start with a skeleton equation, and use the steps given in **Table 9.2** for balancing chemical equations.

#### Math Handbook

Ratios  
page 964

### 2 Solve for the Unknown

Write the skeleton equation for the chemical reaction. Be sure to put the reactants on the left side of the arrow and the products on the right. Separate the substances with plus signs, and indicate their physical states.



1 Na, 1 O, 1 H, 1 Ca, 2 Br

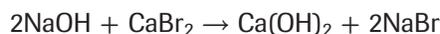
Count the atoms of each element in the reactants.

1 Na, 2 O, 2 H, 1 Ca, 1 Br

Count the atoms of each element in the products.



Insert the coefficient 2 in front of NaOH to balance the hydroxide ions.



Insert the coefficient 2 in front of NaBr to balance the Na and Br atoms.

The ratio of the coefficients is 2:1:1:2.

Write the coefficients in their lowest-possible ratio.

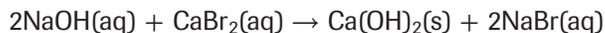
Reactants: 2 Na, 2 OH, 1 Ca, 2 Br

Products: 2 Na, 2 OH, 1 Ca, 2 Br

Check to make sure that the number of atoms of each element is equal on both sides of the equation.

### 3 Evaluate the Answer

The chemical formulas for all substances are written correctly. The number of atoms of each element is equal on both sides of the equation. The coefficients are written in the lowest possible ratio. The balanced chemical equation for the reaction is



## Real-World Chemistry Calcium Hydroxide



**Reef aquariums** An aqueous solution of calcium hydroxide is used in reef aquariums to provide calcium for animals such as snails and corals. Calcium hydroxide reacts with the carbon dioxide in the water to produce calcium and bicarbonate ions. Reef animals use the calcium to grow shells and strong skeletal systems.

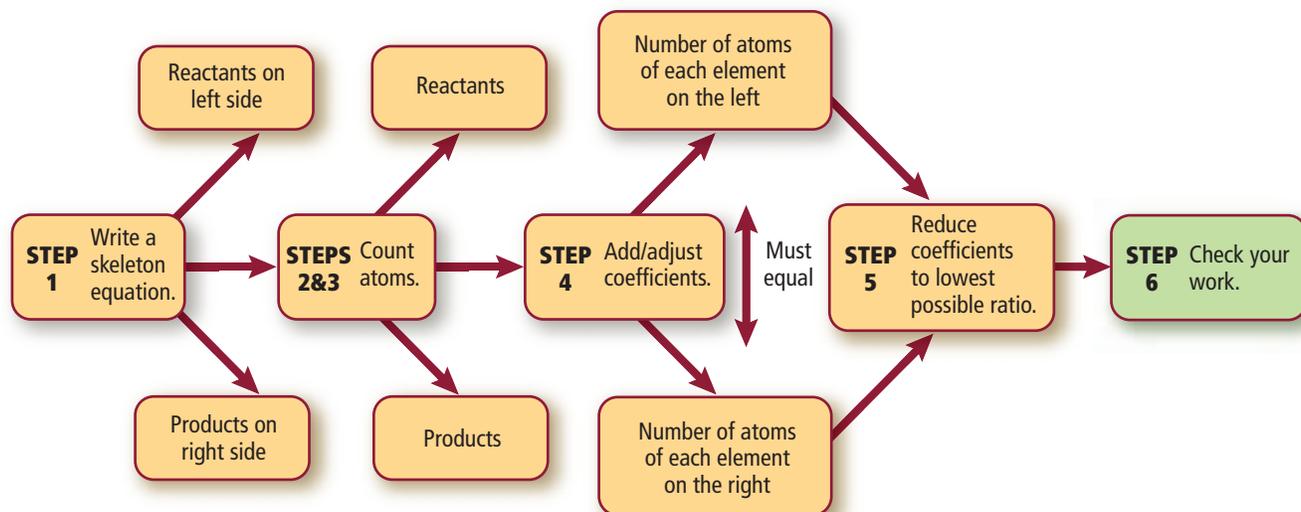
## PRACTICE Problems

Extra Practice Page 980 and [glencoe.com](http://glencoe.com)

Write chemical equations for each of the following reactions.

- In water, iron(III) chloride reacts with sodium hydroxide, producing solid iron(III) hydroxide and sodium chloride.
- Liquid carbon disulfide reacts with oxygen gas, producing carbon dioxide gas and sulfur dioxide gas.
- Challenge** A piece of zinc metal is added to a solution of hydrogen sulfate. This reaction produces a gas and a solution of zinc sulfate.

## Balancing Chemical Equations



■ **Figure 9.6** It is imperative to your study of chemistry to be able to balance chemical equations. Use this flowchart to help you master the skill. Notice that the numbered steps correspond to the steps in **Table 9.2**.

**Obeying the law of conservation of mass** Probably the most fundamental concept of chemistry is the law of conservation of mass that you first encountered in Chapter 3. All chemical reactions obey the law that matter is neither created nor destroyed. Therefore, it is also fundamental that the equations that represent chemical reactions include sufficient information to show that the reaction obeys the law of conservation of mass.

You have learned how to show this relationship with balanced chemical equations. The flowchart shown in **Figure 9.6** summarizes the steps for balancing equations. You will probably find that some chemical equations can be balanced easily, whereas others are more difficult to balance. All chemical equations, however, can be balanced by the process you learned in this section.

## Section 9.1 Assessment

### Section Summary

- ▶ Some physical changes are evidence that indicate a chemical reaction has occurred.
- ▶ Word equations and skeleton equations provide important information about a chemical reaction.
- ▶ A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction.
- ▶ Balancing an equation involves adjusting the coefficients until the number of atoms of each element is equal on both sides of the equation.

7. **MAIN Idea** Explain why it is important that a chemical equation be balanced.
8. **List** three types of physical evidence that indicate a chemical reaction has occurred.
9. **Compare and contrast** a skeleton equation and a chemical equation.
10. **Explain** why it is important to reduce coefficients in a balanced equation to the lowest-possible whole-number ratio.
11. **Analyze** When balancing a chemical equation, can you adjust the subscript in a formula? Explain.
12. **Assess** Is the following equation balanced? If not, correct the coefficients to balance the equation.  

$$2\text{K}_2\text{CrO}_4(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{PbCrO}_4(\text{s})$$
13. **Evaluate** Aqueous phosphoric acid and aqueous calcium hydroxide react to form solid calcium phosphate and water. Write a balanced chemical equation for this reaction.

## Section 9.2

### Objectives

- **Classify** chemical reactions.
- **Identify** the characteristics of different classes of chemical reactions.

### Review Vocabulary

**metal:** an element that is a solid at room temperature, a good conductor of heat and electricity, and is generally shiny

### New Vocabulary

synthesis reaction  
combustion reaction  
decomposition reaction  
single-replacement reaction  
double-replacement reaction  
precipitate

# Classifying Chemical Reactions

**MAIN Idea** There are four types of chemical reactions: synthesis, combustion, decomposition, and replacement reactions.

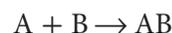
**Real-World Reading Link** It could take you a long time to find a specific novel in an unorganized bookstore. Bookstores classify and organize books into different categories to make your search easier. Chemical reactions are also classified and organized into different categories.

## Types of Chemical Reactions

Chemists classify chemical reactions in order to organize the many reactions that occur daily. Knowing the categories of chemical reactions can help you remember and understand them. It can also help you recognize patterns and predict the products of many chemical reactions. One way chemists classify reactions is to distinguish among the four types: synthesis, combustion, decomposition, and replacement reactions. Some reactions fit into more than one of these types.

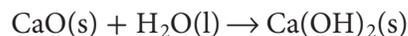
## Synthesis Reactions

In **Figure 9.7**, sodium and chlorine react to produce sodium chloride. This reaction is a **synthesis reaction**—a chemical reaction in which two or more substances (A and B) react to produce a single product (AB).

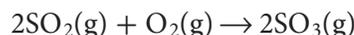


When two elements react, the reaction is always a synthesis reaction.

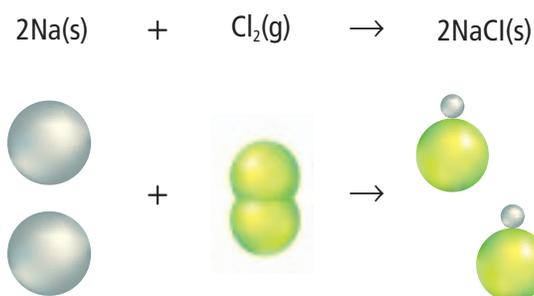
Two compounds can also combine to form one compound. For example, the reaction between calcium oxide (CaO) and water (H<sub>2</sub>O) to form calcium hydroxide (Ca(OH)<sub>2</sub>) is a synthesis reaction.



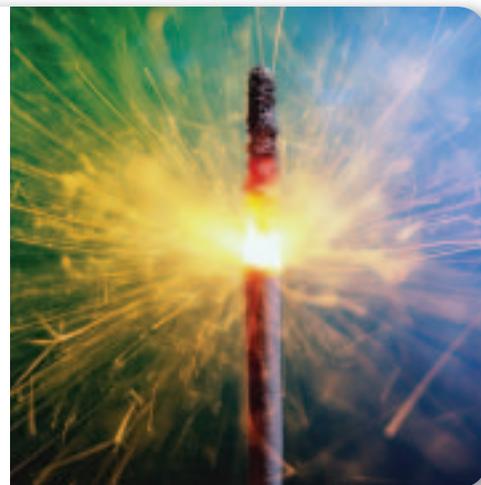
Another type of synthesis reaction involves a reaction between a compound and an element, as happens when sulfur dioxide gas (SO<sub>2</sub>) reacts with oxygen gas (O<sub>2</sub>) to form sulfur trioxide (SO<sub>3</sub>).



■ **Figure 9.7** In this synthesis reaction, two elements, sodium and chlorine, react to produce one compound, sodium chloride.



■ **Figure 9.8** The light produced by a sparkler is the result of a combustion reaction between oxygen and different metals.



## VOCABULARY

### WORD ORIGIN

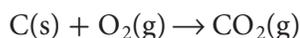
#### Combustion

comes from the Latin word *comburare*, meaning *to burn*

## Combustion Reactions

The synthesis reaction between sulfur dioxide and oxygen can also be classified as a combustion reaction. In a **combustion reaction**, such as the one shown in **Figure 9.8**, oxygen combines with a substance and releases energy in the form of heat and light. Oxygen can combine in this way with many different substances, making combustion reactions common. To learn more about the discovery of the chemical reaction for combustion and other reactions, review **Figure 9.9**.

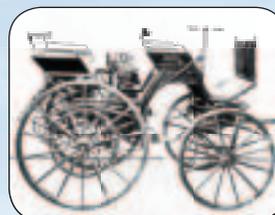
A combustion reaction occurs between hydrogen and oxygen when hydrogen is heated, as illustrated in **Figure 9.10**. Water is formed during the reaction, and a large amount of energy is released. Another important combustion reaction occurs when coal is burned to produce energy. Coal is called a fossil fuel because it contains the remains of plants that lived long ago. It is composed primarily of the element carbon. Coal-burning power plants generate electric power in many parts of the United States. The primary reaction that occurs in these plants is between carbon and oxygen.



## ■ Figure 9.9 Real-World Chemical Reactions

Throughout history, people have worked to understand and apply the power of chemical reactions to solve problems.

**CIRCA 1800**  
Experiments with plants result in the discovery of the balanced chemical equation for photosynthesis.



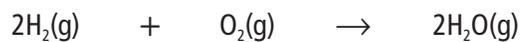
**1885** The internal combustion engine is invented. It later becomes the prototype for the modern gas engine.

**1775** Antoine Lavoisier demonstrates that combustion is an exothermic chemical reaction involving oxygen.



**1635** America's first chemical plant opens in Boston. Products include saltpeter, a component of gunpowder, and alum, a chemical used in tanning animal skins.

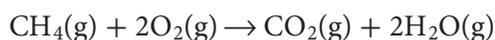
**1909–1910** German chemists Fritz Haber and Carl Bosch develop the Haber-Bosch process for synthesizing ammonia.



■ **Figure 9.10** During a combustion reaction between oxygen and hydrogen, water is formed.

**Analyze** Why is this chemical reaction both a synthesis reaction and a combustion reaction?

Note that the combustion reactions just mentioned are also synthesis reactions. However, not all combustion reactions are synthesis reactions. For example, the reaction involving methane gas ( $\text{CH}_4$ ) and oxygen illustrates a combustion reaction in which one substance replaces another in the formation of products.



Methane, which belongs to a group of substances called hydrocarbons, is the major component of natural gas. All hydrocarbons contain carbon and hydrogen and burn in oxygen to yield carbon dioxide and water. You will learn more about hydrocarbons in Chapter 21.

## PRACTICE Problems

Extra Practice Page 980 and [glencoe.com](http://glencoe.com)

Write chemical equations for the following reactions. Classify each reaction into as many categories as possible.

- The solids aluminum and sulfur react to produce aluminum sulfide.
- Water and dinitrogen pentoxide gas react to produce aqueous hydrogen nitrate.
- The gases nitrogen dioxide and oxygen react to produce dinitrogen pentoxide gas.
- Challenge** Sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and sodium hydroxide solutions react to produce aqueous sodium sulfate and water.

1950



1952 A heavy smog—sulfur dioxide and other coal-burning products—settles over London for five days in December, causing 4000 deaths.

1974–1978 Researchers demonstrate that chlorofluorocarbons (CFCs) can deplete the ozone layer. The use of CFCs as spray propellants is banned in the United States.

1965



2004 Scientists discover that migrating birds are guided by chemical reactions in their bodies that are influenced by Earth's magnetic field.

1980

1995 Researchers use the atomic force microscope to create and observe chemical reactions as they occur molecule by molecule, paving the way for nanoscale engineering.

1995

2010

Concepts in Motion

**Interactive Time Line** To learn more about these discoveries and others, visit [glencoe.com](http://glencoe.com).





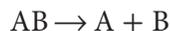
■ **Figure 9.11** The decomposition of sodium azide, which produces a gas, is the chemical reaction that inflates air bags.

**FOLDABLES**

Incorporate information from this section into your Foldable.

## Decomposition Reactions

Some chemical reactions are essentially the opposite of synthesis reactions. These reactions are classified as decomposition reactions. A **decomposition reaction** is one in which a single compound breaks down into two or more elements or new compounds. In generic terms, decomposition reactions can be represented as follows.



Decomposition reactions often require an energy source, such as heat, light, or electricity, to occur. For example, ammonium nitrate breaks down into dinitrogen monoxide and water when the reactant is heated to a high temperature.



Notice that this decomposition reaction involves one reactant breaking down into more than one product.

The outcome of another decomposition reaction is shown in **Figure 9.11**. Automobile safety air bags inflate rapidly as sodium azide pellets decompose. A device that can provide an electric signal to start the reaction is packaged inside air bags along with the sodium azide pellets. When the device is activated, sodium azide decomposes, producing nitrogen gas that quickly inflates the air bag.

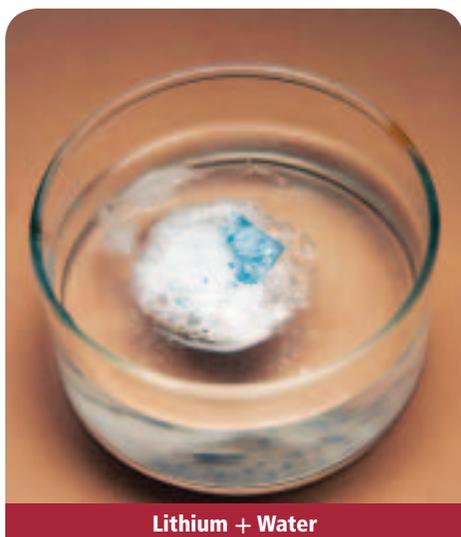


### PRACTICE Problems

Extra Practice Page 980 and [glencoe.com](http://glencoe.com)

Write chemical equations for the following decomposition reactions.

18. Aluminum oxide(s) decomposes when electricity passes through it.
19. Nickel(II) hydroxide(s) decomposes to produce nickel(II) oxide(s) and water.
20. **Challenge** Heating sodium hydrogen carbonate(s) produces sodium carbonate(aq) and water. Carbon dioxide gas is also produced.



Lithium + Water



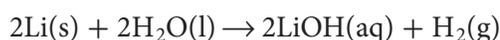
Copper + Silver Nitrate

■ **Figure 9.12** In a single-replacement reaction, the atoms of one element replace the atoms of another element in a compound.

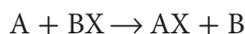
## Replacement Reactions

In contrast to synthesis, combustion, and decomposition reactions, many chemical reactions are replacement reactions and involve the replacement of an element in a compound. These reactions are also called displacement reactions. There are two types of replacement reactions: single-replacement reactions and double-replacement reactions.

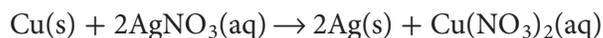
**Single-replacement reactions** The reaction between lithium and water is shown in **Figure 9.12**. The following chemical equation shows that a lithium atom replaces one of the hydrogen atoms in a water molecule.



A reaction in which the atoms of one element replace the atoms of another element in a compound is called a **single-replacement reaction**.



**Metal replaces hydrogen or another metal** The reaction between lithium and water is one type of single-replacement reaction, in which a metal replaces a hydrogen atom in a water molecule. Another type of single-replacement reaction occurs when one metal replaces another metal in a compound dissolved in water. **Figure 9.12** shows a single-replacement reaction occurring when a bar of pure copper is placed in aqueous silver nitrate. The crystals that are accumulating on the copper bar are the silver atoms that the copper atoms replaced.



A metal will not always replace another metal in a compound dissolved in water because metals differ in their reactivities. Reactivity is the ability to react with another substance. An activity series of some metals is shown in **Figure 9.13**. This series orders metals by reactivity with other metals. Single-replacement reactions are used to determine a metal's position on the list. The most active metals are at the top of the list. The least active metals are at the bottom. Similarly, the reactivity of each halogen has been determined and listed, as shown in **Figure 9.13**.

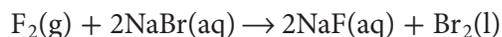
■ **Figure 9.13** An activity series, similar to the series shown here for various metals and halogens, is a useful tool for determining whether a chemical reaction will occur and for determining the result of a single-replacement reaction.

<b>Most active</b>  <b>Least active</b>	<b>METALS</b>
	Lithium
	Rubidium
	Potassium
	Calcium
	Sodium
	Magnesium
	Aluminum
	Manganese
	Zinc
	Iron
	Nickel
	Tin
	Lead
Copper	
Silver	
Platinum	
Gold	
<b>Most active</b>  <b>Least active</b>	<b>HALOGENS</b>
	Fluorine
	Chlorine
	Bromine
Iodine	

You can use the activity series to predict whether or not certain reactions will occur. A specific metal can replace any metal listed below it that is in a compound. It cannot replace any metal listed above it. For example, copper atoms replace silver atoms in a solution of silver nitrate. However, if you place a silver wire in aqueous copper(II) nitrate, the silver atoms will not replace the copper. Silver is listed below copper in the activity series, so no reaction occurs. The letters NR (no reaction) are commonly used to indicate that a reaction will not occur.



**Nonmetal replaces nonmetal** A third type of single-replacement reaction involves the replacement of a nonmetal in a compound by another nonmetal. Halogens are frequently involved in these types of reactions. Like metals, halogens exhibit different activity levels in single-replacement reactions. The reactivities of halogens, determined by single-replacement reactions, are also shown in **Figure 9.13**. The most active halogen is fluorine, and the least active is iodine. A more reactive halogen replaces a less reactive halogen that is part of a compound dissolved in water. For example, fluorine replaces bromine in water containing dissolved sodium bromide. However, bromine does not replace fluorine in water containing dissolved sodium fluoride.



**Reading Check** Explain how a single-replacement reaction works.

## PROBLEM-SOLVING LAB

### Analyze Trends

**How can you explain the reactivities of halogens?** The location of all the halogens in group 17 in the periodic table tells you that halogens have common characteristics. Indeed, halogens are all nonmetals and have seven electrons in their outermost orbitals. However, each halogen also has its own characteristics, such as the ability to react with other substances.

#### Analysis

Examine the accompanying data table. It includes data about the atomic radii, ionization energies, and electronegativities of the halogens.

#### Think Critically

- 1. Make graphs** Use the information in the data table to make three line graphs.
- 2. Describe** any periodic trends that you identify in the data.
- 3. Relate** any periodic trends that you identify among the halogens to the activity series of halogens shown in **Figure 9.13**.
- 4. Predict** the location of the element astatine in the activity series of halogens. Explain.

**Properties of Halogens**

Halogen	Atomic Radius (pm)	Ionization Energy (kJ/mol)	Electronegativity
Fluorine	72	1681	3.98
Chlorine	100	1251	3.16
Bromine	114	1140	2.96
Iodine	133	1008	2.66
Astatine	140	920	2.2

## EXAMPLE Problem 9.2

**Single-Replacement Reactions** Predict the products that will result when these reactants combine, and write a balanced chemical equation for each reaction.

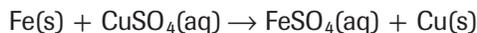
- $\text{Fe(s)} + \text{CuSO}_4\text{(aq)} \rightarrow$
- $\text{Br}_2\text{(l)} + \text{MgCl}_2\text{(aq)} \rightarrow$
- $\text{Mg(s)} + \text{AlCl}_3\text{(aq)} \rightarrow$

### 1 Analyze the Problem

You are given three sets of reactants. Using **Figure 9.13**, you must first determine if each reaction occurs. Then, if a reaction is predicted, you can determine the product(s) of the reaction. With this information you can write a skeleton equation for the reaction. Finally, you can use the steps for balancing chemical equations to write the complete balanced chemical equation.

### 2 Solve for the Unknown

- Iron is listed above copper in the activity series. Therefore, the first reaction will occur because iron is more reactive than copper. In this case, iron will replace copper. The skeleton equation for this reaction is



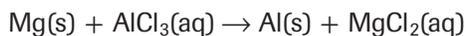
This equation is balanced.

- In the second reaction, chlorine is more reactive than bromine because bromine is listed below chlorine in the activity series. Therefore, the reaction will not occur. The skeleton equation for this situation is

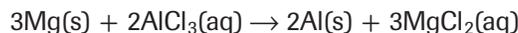


No balancing is required.

- Magnesium is listed above aluminum in the activity series. Therefore, the third reaction will occur because magnesium is more reactive than aluminum. In this case, magnesium will replace aluminum. The skeleton equation for this reaction is



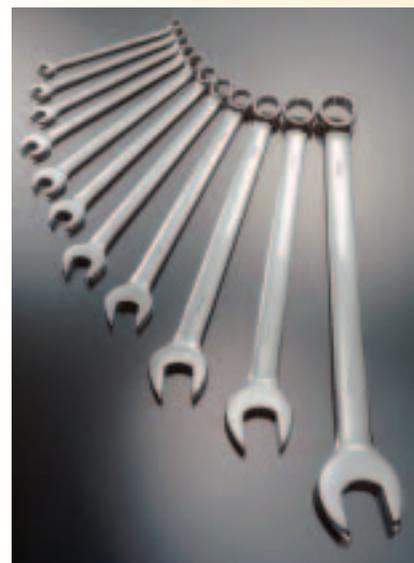
This equation is not balanced. The balanced equation is



### 3 Evaluate the Answer

The activity series shown in **Figure 9.13** supports the reaction predictions. The chemical equations are balanced because the number of atoms of each substance is equal on both sides of the equation.

## Real-World Chemistry Single-Replacement Reactions



**Zinc plating** Tools made of steel are often covered with a layer of zinc to prevent corrosion. Zinc is more reactive than the lead in steel. During zinc plating, the zinc replaces some of the surface lead, coating the steel.

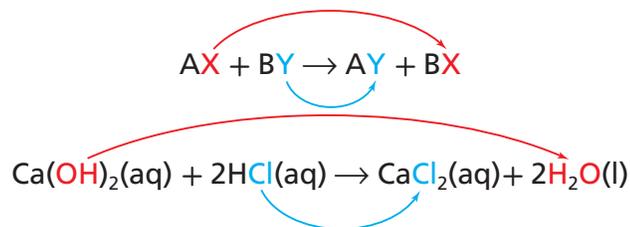
## PRACTICE Problems

Extra Practice Pages 980–981 and [glencoe.com](http://glencoe.com)

Predict whether the following single-replacement reactions will occur. If a reaction occurs, write a balanced equation for the reaction.

- $\text{K(s)} + \text{ZnCl}_2\text{(aq)} \rightarrow$
- $\text{Cl}_2\text{(g)} + \text{HF(aq)} \rightarrow$
- $\text{Fe(s)} + \text{Na}_3\text{PO}_4\text{(aq)} \rightarrow$
- Challenge**  $\text{Al(s)} + \text{Pb(NO}_3)_2\text{(aq)} \rightarrow$

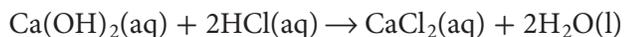
■ **Figure 9.14** The color-coding in the generic equation for a double-replacement reaction and in the equation for the reaction between calcium hydroxide and hydrochloric acid shows the anions changing places.



**Double-replacement reactions** The final type of replacement reaction, which involves an exchange of ions between two compounds, is called a **double-replacement reaction**.

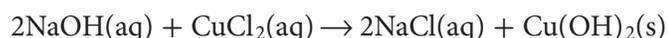
In the generic equation in **Figure 9.14**, A and B represent positively charged ions (cations), and X and Y represent negatively charged ions (anions). Notice that the anions have switched places and are now bonded to the other cations in the reaction. In other words, X replaces Y and Y replaces X—a double replacement. More simply, the positive and negative ions of two compounds switch places.

The reaction between calcium hydroxide and hydrochloric acid is a double-replacement reaction.



The ionic components of the reaction are  $\text{Ca}^{2+}$ ,  $\text{OH}^-$ ,  $\text{H}^+$ , and  $\text{Cl}^-$ . Knowing this, you can now see the two replacements of the reaction. The anions ( $\text{OH}^-$  and  $\text{Cl}^-$ ) have changed places and are now bonded to the other cations ( $\text{Ca}^{2+}$  and  $\text{H}^+$ ), as shown in **Figure 9.14**.

The reaction between sodium hydroxide and copper(II) chloride in solution is also a double-replacement reaction.



In this case, the anions ( $\text{OH}^-$  and  $\text{Cl}^-$ ) changed places and bonded to the other cations ( $\text{Na}^+$  and  $\text{Cu}^{2+}$ ). **Figure 9.15** shows that the result of this reaction is a solid product, copper(II) hydroxide. A solid produced during a chemical reaction in a solution is called a **precipitate**.

■ **Figure 9.15** When aqueous sodium hydroxide is added to a solution of copper(II) chloride, the anions ( $\text{OH}^-$  and  $\text{Cl}^-$ ) change places. The resulting products are sodium chloride, which remains in solution, and copper(II) hydroxide, the blue solid in the test tube.



Concepts in Motion

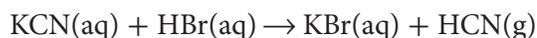
**Interactive Figure** To see an animation of a precipitate forming, visit [glencoe.com](http://glencoe.com).



## Table 9.3 Guidelines for Writing Double-Replacement Reactions

Step	Example
1. Write the components of the reactants in a skeleton equation.	$\text{Al}(\text{NO}_3)_3 + \text{H}_2\text{SO}_4$
2. Identify the cations and the anions in each compound.	$\text{Al}(\text{NO}_3)_3$ has $\text{Al}^{3+}$ and $\text{NO}_3^-$ $\text{H}_2\text{SO}_4$ has $\text{H}^+$ and $\text{SO}_4^{2-}$
3. Pair up each cation with the anion from the other compound.	$\text{Al}^{3+}$ pairs with $\text{SO}_4^{2-}$ $\text{H}^+$ pairs with $\text{NO}_3^-$
4. Write the formulas for the products using the pairs from Step 3.	$\text{Al}_2(\text{SO}_4)_3$ $\text{HNO}_3$
5. Write the complete equation for the double-replacement reaction.	$\text{Al}(\text{NO}_3)_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + \text{HNO}_3$
6. Balance the equation.	$2\text{Al}(\text{NO}_3)_3 + 3\text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 6\text{HNO}_3$

**Products of double-replacement reactions** One of the key characteristics of double-replacement reactions is the type of product that is formed when the reaction takes place. All double-replacement reactions produce either water, a precipitate, or a gas. Refer back to the two double-replacement reactions previously discussed in this section. The reaction between calcium hydroxide and hydrochloric acid produces water. A precipitate is produced in the reaction between sodium hydroxide and copper(II) chloride. An example of a double-replacement reaction that forms a gas is that of potassium cyanide and hydrobromic acid.



It is important to be able to evaluate the chemistry of double-replacement reactions and predict the products of these reactions. The basic steps to write double-replacement reactions are given in **Table 9.3**.

 **Reading Check** Describe what happens to the anions in a double-replacement reaction.

### PRACTICE Problems

Extra Practice Page 981 and [glencoe.com](http://glencoe.com)

Write the balanced chemical equations for the following double-replacement reactions.

- The two substances at right react to produce solid silver iodide and aqueous lithium nitrate.
- Aqueous barium chloride and aqueous potassium carbonate react to produce solid barium carbonate and aqueous potassium chloride.
- Aqueous sodium oxalate and aqueous lead(II) nitrate react to produce solid lead(II) oxalate and aqueous sodium nitrate.
- Challenge** Acetic acid ( $\text{CH}_3\text{COOH}$ ) and potassium hydroxide react to produce potassium acetate and water.



$\text{LiI}(\text{aq})$

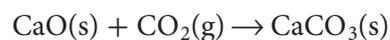


$\text{AgNO}_3(\text{aq})$

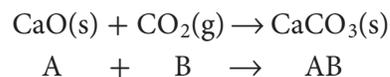

**Table 9.4**
**Predicting Products of Chemical Reactions**

Type of Reaction	Reactants	Probable Products	Generic Equation
Synthesis	<ul style="list-style-type: none"> <li>two or more substances</li> </ul>	<ul style="list-style-type: none"> <li>one compound</li> </ul>	$A + B \rightarrow AB$
Combustion	<ul style="list-style-type: none"> <li>a metal and oxygen</li> <li>a nonmetal and oxygen</li> <li>a compound and oxygen</li> </ul>	<ul style="list-style-type: none"> <li>the oxide of the metal</li> <li>the oxide of the nonmetal</li> <li>two or more oxides</li> </ul>	$A + O_2 \rightarrow AO$
Decomposition	<ul style="list-style-type: none"> <li>one compound</li> </ul>	<ul style="list-style-type: none"> <li>two or more elements and/or compounds</li> </ul>	$AB \rightarrow A + B$
Single-replacement	<ul style="list-style-type: none"> <li>a metal and a compound</li> <li>a nonmetal and a compound</li> </ul>	<ul style="list-style-type: none"> <li>a new compound and the replaced metal</li> <li>a new compound and the replaced nonmetal</li> </ul>	$A + BX \rightarrow AX + B$
Double-replacement	<ul style="list-style-type: none"> <li>two compounds</li> </ul>	<ul style="list-style-type: none"> <li>two different compounds, one of which is a solid, water, or a gas</li> </ul>	$AX + BY \rightarrow AY + BX$

**Table 9.4** summarizes the various types of chemical reactions. Use the table to help you organize the reactions, so that you can identify each and predict its products. For example, how would you determine what type of reaction occurs when solid calcium oxide and carbon dioxide gas react to produce solid calcium carbonate? First, write the chemical equation.



Second, determine what is happening in the reaction. In this case, two substances are reacting to form one compound. Third, use the table to identify the type of reaction. The reaction is a synthesis reaction. Fourth, check your answer by comparing the chemical equation to the generic equation for that type of reaction.

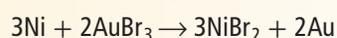


## Section 9.2 Assessment

### Section Summary

- Classifying chemical reactions makes them easier to understand, remember, and recognize.
- Activity series of metals and halogens can be used to predict if single-replacement reactions will occur.

- MAIN Idea** Describe the four types of chemical reactions and their characteristics.
- Explain** how an activity series of metals is organized.
- Compare and contrast** single-replacement reactions and double-replacement reactions.
- Describe** the result of a double-replacement reaction.
- Classify** What type of reaction is most likely to occur when barium reacts with fluorine? Write the chemical equation for the reaction.
- Interpret Data** Could the following reaction occur? Explain your answer.



## Section 9.3

### Objectives

- Describe aqueous solutions.
- Write complete ionic and net ionic equations for chemical reactions in aqueous solutions.
- Predict whether reactions in aqueous solutions will produce a precipitate, water, or a gas.

### Review Vocabulary

**solution:** a uniform mixture that might contain solids, liquids, or gases

### New Vocabulary

aqueous solution  
solute  
solvent  
complete ionic equation  
spectator ion  
net ionic equation

# Reactions in Aqueous Solutions

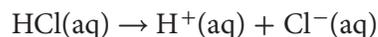
**MAIN Idea** Double-replacement reactions occur between substances in aqueous solutions and produce precipitates, water, or gases.

**Real-World Reading Link** One way to make lemonade involves using a powdered drink mix and water. When the powdered drink mix is added to the water, the lemonade crystals dissolve in the water, forming a solution. This solution is lemonade.

## Aqueous Solutions

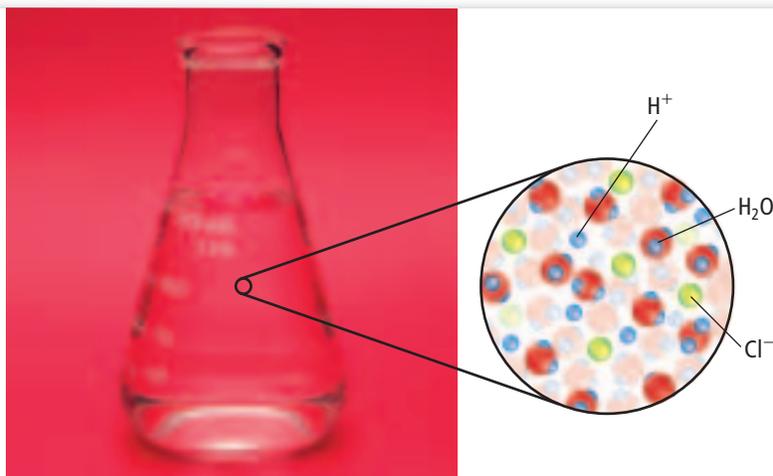
You read in Chapter 3 that a solution is a homogeneous mixture. Many of the reactions discussed in the previous section involve substances dissolved in water. When a substance dissolves in water, a solution forms. An **aqueous solution** contains one or more substances called **solutes** dissolved in the water. In this case, water is the **solvent**—the most plentiful substance in the solution.

**Molecular compounds in solution** Although water is always the solvent in aqueous solutions, there are many possible solutes. Some solutes, such as sucrose (table sugar) and ethanol (grain alcohol), are molecular compounds that exist as molecules in aqueous solutions. Other solutes are molecular compounds that form ions when they dissolve in water. For example, the molecular compound hydrogen chloride forms hydrogen ions and chloride ions when it dissolves in water, as shown in **Figure 9.16**. An equation can be used to show this ionization process.



Compounds such as hydrogen chloride that produce hydrogen ions in aqueous solution are acids. In fact, an aqueous solution of hydrogen chloride is often referred to as hydrochloric acid. You will read more about acids in Chapter 18.

■ **Figure 9.16** In water, hydrogen chloride (HCl) breaks apart into hydrogen ions ( $\text{H}^+$ ) and chloride ions ( $\text{Cl}^-$ ).



## VOCABULARY

### SCIENCE USAGE V. COMMON USAGE

#### Compound

**Science usage:** a chemical combination of two or more different elements

*Salt is a compound comprised of the elements sodium and chlorine.*

**Common usage:** a word that consists of two or more words

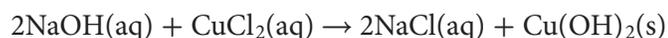
*Two compound words are basketball and textbook.*

**Ionic compounds in solution** In addition to molecular compounds, ionic compounds might be solutes in aqueous solutions. Recall from Chapter 7 that ionic compounds consist of positive ions and negative ions held together by ionic bonds. When ionic compounds dissolve in water, their ions can separate—a process called dissociation. For example, an aqueous solution of the ionic compound sodium hydroxide contains  $\text{Na}^+$  and  $\text{Cl}^-$  ions.

## Types of Reactions in Aqueous Solutions

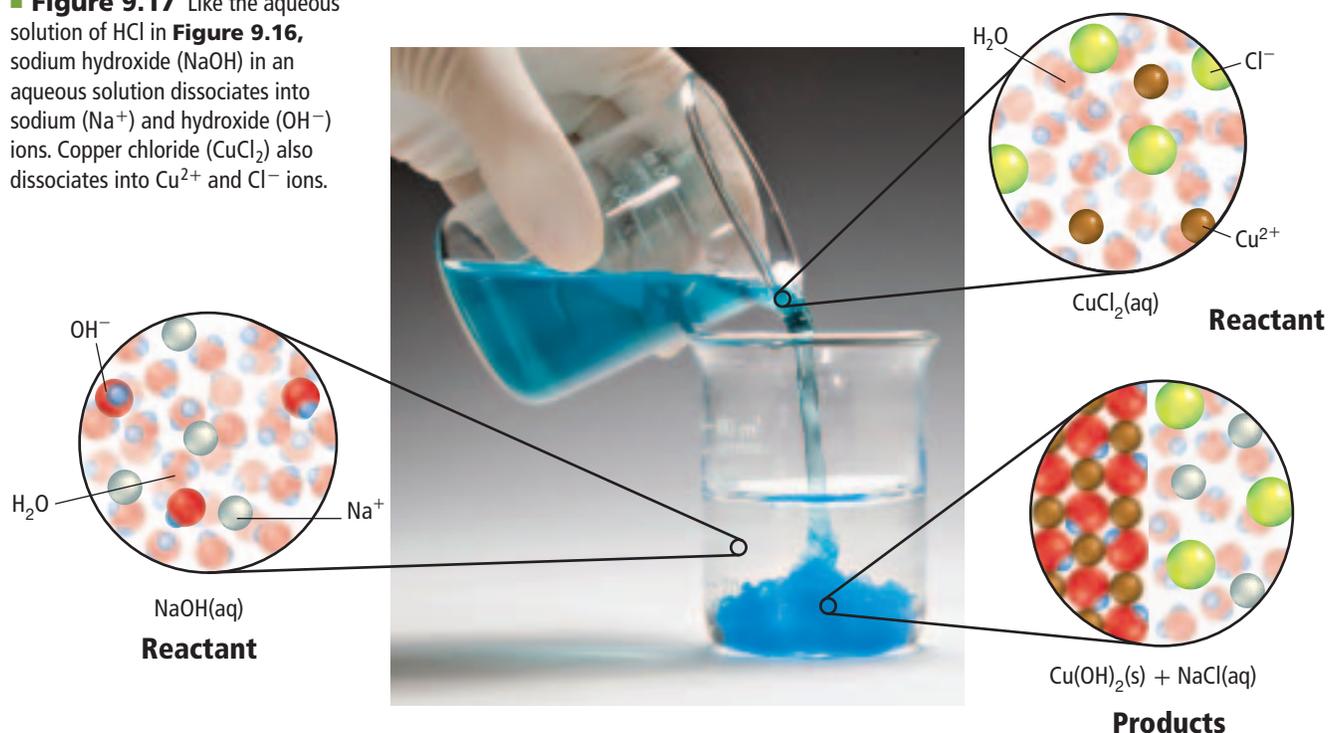
When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. These reactions are always double-replacement reactions. The solvent molecules, which are all water molecules, do not usually react. Three types of products can form from the double-replacement reaction: a precipitate, water, or a gas.

**Reactions that form precipitates** Some reactions that occur in aqueous solutions produce precipitates. For example, recall from Section 9.2 that when aqueous solutions of sodium hydroxide and copper(II) chloride are mixed, a double-replacement reaction occurs in which the precipitate copper(II) hydroxide forms.



Note that the chemical equation does not show some details of this reaction. Sodium hydroxide and copper(II) chloride are ionic compounds. Therefore, in aqueous solutions they exist as  $\text{Na}^+$ ,  $\text{OH}^-$ ,  $\text{Cu}^{2+}$ , and  $\text{Cl}^-$  ions, as shown in **Figure 9.17**. When their solutions are combined,  $\text{Cu}^{2+}$  ions in one solution and  $\text{OH}^-$  ions in the other solution react to form the precipitate copper(II) hydroxide,  $\text{Cu}(\text{OH})_2(\text{s})$ . The  $\text{Na}^+$  and  $\text{Cl}^-$  ions remain dissolved in the new solution.

■ **Figure 9.17** Like the aqueous solution of HCl in **Figure 9.16**, sodium hydroxide ( $\text{NaOH}$ ) in an aqueous solution dissociates into sodium ( $\text{Na}^+$ ) and hydroxide ( $\text{OH}^-$ ) ions. Copper chloride ( $\text{CuCl}_2$ ) also dissociates into  $\text{Cu}^{2+}$  and  $\text{Cl}^-$  ions.



## Observe a Precipitate-Forming Reaction

How do two liquids form a solid?

### Procedure

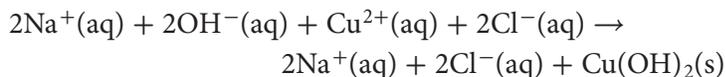
1. Read and complete the lab safety form.
2. Place 50 mL **distilled water** in a **150-mL beaker**.
3. Measure about 4 g **NaOH pellets** on a **balance**. Add the NaOH pellets to the beaker one at a time. Mix with a **stirring rod** until each NaOH pellet dissolves before adding the next pellet.
4. Measure about 6 g **Epsom salts (MgSO<sub>4</sub>)** and place it in another 150-mL beaker. Add 50 mL distilled water to the Epsom salts. Mix with another stirring rod until the Epsom salts dissolve.

5. Slowly pour the Epsom salts solution into the NaOH solution. Record your observations.
6. Stir the new solution. Record your observations.
7. Allow the precipitate to settle, then decant the liquid from the solid into a **100-mL graduated cylinder**.
8. Dispose of the solid as instructed by your teacher.

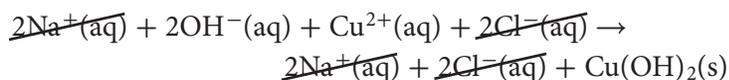
### Analysis

1. **Write** a balanced chemical equation for the reaction between the NaOH and MgSO<sub>4</sub>. Note that most sulfate compounds exist as ions in aqueous solutions.
2. **Write** the complete ionic equation for this reaction.
3. **Determine** which ions are spectator ions, then write the net ionic equation for this reaction.

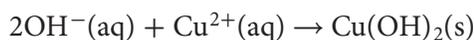
**Ionic equations** To show the details of reactions that involve ions in aqueous solutions, chemists use ionic equations. Ionic equations differ from chemical equations in that substances that are ions in solution are written as ions in the equation. Look again at the reaction between aqueous solutions of sodium hydroxide and copper(II) chloride. To write the ionic equation for this reaction, you must show the reactants, NaOH(aq) and CuCl<sub>2</sub>(aq), and the product, NaCl(aq), as ions.



An ionic equation that shows all of the particles in a solution as they exist is called a **complete ionic equation**. Note that the sodium ions and the chloride ions are both reactants and products. Because they are both reactants and products, they do not participate in the reaction. Ions that do not participate in a reaction are called **spectator ions** and are not usually shown in ionic equations. **Net ionic equations** are ionic equations that include only the particles that participate in the reaction. Net ionic equations are written from complete ionic equations by removing all spectator ions. For example, a net ionic equation is what remains after the sodium and chloride ions are crossed out of this complete ionic equation.



Only the hydroxide and copper ions are left in the net ionic equation shown below.



**Reading Check Compare** How are ionic equations different from chemical equations?

## EXAMPLE Problem 9.3

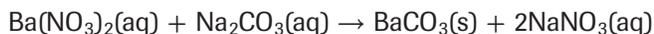
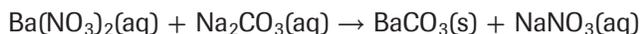
**Reactions That Form a Precipitate** Write the chemical, complete ionic, and net ionic equations for the reaction between aqueous solutions of barium nitrate and sodium carbonate that forms the precipitate barium carbonate.

### 1 Analyze the Problem

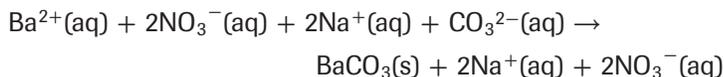
You are given the word equation for the reaction between barium nitrate and sodium carbonate. You must determine the chemical formulas and relative amounts of all reactants and products to write the balanced chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation, you can write the net ionic equation. The net ionic equation will include fewer substances than the other equations.

### 2 Solve for the Unknown

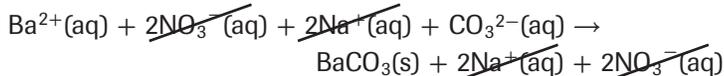
Write the correct chemical formulas and physical states for all substances involved in the reaction.



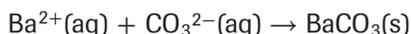
Balance the skeleton equation.



Show the ions of the reactants and the products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation.

### 3 Evaluate the Answer

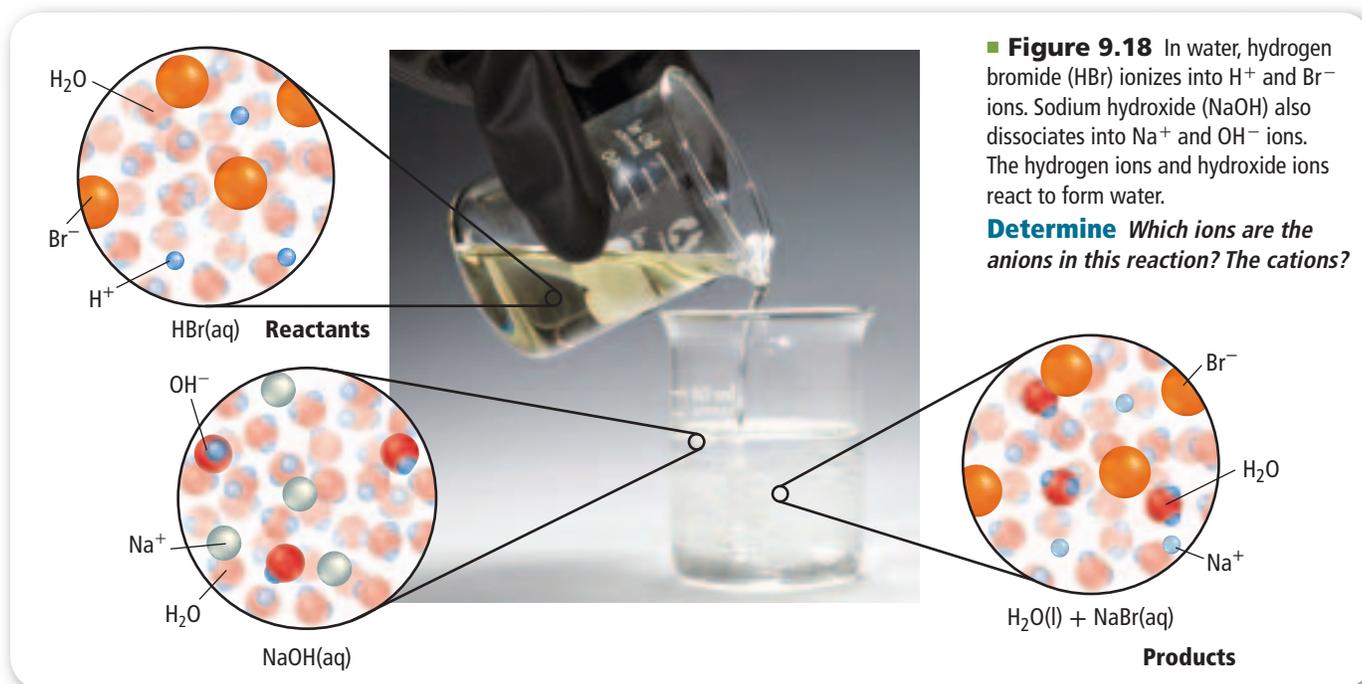
The net ionic equation includes fewer substances than the other equations because it shows only the reacting particles. The particles composing the solid precipitate that is the result of the reaction are no longer ions.

## PRACTICE Problems

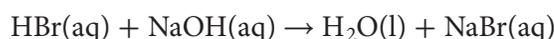
Extra Practice Page 981 and [glencoe.com](http://glencoe.com)

**Write chemical, complete ionic, and net ionic equations for each of the following reactions that might produce a precipitate. Use *NR* to indicate that no reaction occurs.**

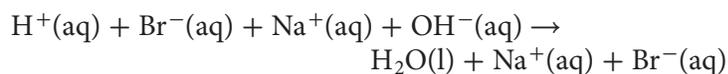
35. Aqueous solutions of potassium iodide and silver nitrate are mixed, forming the precipitate silver iodide.
36. Aqueous solutions of ammonium phosphate and sodium sulfate are mixed. No precipitate forms and no gas is produced.
37. Aqueous solutions of aluminum chloride and sodium hydroxide are mixed, forming the precipitate aluminum hydroxide.
38. Aqueous solutions of lithium sulfate and calcium nitrate are mixed, forming the precipitate calcium sulfate.
39. **Challenge** When aqueous solutions of sodium carbonate and manganese(V) chloride are mixed, a precipitate forms. The precipitate is a compound containing manganese.



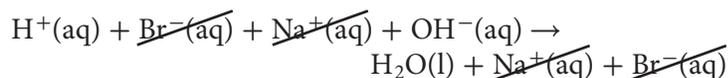
**Reactions that form water** Another type of double-replacement reaction that occurs in an aqueous solution produces water molecules. The water molecules produced in the reaction increase the number of solvent particles. Unlike reactions in which a precipitate forms, no evidence of a chemical reaction is observable because water is colorless, odorless, and already makes up most of the solution. For example, when you mix hydrobromic acid (HBr) with a sodium hydroxide solution (NaOH), as shown in **Figure 9.18**, a double-replacement reaction occurs and water is formed. The chemical equation for this reaction is shown below.



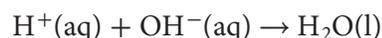
In this case, the reactants and the product sodium bromide exist as ions in an aqueous solution. The complete ionic equation for this reaction shows these ions.



Look carefully at the complete ionic equation. The reacting solute ions are the hydrogen ions and hydroxide ions because the sodium ions and bromine ions are both spectator ions. If you cross out the spectator ions, you are left with the ions that take part in the reaction.



This equation is the net ionic equation for the reaction.



 **Reading Check Analyze** In the reaction between hydrobromic acid and sodium hydroxide, why are the sodium ions and bromine ions called spectator ions?

## EXAMPLE Problem 9.4

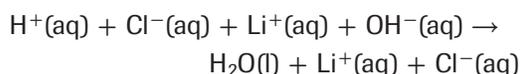
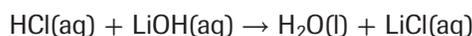
**Reactions That Form Water** Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous lithium hydroxide. This reaction produces water and aqueous lithium chloride.

### 1 Analyze the Problem

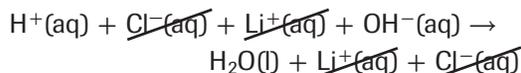
You are given the word equation for the reaction that occurs between hydrochloric acid and aqueous lithium hydroxide to produce water and aqueous lithium chloride. You must determine the chemical formulas for and relative amounts of all reactants and products to write the balanced chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation, you can write the net ionic equation.

### 2 Solve for the Unknown

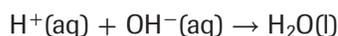
Write the skeleton equation for the reaction and balance it.



Show the ions of the reactants and the products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation.

### 3 Evaluate the Answer

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produces water. The particles that compose the product water are no longer ions.

## PRACTICE Problems

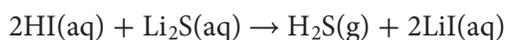
Extra Practice Page 981 and [glencoe.com](http://glencoe.com)

Write chemical, complete ionic, and net ionic equations for the reactions between the following substances, which produce water.

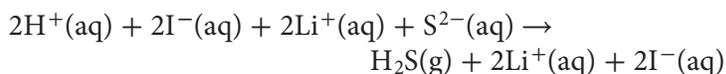
- Mixing sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and aqueous potassium hydroxide produces water and aqueous potassium sulfate.
- Mixing hydrochloric acid ( $\text{HCl}$ ) and aqueous calcium hydroxide produces water and aqueous calcium chloride.
- Mixing nitric acid ( $\text{HNO}_3$ ) and aqueous ammonium hydroxide produces water and aqueous ammonium nitrate.
- Mixing hydrosulfuric acid ( $\text{H}_2\text{S}$ ) and aqueous calcium hydroxide produces water and aqueous calcium sulfate.
- Challenge** When benzoic acid ( $\text{C}_6\text{H}_5\text{COOH}$ ) and magnesium hydroxide are mixed, water and magnesium benzoate are produced.

**Reactions that form gases** A third type of double-replacement reaction that occurs in aqueous solutions results in the formation of a gas. Some gases commonly produced in these reactions are carbon dioxide, hydrogen cyanide, and hydrogen sulfide.

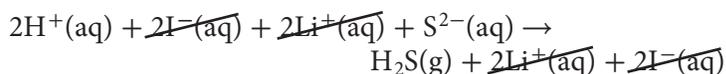
A gas-producing reaction occurs when you mix hydroiodic acid (HI) with an aqueous solution of lithium sulfide. Bubbles of hydrogen sulfide gas form in the container during the reaction. Lithium iodide is also produced in this reaction and remains dissolved in the solution.



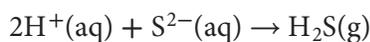
The reactants hydroiodic acid and lithium sulfide exist as ions in aqueous solution. Therefore, you can write an ionic equation for this reaction. The complete ionic equation includes all of the substances in the solution.



Note that there are many spectator ions in the equation. When the spectator ions are crossed out, only the substances involved in the reaction remain in the equation.



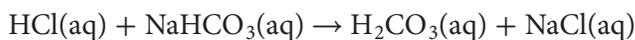
This is the net ionic equation.



If you completed the Launch Lab at the beginning of this chapter, you observed another gas-producing reaction. In that reaction, carbon dioxide gas was produced and bubbled out of the solution. Another reaction that produces carbon dioxide gas occurs in your kitchen when you mix vinegar and baking soda. Vinegar is an aqueous solution of acetic acid and water. Baking soda essentially consists of sodium hydrogen carbonate. Rapid bubbling occurs when vinegar and baking soda are combined. The bubbles are carbon dioxide gas escaping from the solution. You can see this reaction occurring in **Figure 9.19**.

A reaction similar to the one between vinegar and baking soda occurs when you combine any acidic solution and sodium hydrogen carbonate. In all cases, two reactions must occur almost simultaneously in the solution to produce the carbon dioxide gas. One of these is a double-replacement reaction and the other is a decomposition reaction.

For example, when you dissolve sodium hydrogen carbonate in hydrochloric acid, a gas-producing double-replacement reaction occurs. The hydrogen in the hydrochloric acid and the sodium in the sodium hydrogen carbonate replace each other.



Sodium chloride is an ionic compound, and its ions remain separate in the aqueous solution. However, as the carbonic acid ( $\text{H}_2\text{CO}_3$ ) forms, it decomposes immediately into water and carbon dioxide.



■ **Figure 9.19** When vinegar and baking soda (sodium hydrogen carbonate,  $\text{NaHCO}_3$ ) combine, the result is a vigorous bubbling that releases carbon dioxide ( $\text{CO}_2$ ).



## EXAMPLE Problem 9.5

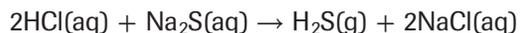
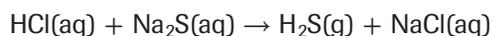
**Reactions That Form Gases** Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous sodium sulfide, which produces hydrogen sulfide gas.

### 1 Analyze the Problem

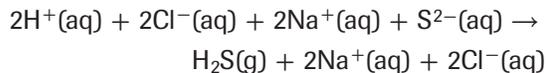
You are given the word equation for the reaction between hydrochloric acid (HCl) and sodium sulfide (Na<sub>2</sub>S). You must write the skeleton equation and balance it. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions in the complete ionic equation, you can write the net ionic equation.

### 2 Solve for the Unknown

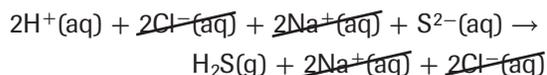
Write the correct skeleton equation for the reaction.



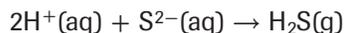
**Balance the skeleton equation.**



**Show the ions of the reactants and the products.**



**Cross out the spectator ions from the complete ionic equation.**



**Write the net ionic equation in its smallest whole-number ratio.**

### 3 Evaluate the Answer

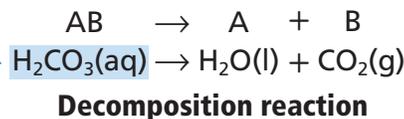
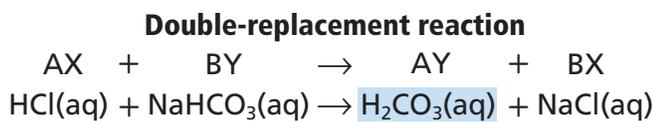
The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produce hydrogen sulfide. The particles that compose the product are no longer ions.

## PRACTICE Problems

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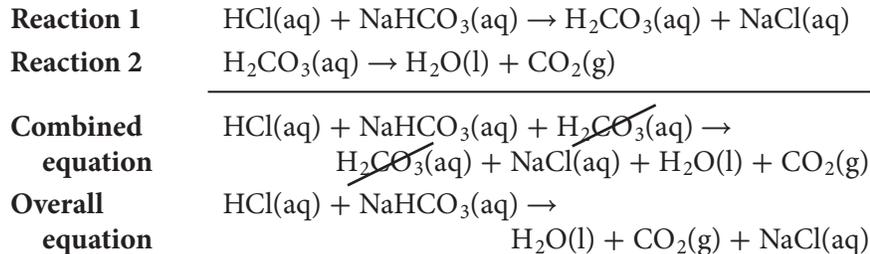
**Write chemical, complete ionic, and net ionic equations for these reactions.**

- Perchloric acid (HClO<sub>4</sub>) reacts with aqueous potassium carbonate, forming carbon dioxide gas and water.
- Sulfuric acid (H<sub>2</sub>SO<sub>4</sub>) reacts with aqueous sodium cyanide, forming hydrogen cyanide gas and aqueous sodium sulfate.
- Hydrobromic acid (HBr) reacts with aqueous ammonium carbonate, forming carbon dioxide gas and water.
- Nitric acid (HNO<sub>3</sub>) reacts with aqueous potassium rubidium sulfide, forming hydrogen sulfide gas.
- Challenge** Aqueous potassium iodide reacts with lead nitrate in solution, forming solid lead iodide.

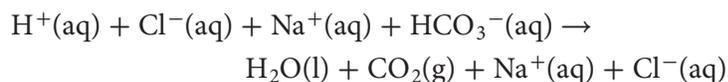


■ **Figure 9.20** When HCl is combined with NaHCO<sub>3</sub>, a double-replacement reaction takes place, followed immediately by a decomposition reaction.

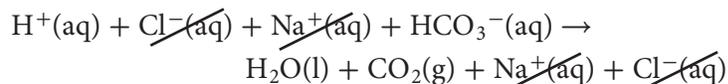
**Overall equations** Recall that when you combine an acidic solution, such as hydrochloric acid, and sodium hydrogen carbonate, two reactions occur—a double-replacement reaction and a decomposition reaction. These reactions are shown in **Figure 9.20**. The two reactions can be combined and represented by one chemical equation in a process similar to adding mathematical equations. An equation that combines two reactions is called an overall equation. To write an overall equation, the reactants in the two reactions are written on the reactant side of the combined equation, and the products of the two reactions are written on the product side. Then, any substances that are on both sides of the equation are crossed out.



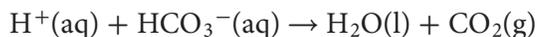
In this case, the reactants in the overall equation exist as ions in aqueous solutions. Therefore, a complete ionic equation can be written for the reaction.



Note that the sodium and chloride ions are the spectator ions. When you cross them out, only the substances that take part in the reaction remain.

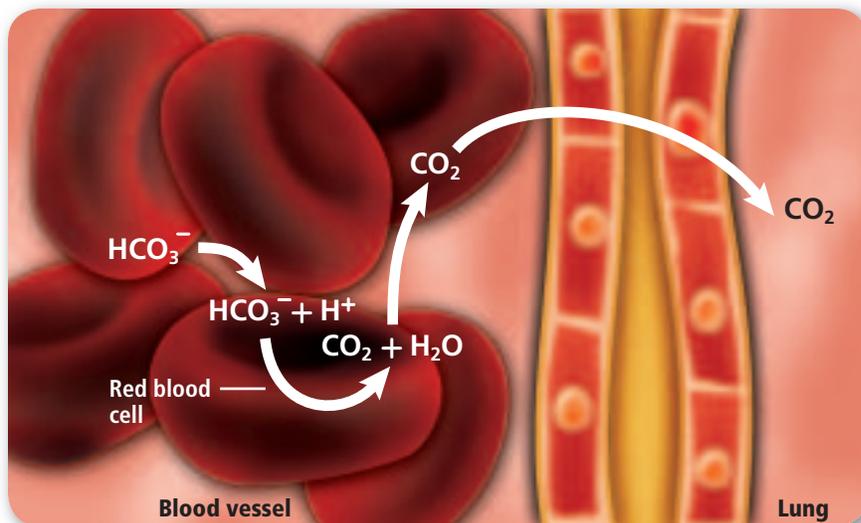


The net ionic equation shows that both water and carbon dioxide gas are produced in this reaction.



**Reading Check Describe** What is an overall equation?

■ **Figure 9.21** After a bicarbonate ion ( $\text{HCO}_3^-$ ) enters a red blood cell, it reacts with a hydrogen ion ( $\text{H}^+$ ) to form water and carbon dioxide ( $\text{CO}_2$ ). The  $\text{CO}_2$  is exhaled from the lungs during respiration.



### CAREERS IN CHEMISTRY

**Biochemist** A biochemist is a scientist who studies the chemical processes of living organisms. A biochemist might study functions of the human body or research how food, drugs, and other substances affect living organisms. For more information on chemistry careers, visit [glencoe.com](http://glencoe.com).

**Connection to Biology** The reaction between hydrogen ions and bicarbonate ions to produce water and carbon dioxide is an important one in your body. This reaction is occurring in the blood vessels of your lungs as you read these words. As shown in **Figure 9.21**, the carbon dioxide gas produced in your cells is transported in your blood in the form of bicarbonate ions ( $\text{HCO}_3^-$ ). In the blood vessels of your lungs, the  $\text{HCO}_3^-$  ions combine with  $\text{H}^+$  ions to produce  $\text{CO}_2$ , which you exhale.

This reaction also occurs in products that are made with baking soda, which contains sodium bicarbonate. Sodium bicarbonate makes baked goods rise. It is used as an antacid and in deodorants to absorb moisture and odors. Baking soda can be added to toothpaste to whiten teeth and freshen breath. As a paste, sodium bicarbonate can be used in cleaning and scrubbing. It is also used as a fire-suppression agent in some fire extinguishers.

## Section 9.3 Assessment

### Section Summary

- ▶ In aqueous solutions, the solvent is always water. There are many possible solutes.
- ▶ Many molecular compounds form ions when they dissolve in water. When some ionic compounds dissolve in water, their ions separate.
- ▶ When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. The solvent molecules do not usually react.
- ▶ Reactions that occur in aqueous solutions are double-replacement reactions.

50. **MAIN Idea** List three common types of products produced by reactions that occur in aqueous solutions.
51. **Describe** solvents and solutes in an aqueous solution.
52. **Distinguish** between a complete ionic equation and a net ionic equation.
53. **Write** complete ionic and net ionic equations for the reaction between sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and calcium carbonate ( $\text{CaCO}_3$ ).  

$$\text{H}_2\text{SO}_4(\text{aq}) + \text{CaCO}_3(\text{s}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) + \text{CaSO}_4(\text{aq})$$
54. **Analyze** Complete and balance the following equation.  

$$\text{CO}_2(\text{g}) + \text{HCl}(\text{aq}) \rightarrow$$
55. **Predict** What type of product would the following reaction be most likely to produce? Explain your reasoning.  

$$\text{Ba}(\text{OH})_2(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow$$
56. **Formulate Equations** A reaction occurs when nitric acid ( $\text{HNO}_3$ ) is mixed with an aqueous solution of potassium hydrogen carbonate. Aqueous potassium nitrate is produced. Write the chemical and net ionic equations for the reaction.

# HOW IT WORKS

## Lighting Up the Night: Bioluminescence

In the gathering darkness, a male firefly announces his presence by sending a signal in yellow-green light. A female near the ground answers his call, and he descends. The result might be a successful mating, or, if the female of another firefly species has fooled the male, he might be greedily devoured. The production of light by the firefly is the result of a chemical process called bioluminescence. This process is a strategy used by a wide variety of living things in many different environments. How does it work?

**1 Flashy Beetles** Fireflies (or lightning bugs) are not flies at all, but a group of beetles that flash their mating signals. They also use their light to lure their prey. The yellow-green light comes from cells in their lower abdomen. The wavelength for this light is between 510 and 670 nm.

**2 Bioluminescence** The glow of the firefly is the result of a chemical reaction. The reactants are oxygen and luciferin, a light-emitting substance found in some organisms. An enzyme, luciferase, speeds up the reaction. The products of this reaction are oxyluciferin and energy, in the form of light.



**3 Glowing Discoveries** Research into bioluminescence led to the discovery of green fluorescent protein (GFP), which is found in some species of jellyfish. GFP emits a green light when exposed to UV light. Researchers have inserted GFP into various organisms, such as mice, for research purposes. Examples of what scientists are using GFP to study include cancer, malaria, and cellular processes.



### WRITING in Chemistry

**Research** Identify different life forms that use bioluminescence and create a pamphlet showing how bioluminescence is effective in each of these organisms. For more information, visit [glencoe.com](http://glencoe.com).

## DEVELOP AN ACTIVITY SERIES

**Background:** Some metals are more reactive than others. By comparing how different metals react with the known ions in aqueous solutions, an activity series for the tested materials can be developed. The activity series will reflect the relative reactivity of the tested metals.

**Question:** How is an activity series developed?

## Materials

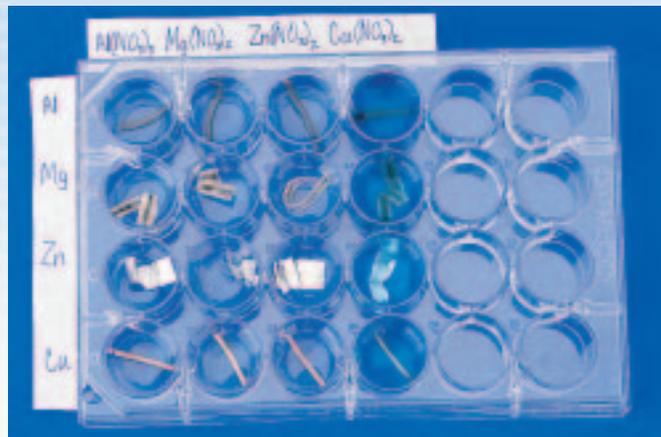
1.0M $\text{Zn}(\text{NO}_3)_2$	Al wire
1.0M $\text{Al}(\text{NO}_3)_3$	Mg ribbon
1.0M $\text{Cu}(\text{NO}_3)_2$	Zn metal strips (4)
1.0M $\text{Mg}(\text{NO}_3)_2$	Emery cloth or sandpaper
pipettes (4)	24-well microscale
wire cutters	reaction plate
Cu wire	

## Safety Precautions



## Procedure

1. Read and complete the lab safety form.
2. Create a table to record your data.
3. Use a pipette to fill each of the four wells in column 1 of the reaction plate with 2 mL of 1.0M  $\text{Al}(\text{NO}_3)_3$  solution.
4. Repeat the procedure in Step 3 to fill the four wells in column 2 with 2 mL of 1.0M  $\text{Mg}(\text{NO}_3)_2$ .
5. Repeat the procedure in Step 3 to fill the four wells in column 3 with 2 mL of 1.0M  $\text{Zn}(\text{NO}_3)_2$ .
6. Repeat the procedure in Step 3 to fill the four wells in column 4 with 2 mL of 1.0M  $\text{Cu}(\text{NO}_3)_2$ .
7. With the emery cloth or sandpaper, polish 10 cm of aluminum wire until it is shiny. Use wire cutters to carefully cut the aluminum wire into four 2.5-cm pieces. Place a piece of the aluminum wire in each well of row A containing solution.
8. Repeat the procedure in Step 7 using 10 cm of magnesium ribbon. Place a piece of Mg ribbon in each well of row B containing solution.
9. Use the emery cloth or sandpaper to polish each small strip of zinc metal. Place a piece of Zn metal in each well of row C containing solution.
10. Observe what happens in each well. After 5 minutes, record your observations in the data table you made.



11. **Cleanup and Disposal** Dispose of the chemicals, solutions, and pipettes as directed by your teacher. Wash and return all lab equipment to the designated location. Wash your hands thoroughly.

## Analyze and Conclude

1. **Observe and Infer** In which wells of the reaction plate did chemical reactions occur? Which metal reacted with the most solutions? Which metal reacted with the fewest solutions? Which metal is the most reactive?
2. **Sequence** The most-active metal reacted with the most solutions. The least-active metal reacted with the fewest solutions. Order the four metals from most active to least active.
3. **Apply** Write a chemical equation for each single-replacement reaction that occurred on your reaction plate.
4. **Real-World Chemistry** Under what circumstances might it be important to know the activity tendencies of a series of elements?
5. **Error Analysis** How does your answer from Question 2 above compare with the activity series in **Figure 9.13**? What could account for the differences?

## INQUIRY EXTENSION

**Design an Experiment** Think of three “what if” questions about this investigation that might affect your results. Design an experiment to test one of them.



**BIG Idea** Millions of chemical reactions in and around you transform reactants into products, resulting in the absorption or release of energy.

## Section 9.1 Reactions and Equations

**MAIN Idea** Chemical reactions are represented by balanced chemical equations.

### Vocabulary

- chemical equation (p. 285)
- chemical reaction (p. 282)
- coefficient (p. 285)
- product (p. 283)
- reactant (p. 283)

### Key Concepts

- Some physical changes are evidence that indicate a chemical reaction has occurred.
- Word equations and skeleton equations provide important information about a chemical reaction.
- A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction.
- Balancing an equation involves adjusting the coefficients until the number of atoms of each element is equal on both sides of the equation.

## Section 9.2 Classifying Chemical Reactions

**MAIN Idea** There are four types of chemical reactions: synthesis, combustion, decomposition, and replacement reactions.

### Vocabulary

- combustion reaction (p. 290)
- decomposition reaction (p. 292)
- double-replacement reaction (p. 296)
- precipitate (p. 296)
- single-replacement reaction (p. 293)
- synthesis reaction (p. 289)

### Key Concepts

- Classifying chemical reactions makes them easier to understand, remember, and recognize.
- Activity series of metals and halogens can be used to predict if single-replacement reactions will occur.

## Section 9.3 Reactions in Aqueous Solutions

**MAIN Idea** Double-replacement reactions occur between substances in aqueous solutions and produce precipitates, water, or gases.

### Vocabulary

- aqueous solution (p. 299)
- complete ionic equation (p. 301)
- net ionic equation (p. 301)
- solute (p. 299)
- solvent (p. 299)
- spectator ion (p. 301)

### Key Concepts

- In aqueous solutions, the solvent is always water. There are many possible solutes.
- Many molecular compounds form ions when they dissolve in water. When some ionic compounds dissolve in water, their ions separate.
- When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. The solvent molecules do not usually react.
- Reactions that occur in aqueous solutions are double-replacement reactions.

## Section 9.1

## Mastering Concepts

57. Define *chemical equation*.
58. Distinguish between a chemical reaction and a chemical equation.
59. Explain the difference between reactants and products.
60. What do the arrows and coefficients in equations communicate?
61. Does a conversion of a substance into a new substance always indicate that a chemical reaction has occurred? Explain.
62. Write formulas for the following substances and designate their physical states.
- nitrogen dioxide gas
  - liquid gallium
  - barium chloride dissolved in water
  - solid ammonium carbonate
63. Identify the reactants in the following reaction: When potassium is dropped into aqueous zinc nitrate, zinc and aqueous potassium nitrate form.
64. Balance the reaction of hydrogen sulfide with atmospheric oxygen gas.
- $$\text{H}_2\text{S}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{s}) + \text{H}_2\text{O}(\text{g})$$
65. Write word equations for the following skeleton equations.
- $\text{Cu}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CuO}(\text{s})$
  - $\text{K}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{KOH}(\text{aq}) + \text{H}_2(\text{g})$
  - $\text{CaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + \text{NaCl}(\text{aq})$
66. Balance the following reactions.
- $(\text{NH}_4)_2\text{Cr}_2\text{O}_7(\text{s}) \rightarrow \text{Cr}_2\text{O}_3(\text{s}) + \text{N}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
  - $\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(\text{s}) + \text{O}_2(\text{g})$

## Mastering Problems

67. Hydrogen iodide gas breaks down into hydrogen gas and iodine gas during a decomposition reaction. Write a skeleton equation for this reaction.
68. Write skeleton equations for these reactions.
- sodium carbonate(s)  $\rightarrow$   
sodium oxide(s) + carbon dioxide(g)
  - aluminum(s) + iodine(s)  $\rightarrow$  aluminum iodide(s)
  - iron(II) oxide(s) + oxygen(g)  $\rightarrow$  iron(III) oxide(s)
69. Write skeleton equations for these reactions.
- butane ( $\text{C}_4\text{H}_{10}$ )(l) + oxygen(g)  $\rightarrow$   
carbon dioxide(g) + water(l)
  - aluminum carbonate(s)  $\rightarrow$   
aluminum oxide(s) + carbon dioxide(g)
  - silver nitrate(aq) + sodium sulfide(aq)  $\rightarrow$   
silver sulfide(s) + sodium nitrate(aq)

70. Write a skeleton equation for the reaction between lithium(s) and chlorine gas to produce lithium chloride(s).
71. Write skeleton equations for these reactions.
- iron(s) + fluorine(g)  $\rightarrow$  iron(III) fluoride(s)
  - sulfur trioxide(g) + water(l)  $\rightarrow$  sulfuric acid(aq)
  - sodium(s) + magnesium iodide(aq)  $\rightarrow$   
sodium iodide(aq) + magnesium(s)
  - vanadium(s) + oxygen(g)  $\rightarrow$  vanadium(V) oxide(s)
72. Write skeleton equations for these reactions.
- lithium(s) + gold(III) chloride(aq)  $\rightarrow$   
lithium chloride(aq) + gold(s)
  - iron(s) + tin(IV) nitrate(aq)  $\rightarrow$   
iron(III) nitrate(aq) + tin(s)
  - nickel(II) chloride(s) + oxygen(g)  $\rightarrow$   
nickel(II) oxide(s) + dichlorine pentoxide(g)
  - lithium chromate(aq) + barium chloride(aq)  $\rightarrow$   
lithium chloride(aq) + barium chromate(s)
73. Balance the skeleton equations for the reactions described in Question 71.
74. Balance the skeleton equations for the reactions described in Question 72.
75. Write chemical equations for these reactions.
- When solid naphthalene ( $\text{C}_{10}\text{H}_8$ ) burns in air, the reaction yields gaseous carbon dioxide and liquid water.
  - Bubbling hydrogen sulfide gas through manganese(II) chloride dissolved in water results in the formation of the precipitate manganese(II) sulfide and hydrochloric acid.
  - Solid magnesium reacts with nitrogen gas to produce solid magnesium nitride.
  - Heating oxygen difluoride gas yields oxygen gas and fluorine gas.

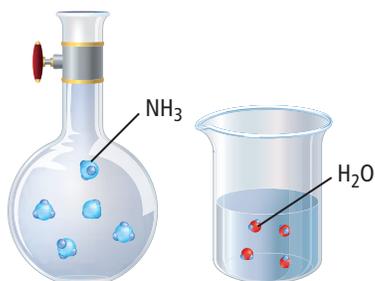
## Section 9.2

## Mastering Concepts

76. List each of the four types of chemical reactions and give an example for each type.
77. How would you classify a chemical reaction between two reactants that produces one product?
78. Under what conditions does a precipitate form in a chemical reaction?
79. Will a metal always replace another metal in a compound dissolved in water? Explain.
80. In each of the following pairs, which element will replace the other in a reaction?
- tin and sodium
  - lead and silver
  - fluorine and iodine
  - copper and nickel

## Mastering Problems

81. Classify each of the reactions represented by the chemical equations in Question 71.
82. Classify each of the reactions represented by the chemical equations in Question 72.



■ Figure 9.22

83. Use **Figure 9.22** to answer the following questions.
- Write a chemical equation for the reaction between the two compounds shown in the figure.
  - Classify this reaction.
84. Write a balanced chemical equation for the combustion of liquid methanol ( $\text{CH}_3\text{OH}$ ).
85. Write chemical equations for each of the following synthesis reactions.
- boron + fluorine  $\rightarrow$
  - germanium + sulfur  $\rightarrow$
  - zirconium + nitrogen  $\rightarrow$
  - tetraphosphorus decoxide + water  $\rightarrow$  phosphoric acid
86. **Combustion** Write a chemical equation for the combustion of each of the following substances. If a compound contains carbon and hydrogen, assume that carbon dioxide gas and liquid water are produced.
- solid barium
  - solid boron
  - liquid acetone ( $\text{C}_3\text{H}_6\text{O}$ )
  - liquid octane ( $\text{C}_8\text{H}_{18}$ )
87. Write chemical equations for each of the following decomposition reactions. One or more products might be identified.
- magnesium bromide  $\rightarrow$
  - cobalt(II) oxide  $\rightarrow$
  - titanium(IV) hydroxide  $\rightarrow$   
titanium(IV) oxide + water
  - barium carbonate  $\rightarrow$  barium oxide + carbon dioxide
88. Write chemical equations for the following single-replacement reactions that might occur in water. If no reaction occurs, write *NR* in place of the products.
- nickel + magnesium chloride  $\rightarrow$
  - calcium + copper(II) bromide  $\rightarrow$
  - potassium + aluminum nitrate  $\rightarrow$
  - magnesium + silver nitrate  $\rightarrow$

## Section 9.3

### Mastering Concepts

89. Complete the following word equation.
- $$\text{Solute} + \text{Solvent} \rightarrow$$
90. Define each of the following terms: *solution*, *solvent*, and *solute*.
91. When reactions occur in aqueous solutions, what common types of products are produced?
92. Compare and contrast chemical equations and ionic equations.
93. What is a net ionic equation? How does it differ from a complete ionic equation?
94. Define *spectator ion*.
95. Write the net ionic equation for a chemical reaction that occurs in an aqueous solution and produces water.

### Mastering Problems

96. Complete the following chemical equations.
- $\text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow$
  - $\text{K(s)} + \text{H}_2\text{O(l)} \rightarrow$
97. Complete the following chemical equation.
- $$\text{CuCl}_2\text{(s)} + \text{Na}_2\text{SO}_4\text{(aq)} \rightarrow$$
98. Write complete ionic and net ionic equations for the chemical reaction in Question 97.
99. Write complete ionic and net ionic equations for each of the following reactions.
- $\text{K}_2\text{S(aq)} + \text{CoCl}_2\text{(aq)} \rightarrow 2\text{KCl(aq)} + \text{CoS(s)}$
  - $\text{H}_2\text{SO}_4\text{(aq)} + \text{CaCO}_3\text{(s)} \rightarrow$   
 $\text{H}_2\text{O(l)} + \text{CO}_2\text{(g)} + \text{CaSO}_4\text{(s)}$
  - $2\text{HClO(aq)} + \text{Ca(OH)}_2\text{(aq)} \rightarrow$   
 $2\text{H}_2\text{O(l)} + \text{Ca(ClO)}_2\text{(aq)}$
100. A reaction occurs when hydrosulfuric acid ( $\text{H}_2\text{S}$ ) is mixed with an aqueous solution of iron(III) bromide. The reaction produces solid iron(III) sulfide and aqueous hydrogen bromide. Write the chemical and net ionic equations for the reaction.
101. Write complete ionic and net ionic equations for each of the following reactions.
- $\text{H}_3\text{PO}_4\text{(aq)} + 3\text{RbOH(aq)} \rightarrow 3\text{H}_2\text{O(l)} + \text{Rb}_3\text{PO}_4\text{(aq)}$
  - $\text{HCl(aq)} + \text{NH}_4\text{OH(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{NH}_4\text{Cl(aq)}$
  - $2\text{HI} + (\text{NH}_4)_2\text{S(aq)} \rightarrow \text{H}_2\text{S(g)} + 2\text{NH}_4\text{I(aq)}$
  - $\text{HNO}_3\text{(aq)} + \text{KCN(aq)} \rightarrow \text{HCN(g)} + \text{KNO}_3\text{(aq)}$
102. **Paper** A reaction occurs when sulfurous acid ( $\text{H}_2\text{SO}_3$ ) is mixed with an aqueous solution of sodium hydroxide. The reaction produces aqueous sodium sulfite, a chemical used in manufacturing paper. Write the chemical and net ionic equations for the reaction.

**Mixed Review**

- 103. Photosynthesis** Identify the products in the following reaction that occurs in plants: Carbon dioxide and water react to produce glucose and oxygen.
- 104.** How will aqueous solutions of sucrose and hydrogen chloride differ?
- 105.** Write the word equation for each of these skeleton equations.  $C_6H_6$  is the formula for benzene.
- $C_6H_6(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
  - $CO(g) + O_2(g) \rightarrow CO_2(g)$
  - $Cl_2(g) + NaBr(s) \rightarrow NaCl(s) + Br_2(g)$
  - $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
- 106.** Classify each of the reactions represented by the chemical equations in Question 105.
- 107.** Write skeleton equations for the following reactions.
- ammonium phosphate(aq) + chromium(III) bromide(aq)  $\rightarrow$  ammonium bromide(aq) + chromium(III) phosphate(s)
  - chromium(VI) hydroxide(s)  $\rightarrow$  chromium(VI) oxide(s) + water(l)
  - aluminum(s) + copper(I) chloride(aq)  $\rightarrow$  aluminum chloride(aq) + copper(s)
  - potassium iodide(aq) + mercury(I) nitrate(aq)  $\rightarrow$  potassium nitrate(aq) + mercury(I) iodide(s)
- 108.** Balance the skeleton equations for the reactions described in Question 107.
- 109.** Classify each of the reactions represented by the chemical equations in Question 108.
- 110.** Predict whether each of the following reactions will occur in aqueous solutions. If you predict that a reaction will not occur, explain your reasoning. *Note: Barium sulfate and silver bromide precipitate in aqueous solutions.*
- sodium hydroxide + ammonium sulfate  $\rightarrow$
  - niobium(V) sulfate + barium nitrate  $\rightarrow$
  - strontium bromide + silver nitrate  $\rightarrow$
- 111.** Complete the missing information in the following skeleton equation and balance the chemical equation:  
 $NaOH(aq) + \underline{\hspace{1cm}} \rightarrow 3NaCl(aq) + Al(OH)_3(aq)$
- 112. Precipitate Formation** The addition of hydrochloric acid to beakers containing solutions of either sodium chloride (NaCl) or silver nitrate ( $KNO_3$ ) causes a white precipitate in one of the beakers.
- Which beaker contains a precipitate?
  - What is the precipitate?
  - Write a chemical equation showing the reaction.
  - Classify the reaction.
- 113.** Write the skeleton equation and the balanced chemical equation for the reaction between iron and chlorine.
- 114.** Write a chemical equation representing the decomposition of water into two gaseous products. What are the products?
- 115.** Distinguish between an ionic compound and a molecular compound dissolved in water. Do all molecular compounds ionize when dissolved in water? Explain.
- 116.** Classify the type of reactions that occur in aqueous solutions, and give an example to support your answer.

**Think Critically**

- 117. Explain** how an equation can be balanced even if the number of reactant particles differs from the number of product particles.
- 118. Apply** Describe the reaction of aqueous solutions of sodium sulfide and copper(II) sulfate, producing the precipitate copper(II) sulfide.
- 119. Predict** A piece of aluminum metal is placed in aqueous KCl. Another piece of aluminum is placed in an aqueous  $AgNO_3$  solution. Explain why a chemical reaction does or does not occur in each instance.
- 120. Design an Experiment** You suspect that the water in a lake close to your school might contain lead in the form of  $Pb^{2+}(aq)$  ions. Formulate your suspicion as a hypothesis and design an experiment to test your theory. Write the net ionic equations for the reactions of your experiment. (*Hint: In aqueous solution,  $Pb^{2+}$  forms compounds that are solids with  $Cl^-$ ,  $Br^-$ ,  $I^-$ , and  $SO_4^{2-}$  ions.*)
- 121. Predict** When sodium metal reacts with water, it produces sodium hydroxide, hydrogen gas, and heat. Write balanced chemical equations for Li, Na, and K reacting with water. Use **Figure 9.13** to predict the order of the amount of heat released from least to most amount of heat released.
- 122. Apply** Write the chemical equations and net ionic equations for each of the following reactions that might occur in aqueous solutions. If a reaction does not occur, write *NR* in place of the products. Magnesium phosphate precipitates in an aqueous solution.
- $KNO_3 + CsCl \rightarrow$
  - $Ca(OH)_2 + KCN \rightarrow$
  - $Li_3PO_4 + MgSO_4 \rightarrow$
  - $HBrO + NaOH \rightarrow$
- 123. Analyze** Explain why a nail exposed to air forms rust, whereas the same nail exposed to a pure nitrogen environment does not form rust.
- 124. Evaluate** Write a balanced chemical equation for the reaction of aluminum with oxygen to produce aluminum oxide.

### Challenge Problem

- 125.** A single-replacement reaction occurs between copper and silver nitrate. When 63.5 g of copper reacts with 339.8 g of silver nitrate, 215.8 g of silver is produced. Write a balanced chemical equation for this reaction. What other product formed? What is the mass of the second product?

### Cumulative Review

- 126.** Complete the following problems in scientific notation. Round off to the correct number of significant figures. (*Chapter 2*)
- $(5.31 \times 10^{-2} \text{ cm}) \times (2.46 \times 10^5 \text{ cm})$
  - $(6.42 \times 10^{-2} \text{ g}) \div (3.21 \times 10^{-3} \text{ g})$
  - $(9.87 \times 10^4 \text{ g}) - (6.2 \times 10^3 \text{ g})$
- 127.** Distinguish between a mixture, a solution, and a compound. (*Chapter 3*)
- 128.** Data from chromium's four naturally occurring isotopes is provided in **Table 9.5**. Calculate chromium's atomic mass. (*Chapter 4*)

**Table 9.5 Chromium Isotope Data**

Isotope	Percent Abundance	Mass (amu)
Cr-50	4.35%	49.946
Cr-52	83.79%	51.941
Cr-53	9.50%	52.941
Cr-54	2.36%	53.939

- 129.** Differentiate between electron configuration and electron-dot structure. (*Chapter 5*)
- 130.** Identify the elements by their electron configuration. (*Chapter 5*)
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
  - $[\text{Ne}] 3s^2 3p^4$
  - $[\text{Xe}] 6s^2$
- 131.** Write the electron configuration for the element fitting each description. (*Chapter 6*)
- a metalloid in group 13
  - a nonmetal in group 15, period 3
- 132.** Describe the formation of positive and negative ions. (*Chapter 7*)
- 133.** Write the formula for the compounds made from each of the following pairs of ions. (*Chapter 7*)
- copper(I) and sulfite
  - tin(IV) and fluoride
  - gold(III) and cyanide
  - lead(II) and sulfide

### Additional Assessment

#### WRITING in Chemistry

- 134. Kitchen Chemistry** Make a poster describing chemical reactions that occur in the kitchen.
- 135. Mathematical Equations** Write a report that compares and contrasts chemical equations and mathematical equations.
- 136. Balance Equations** Create a flowchart describing how to balance a chemical equation.

#### DBQ Document-Based Questions

**Solubility** Scientists, in determining whether a precipitate will occur in a chemical reaction, use a solubility rules chart. **Table 9.6** lists the solubility rules for ionic compounds in water.

Data obtained from: Van Der Sluys, W.G. 2001, *J. Chem. Ed.* 78:111–115

**Table 9.6 Solubility Rules for Ionic Compound in Water**

Ionic Compound	Rule
Soluble salts	Group 1 cations and $\text{NH}_4^+$ ions form soluble salts.
	All nitrates are soluble.
	Most halides are soluble, except those of $\text{Pb}^{2+}$ , $\text{Hg}_2^{2+}$ , $\text{Ag}^+$ , and $\text{Cu}^+$ .
Insoluble salts	Most sulfates are soluble, with the exceptions of those of $\text{Ba}^{2+}$ , $\text{Sr}^{2+}$ , and $\text{Pb}^{2+}$ , $\text{Ag}^+$ , $\text{Ca}^{2+}$ , and $\text{Hg}_2^{2+}$ form slightly soluble sulfates.
	Hydroxides, oxides, and sulfides are usually insoluble, except that those of group 1 ions and $\text{NH}_4^+$ are soluble and those of group 2 ions are slightly soluble.
	Chromates, phosphates, and carbonates are usually insoluble, except that those of group 1 ions and $\text{NH}_4^+$ are soluble.

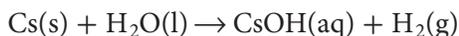
Using the solubility rules provided in the table above, complete the following chemical equations. Indicate whether a precipitate forms or not. Identify the precipitate. If no reaction occurs, write *NR*.

- 137.**  $\text{Ca}(\text{NO}_3)_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow$
- 138.**  $\text{Mg}(\text{s}) + \text{NaOH}(\text{aq}) \rightarrow$
- 139.**  $\text{PbS}(\text{s}) + \text{LiNO}_3(\text{aq}) \rightarrow$

## Standardized Test Practice

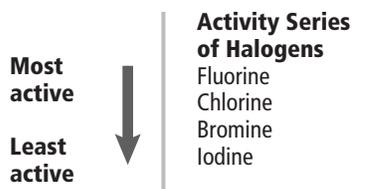
## Multiple Choice

1. What type of reaction is described by the following equation?



- A. synthesis  
B. combustion  
C. decomposition  
D. single-replacement

Use the figure below to answer Question 2.



2. Which reaction between halogens and halide salts will occur?
- A.  $\text{F}_2\text{(g)} + \text{FeI}_2\text{(aq)} \rightarrow \text{FeF}_2\text{(aq)} + \text{I}_2\text{(l)}$   
B.  $\text{I}_2\text{(s)} + \text{MnBr}_2\text{(aq)} \rightarrow \text{MnI}_2\text{(aq)} + \text{Br}_2\text{(g)}$   
C.  $\text{Cl}_2\text{(s)} + \text{SrF}_2\text{(aq)} \rightarrow \text{SrCl}_2\text{(aq)} + \text{F}_2\text{(g)}$   
D.  $\text{Br}_2\text{(l)} + \text{CoCl}_2\text{(aq)} \rightarrow \text{CoBr}_2\text{(aq)} + \text{Cl}_2\text{(g)}$
3. Which is the electron configuration for iron?
- A.  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$   
B.  $[\text{Ar}]3d^6$   
C.  $1s^2 2p^6 3p^6 3d^6$   
D.  $[\text{Ar}]4s^2 4d^6$
4. Which is a description of a pattern displayed by elements in the periodic table?
- A. repetition of their physical properties when arranged by increasing atomic radius  
B. repetition of their chemical properties when arranged by increasing atomic mass  
C. periodic repetition of their properties when arranged by increasing atomic number  
D. periodic repetition of their properties when arranged by increasing atomic mass
5. When moving down a group on the periodic table, which two atomic properties follow the same trend?
- A. atomic radius and ionization energy  
B. ionic radius and atomic radius  
C. ionization energy and ionic radius  
D. ionic radius and electronegativity

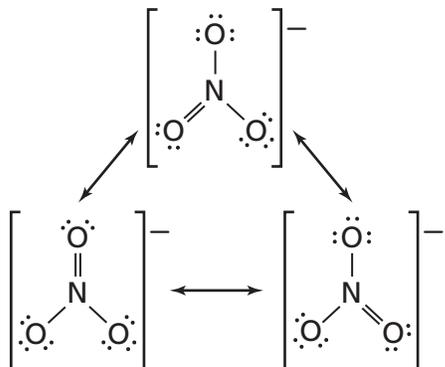
Use the table below to answer Questions 6 to 8.

Physical Properties of Select Ionic Compounds				
Compound	Name	State at 25°C	Soluble in Water?	Melting Point (°C)
$\text{NaClO}_3$	sodium chlorate	solid	yes	248
$\text{Na}_2\text{SO}_4$	sodium sulfate	solid	yes	884
$\text{NiCl}_2$	nickel(II) chloride	solid	yes	1009
$\text{Ni(OH)}_2$	nickel(II) hydroxide	solid	no	230
$\text{AgNO}_3$	silver nitrate	solid	yes	212

6. An aqueous solution of nickel(II) sulfate is mixed with aqueous sodium hydroxide. Will a visible reaction occur?
- A. No, solid nickel(II) hydroxide is soluble in water.  
B. No, solid sodium sulfate is soluble in water.  
C. Yes, solid sodium sulfate will precipitate out of the solution.  
D. Yes, solid nickel(II) hydroxide will precipitate out of the solution.
7. What happens when  $\text{AgClO}_3\text{(aq)}$  and  $\text{NaNO}_3\text{(aq)}$  are mixed?
- A. No visible reaction occurs.  
B. Solid  $\text{NaClO}_3$  precipitates out of the solution.  
C.  $\text{NO}_2$  gas is released during the reaction.  
D. Solid Ag metal is produced.
8. Finely ground nickel(II) hydroxide is placed in a beaker of water. It sinks to the bottom of the beaker and remains unchanged. An aqueous solution of hydrochloric acid (HCl) is then added to the beaker, and the  $\text{Ni(OH)}_2$  disappears. Which equation best describes what occurred in the beaker?
- A.  $\text{Ni(OH)}_2\text{(s)} + \text{HCl(aq)} \rightarrow \text{NiO(aq)} + \text{H}_2\text{(g)} + \text{HCl(aq)}$   
B.  $\text{Ni(OH)}_2\text{(s)} + 2\text{HCl(aq)} \rightarrow \text{NiCl}_2\text{(aq)} + 2\text{H}_2\text{O(l)}$   
C.  $\text{Ni(OH)}_2\text{(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{NiCl}_2\text{(aq)} + 2\text{H}_2\text{O(l)}$   
D.  $\text{Ni(OH)}_2\text{(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{NiCl}_2\text{(aq)} + 3\text{H}_2\text{O(l)} + \text{O}_2\text{(g)}$

## Short Answer

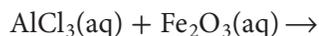
Use the diagram below to answer Questions 9 and 10.



- What is the name for the multiple Lewis structures shown in the diagram?
- Why do these structures form?
- Write the balanced chemical equation for the reaction of solid calcium with water to form calcium hydroxide in solution and hydrogen gas.

## Extended Response

Use the partial chemical equation below to answer Questions 12 and 13.

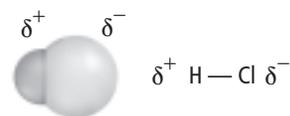


- What type of reaction will this be? Explain how you can tell from the reactants.
- Predict what the products of this reaction will be. Use evidence from the reaction to support your answer.
- What is the electron configuration for the ion  $\text{P}^{3-}$ ? Explain how this configuration is different from the configuration for the neutral atom of phosphorus.

## SAT Subject Test: Chemistry

- Chloroform ( $\text{CHCl}_3$ ) was one of the first anesthetics used in medicine. The chloroform molecule contains 26 valence electrons total. How many of these valence electrons are part of covalent bonds?
  - 26
  - 13
  - 8
  - 4
  - 2
- Which is NOT true of an atom obeying the octet rule?
  - obtains a full set of eight valence electrons
  - acquires the valence configuration of a noble gas
  - electron configuration is unusually stable
  - has an  $s^2p^6$  valence configuration
  - will lose electrons

Use the figure below to answer Question 17.



- Which statement does NOT correctly describe the model of HCl shown above?
  - A nonpolar bond exists between these atoms.
  - Chlorine has a stronger attraction for electrons than does hydrogen.
  - The electrons in the bond are shared unequally.
  - This compound dissolves in a polar substance.
  - Chlorine is the more electronegative atom.
- The combustion of ethanol ( $\text{C}_2\text{H}_6\text{O}$ ) produces carbon dioxide and water vapor. What equation best describes this process?
  - $\text{C}_2\text{H}_6\text{O}(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
  - $\text{C}_2\text{H}_6\text{O}(\text{l}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
  - $\text{C}_2\text{H}_6\text{O}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$
  - $\text{C}_2\text{H}_6\text{O}(\text{l}) \rightarrow 3\text{O}_2(\text{l}) + 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
  - $\text{C}_2\text{H}_6\text{O}(\text{l}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$

### NEED EXTRA HELP?

If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Review Section . . .	9.2	9.2	5.3	6.3	6.3	9.3	9.3	9.3	8.3	8.3	9.1	9.2	9.2	5.3	8.3	8.3	8.5	9.2