7.1 Describing Reactions

Reading Focus

Key Concepts

- What is the law of conservation of mass?
- Why must chemical equations be balanced?
- Why do chemists use the mole?
- How can you calculate the mass of a reactant or product in a chemical reaction?

Vocabulary

- reactants
- products
- chemical equation
- coefficients
- mole
- molar mass

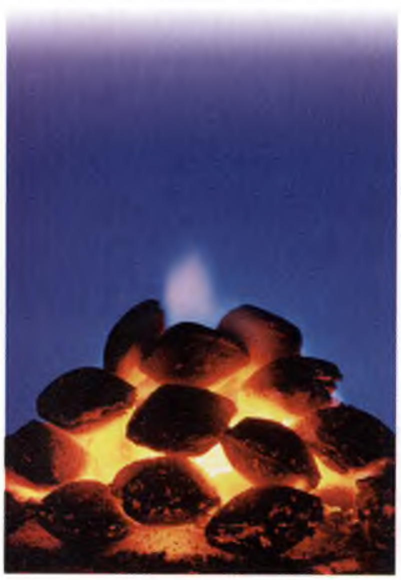
Reading Strategy

Monitoring Your Understanding Preview the Key Concepts, topic headings, vocabulary, and figures in this section. List two things you expect to learn. After reading, state what you learned about each item you listed.

What I Expect to Learn	What I Learned
a. ?	b. ?
c. ?	d. ?

What type of change is happening in Figure 1? When charcoal burns, it changes into other substances while producing heat and light. Burning is a chemical change. When a substance undergoes a chemical change, a chemical reaction is said to take place. In order to understand chemical reactions, you first must be able to describe them.

Figure 1 Burning is an example of a chemical reaction. When charcoal burns, the carbon in the charcoal reacts with oxygen in the air to produce carbon dioxide and heat.



Chemical Equations

A useful way of describing a change is to state what is present before and after the change. For example, suppose you wanted to show how your appearance changed as you grew older. You could compare a photo of yourself when you were younger with a photo that was taken recently.

A useful description of a chemical reaction tells you the substances present before and after the reaction. In a chemical reaction, the substances that undergo change are called **reactants.** The new substances formed as a result of that change are called **products.** In Figure 1, the reactants are the carbon in the charcoal and the oxygen in the air. The product of the reaction is carbon dioxide gas.

Using Equations to Represent Reactions During a chemical reaction, the reactants change into products. You can summarize this process with a word equation.

$$\text{Reactants} \longrightarrow \text{Products}$$

To describe the burning of charcoal, you can substitute the reactants and products of the reaction into the word equation as follows.

$$\text{Carbon} + \text{Oxygen} \longrightarrow \text{Carbon dioxide}$$

You can then simplify the word equation by writing the reactants and products as chemical formulas.

$$C + O_2 \longrightarrow CO_2$$

Now you have a chemical equation. A **chemical equation** is a representation of a chemical reaction in which the reactants and products are expressed as formulas. You can read the equation above as, "Carbon and oxygen react and form carbon dioxide," or, "The reaction of carbon and oxygen yields carbon dioxide."

Reading Checkpoint *What is a chemical equation?*

Conservation of Mass As a piece of charcoal burns, it gets smaller and smaller until it is finally reduced to a tiny pile of ash. Although the charcoal seems to disappear as it burns, it is actually being converted into carbon dioxide gas. If you measured the mass of the carbon dioxide produced, it would equal the mass of the charcoal and oxygen that reacted.

During chemical reactions, the mass of the products is always equal to the mass of the reactants. This principle, established by French chemist Antoine Lavoisier (1743–1794), is known as the law of conservation of mass. **The law of conservation of mass states that mass is neither created nor destroyed in a chemical reaction.** By demonstrating that mass is conserved in various reactions, Lavoisier laid the foundation for modern chemistry.

Figure 2 illustrates how a chemical equation can be restated in terms of atoms and molecules. The equation reads, "One atom of carbon reacts with one molecule of oxygen and forms one molecule of carbon dioxide." Suppose you have six carbon atoms. If each carbon atom reacts with one oxygen molecule to form one carbon dioxide molecule, then six carbon atoms react with six oxygen molecules to form six carbon dioxide molecules. Notice that the number of atoms on the left side of the equation equals the number of atoms on the right. The equation shows that mass is conserved.

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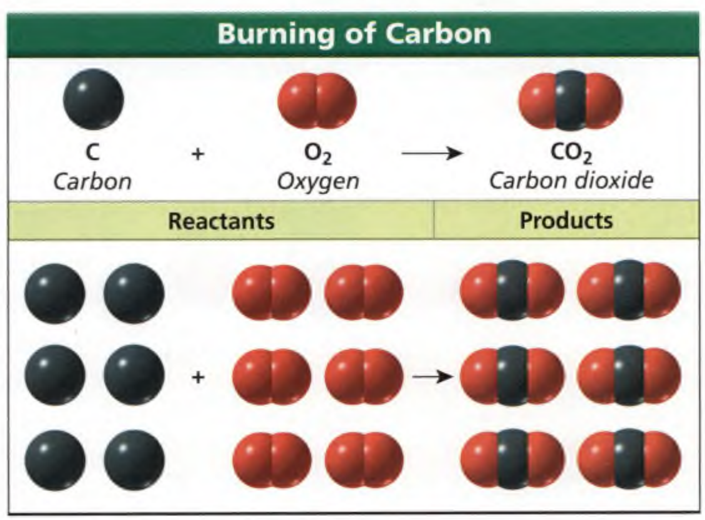
Figure 2 Whether you burn one carbon atom or six carbon atoms, the equation used to describe the reaction is the same. **Using Models** ***How do both models of the reaction below show that mass is conserved?***



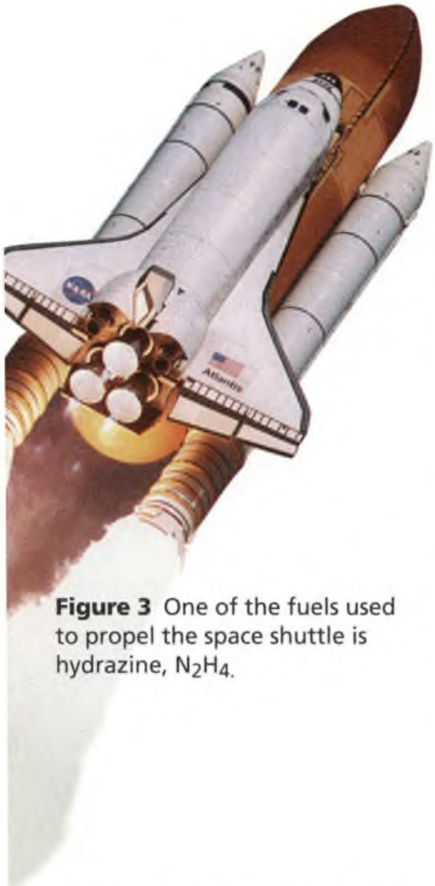


Figure 3 One of the fuels used to propel the space shuttle is hydrazine, N_2H_4.

Balancing Equations

Some reactions are powerful enough to propel spacecraft like the one shown in Figure 3. Many rocket fuels contain a compound called hydrazine, N_2H_4. When hydrazine burns in the presence of oxygen, the reaction produces nitrogen, water vapor, and heat. You can describe this reaction by writing a chemical equation.

$$N_2H_4 + O_2 \longrightarrow N_2 + H_2O$$

If you examine this equation carefully, you will notice that the number of atoms on the left side does not equal the number of atoms on the right. The equation is not balanced. **In order to show that mass is conserved during a reaction, a chemical equation must be balanced.**

You can balance a chemical equation by changing the **coefficients,** the numbers that appear before the formulas. In the unbalanced equation above, the coefficients are understood to be 1. When you change a coefficient, you change the amount of that reactant or product represented in the chemical equation. As you balance equations, you should never change the subscripts in a formula. Changing the formula changes the identity of that reactant or product.

The first step in balancing an equation is to count the number of atoms of each element on each side of the equation. As Figure 4A shows, the left side has two nitrogen atoms, four hydrogen atoms, and two oxygen atoms. The right side has two nitrogen atoms, two hydrogen atoms, and one oxygen atom. The hydrogen and oxygen atoms need to be balanced.

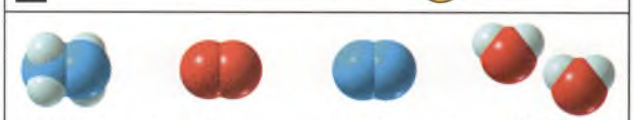
Figure 4 To balance a chemical equation, first count the atoms on each side of the equation. **Applying Concepts** ***Why must chemical equations be balanced?***

Burning of Hydrazine

$N_2H_4 + O_2 \longrightarrow N_2 + H_2O$

Reactants	Products
2 nitrogen atoms	2 nitrogen atoms
4 hydrogen atoms	2 hydrogen atoms
2 oxygen atoms	1 oxygen atom

A Unbalanced equation



$N_2H_4 + O_2 \longrightarrow N_2 + 2H_2O$

Reactants	Products
2 nitrogen atoms	2 nitrogen atoms
4 hydrogen atoms	4 hydrogen atoms
2 oxygen atoms	2 oxygen atoms

B Balanced equation

The next step is to change one or more coefficients until the equation is balanced. You do not have to change the coefficients of N_2H_4 or N_2 because the nitrogen atoms are already balanced. However, the left side has more hydrogen and oxygen atoms than the right side, so you can try increasing the coefficient of water. Try changing the coefficient of water to 2. You then have four hydrogen atoms and two oxygen atoms on the right side, as shown in Figure 4B.

$$N_2H_4 + O_2 \longrightarrow N_2 + 2H_2O$$

The equation is now balanced. It tells you that each molecule of hydrazine requires one molecule of oxygen to react. It also tells you that when one molecule of hydrazine burns, it produces one molecule of nitrogen and two molecules of water.

Reading Checkpoint ***What is a coefficient?***

Math Skills

Balancing Chemical Equations

Write a balanced equation for the reaction between copper and oxygen to produce copper(II) oxide, CuO.

1 Read and Understand

What information are you given?

Reactants: Cu, O_2 Product: CuO

2 Plan and Solve

Write a chemical equation with the reactants on the left side and the product on the right.

$$Cu + O_2 \longrightarrow CuO$$

This equation is not balanced. Change the coefficient of CuO in order to balance the number of oxygen atoms.

$$Cu + O_2 \longrightarrow 2CuO$$

Change the coefficient of Cu in order to balance the number of copper atoms.

$$2Cu + O_2 \longrightarrow 2CuO$$

3 Look Back and Check

Is your answer reasonable?

The number of atoms on the left equals the number of atoms on the right.

Math Practice

1. Hydrogen chloride, or HCl, is an important industrial chemical. Write a balanced equation for the production of hydrogen chloride from hydrogen and chlorine.
2. Balance the following chemical equations.
   a. $H_2O_2 \longrightarrow H_2O + O_2$
   b. $Mg + HCl \longrightarrow H_2 + MgCl_2$
3. Ethylene, C_2H_4, burns in the presence of oxygen to produce carbon dioxide and water vapor. Write a balanced equation for this reaction.

Counting With Moles

How many shoes do you own? Because shoes come in twos, you would most likely count them by the pair rather than individually. The counting units you use depend on what you are counting. For example, you might count eggs by the dozen or paper by the ream.

Chemists also need practical units for counting things. Although you can describe a reaction in terms of atoms and molecules, these units are too small to be practical. **Because chemical reactions often involve large numbers of small particles, chemists use a counting unit called the mole to measure amounts of a substance.**

A **mole** (mol) is an amount of a substance that contains approximately 6.02×10^{23} particles of that substance. This number is known as Avogadro's number. In chemistry, a mole of a substance generally contains 6.02×10^{23} atoms, molecules, or ions of that substance. For instance, a mole of iron is 6.02×10^{23} atoms of iron.

Figure 5 Shoes are often counted by the pair, eggs by the dozen, and paper by the ream (500 sheets). To count particles of a substance, chemists use the mole (6.02×10^{23} particles).

Molar Mass A dozen eggs has a different mass than a dozen oranges. Similarly, a mole of carbon has a different mass than a mole of sulfur, as shown in Figure 6. The mass of one mole of a substance is called a **molar mass.** For an element, the molar mass is the same as its atomic mass expressed in grams. For example, the atomic mass of carbon is 12.0 amu, so the molar mass of carbon is 12.0 grams.

For a compound, you can calculate the molar mass by adding up the atomic masses of its component atoms, and then expressing this sum in grams. A carbon dioxide molecule is composed of one carbon atom (12.0 amu) and two oxygen atoms (2×16.0 amu $= 32.0$ amu). So carbon dioxide has a molar mass of 44.0 grams.

Figure 6 The molar mass of carbon is 12.0 grams. The molar mass of sulfur is 32.1 grams. **Inferring** ***If each of the carbon and sulfur samples contains one mole of atoms, why do the samples have different masses?***

Mole-Mass Conversions Once you know the molar mass of a substance, you can convert moles of that substance into mass, or a mass of that substance into moles. For either calculation, you need to express the molar mass as a conversion factor. For example, the molar mass of CO_2 is 44.0 grams, which means that one mole of CO_2 has a mass of 44.0 grams. This relationship yields the following conversion factors.

$$\frac{44.0 \text{ g } CO_2}{1 \text{ mol } CO_2} \qquad \frac{1 \text{ mol } CO_2}{44.0 \text{ g } CO_2}$$

Suppose you have 55.0 grams of CO_2. To calculate how many moles of CO_2 you have, multiply the mass by the conversion factor on the right.

$$55.0 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.0 \text{ g } CO_2} = 1.25 \text{ mol } CO_2$$

You can check your answer by using the conversion factor on the left.

$$1.25 \text{ mol } CO_2 \times \frac{44.0 \text{ g } CO_2}{1 \text{ mol } CO_2} = 55.0 \text{ g } CO_2$$

Quick Lab

Modeling a Mole

Materials

bolt, 2 nuts, 2 washers, balance

Procedure

1. Measure and record the mass of one bolt, one nut, and one washer. Each piece of hardware will represent an atom of a different element.
2. Assemble the bolt, nuts, and washers together so that they form a model of a molecule known as BN_2W_2.
3. Predict the mass of BN_2W_2.
4. Test your prediction by finding the mass of your model.

Analyze and Conclude

1. **Analyzing Data** Did your prediction match the actual mass of your model? Explain.
2. **Calculating** How many models of your molecule can you make with 20 washers and as many nuts and bolts as you need?
3. **Calculating** How many models of your molecule can you make with 100 grams of nuts and as many bolts and washers as you need?

Chemical Calculations

Think about baking a cake like the one in Figure 7. The directions on a box of cake mix might tell you to add two eggs and one cup of water to the cake mix. Suppose you wanted to make three cakes. Although the directions don't tell you specifically how many eggs and how much water are required for three cakes, you could figure out the amounts. To make three cakes, you would need three times as much of each ingredient—six eggs, three cups of water, and three packages of cake mix.

Chemical equations can be read as recipes for making new substances. Figure 8 shows the balanced equation for the formation of water. You can read this equation as, "Two molecules of hydrogen react with one molecule of oxygen and form two molecules of water." In terms of moles, the equation reads, "Two moles of hydrogen react with one mole of oxygen and form two moles of water." To convert from moles to mass, you need to use the molar masses as conversion factors. Figure 8 shows how the same equation can also be read as, "4.0 grams of H_2 reacts with 32.0 grams of O_2 and forms 36.0 grams of H_2O."

Figure 7 A cake recipe tells you how much of each ingredient to use for each cake you bake. **Using Analogies** ***How is a cake recipe like a chemical equation?***

How many grams of oxygen would you need to make 144 grams of water? **In chemical reactions, the mass of a reactant or product can be calculated by using a balanced chemical equation and molar masses of the reactants and products.** The chemical equation tells you how to relate amounts of reactants to amounts of products. Molar masses let you convert those amounts into masses.

Reading Checkpoint *How do you convert from moles to mass?*

Converting Mass to Moles

To calculate how much oxygen is required to make 144 grams of water, you need to begin with a balanced chemical equation for the reaction.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

The first step in your calculations is to determine how many moles of water you are trying to make. By using the molar mass of water, you can convert the given mass of water into moles.

$$144 \text{ g } H_2O \times \frac{1 \text{ mol H2O}}{18.0 \text{ g } H_2O} = 8.00 \text{ mol } H_2O$$

Formation of Water

Equation	$2H_2$ +	O_2 →	$2H_2O$
Amount	2 mol	1 mol	2 mol
Molar Mass	2.0 g/mol	32.0 g/mol	18.0 g/mol
Mass (Moles × Molar Mass)	4.0 g +	32.0 g →	36.0 g

Figure 8 In a balanced chemical equation, the number of atoms of each element on the left equals the number of atoms of each element on the right. By using molar masses, you can show that the mass of the reactants equals the mass of the products.

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Using Mole Ratios Remember the balanced chemical equation for the formation of water. You can read it as, "Two moles of hydrogen react with one mole of oxygen and form two moles of water." Because each mole of oxygen that reacts will yield two moles of water, you can write the following conversion factors, or mole ratios.

$$\frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} \qquad \frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2}$$

The mole ratio on the left allows you to convert moles of water to moles of oxygen. Now you can calculate how many moles of oxygen are required to produce eight moles of water:

$$8.00 \text{ mol } H_2O \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} = 4.00 \text{ mol } O_2$$

Converting Moles to Mass The last step is to convert moles of O_2 to grams of O_2 by using the molar mass of O_2 as a conversion factor.

$$4.00 \text{ mol } O_2 \times \frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2} = 128 \text{ g } O_2$$

So, in order to produce 144 grams of H_2O, you must supply 128 grams of O_2. Notice that you used the concept of a mole in two ways to solve this problem. In the first and last step, you used a molar mass to convert between mass and moles. In the middle step, you used the mole ratio to convert moles of a product into moles of a reactant.

Section 7.1 Assessment

Reviewing Concepts

1. What is the law of conservation of mass?
2. Why does a chemical equation need to be balanced?
3. Why do chemists use the mole as a counting unit?
4. What information do you need to predict the mass of a reactant or product in a chemical reaction?
5. What is a mole ratio?

Critical Thinking

6. **Applying Concepts** The following equation describes how sodium and chlorine react to produce sodium chloride.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Is the equation balanced? Explain your answer.

7. **Calculating** Ammonia, NH_3, can be made by reacting nitrogen with hydrogen.

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

How many moles of NH_3 can be made if 7.5 moles of H_2 react with enough N_2?

8. **Calculating** What mass of NH_3 can be made from 35.0 g of N_2?

Math Practice

9. Balance the following equation.

$$K + Br_2 \longrightarrow KBr$$

10. Write a balanced chemical equation for the formation of magnesium oxide, MgO, from magnesium and oxygen.

7.2 Types of Reactions

Reading Focus

Key Concepts

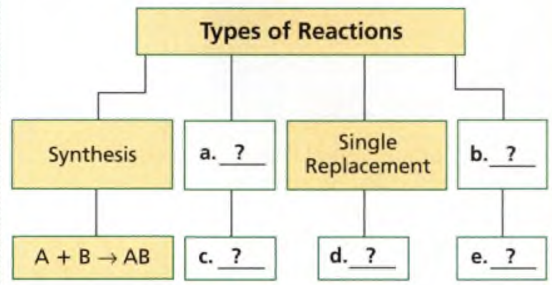
- What are the general types of chemical reactions?
- How did the discovery of subatomic particles affect the classification of reactions?

Vocabulary

- synthesis reaction
- decomposition reaction
- single-replacement reaction
- double-replacement reaction
- combustion reaction
- oxidation-reduction reaction

Reading Strategy

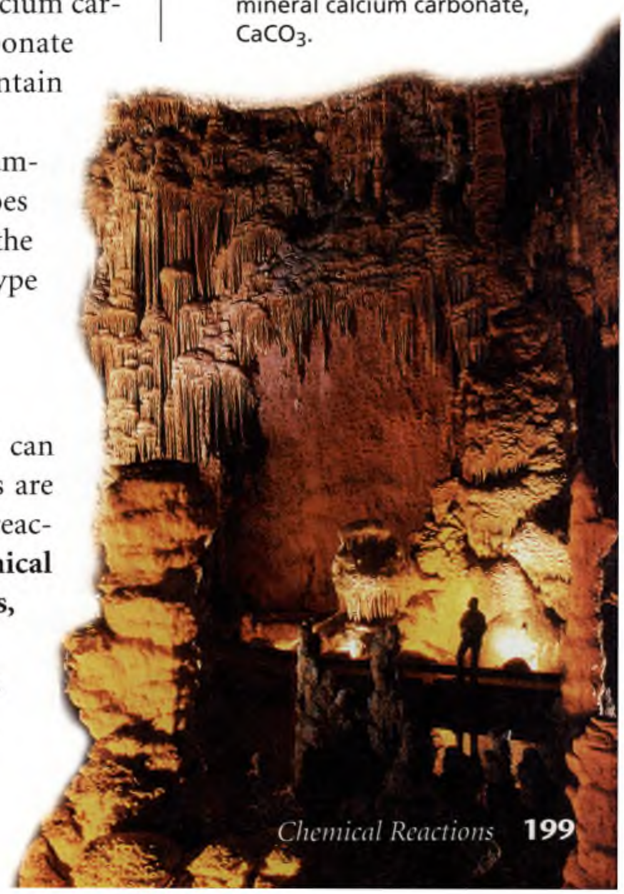
Previewing Skim the section and begin a concept map like the one below that identifies types of reactions with a general form. As you read, add the general form of each type of reaction.



The walls of the cave shown in Figure 9 are solid limestone. When hydrochloric acid is dropped on limestone, a chemical reaction occurs in which a gas is produced. Geologists can use this reaction to determine whether a rock sample contains the mineral calcium carbonate, $CaCO_3$. When a rock containing calcium carbonate reacts with hydrochloric acid, it fizzes. The bubbles contain carbon dioxide gas.

Many other reactions produce carbon dioxide. For example, heating limestone produces carbon dioxide. So does burning gasoline. However, just because two reactions have the same product, you cannot assume that they are the same type of reaction.

Figure 9 The walls and other formations of Blanchard Springs Caverns in Arkansas contain the mineral calcium carbonate, $CaCO_3$.



Classifying Reactions

Just as you can classify matter into different types, you can classify chemical reactions into different types. Reactions are often classified by the type of reactant or the number of reactants and products. **Some general types of chemical reactions are synthesis reactions, decomposition reactions, single-replacement reactions, double-replacement reactions, and combustion reactions.** Each type describes a different way in which reactants interact to form products.

Figure 10 Sodium metal reacts vigorously with chlorine to form sodium chloride, NaCl.
Interpreting Photos ***What evidence in this photograph tells you that a chemical reaction is taking place?***

Synthesis A **synthesis reaction** is a reaction in which two or more substances react to form a single substance. The reactants may be either elements or compounds. The product synthesized is always a compound. The general equation for a synthesis reaction is

$$A + B \longrightarrow AB$$

Figure 10 shows what happens when sodium reacts with chlorine. The product of this reaction is the compound sodium chloride, which appears as a whitish cloud of solid particles. You are probably more familiar with sodium chloride as table salt. You can describe the synthesis of sodium chloride with the following equation.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Another example of a synthesis reaction is hydrogen and oxygen reacting to form water.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

This reaction is used to generate electricity for satellites and spacecraft.

Reading Checkpoint *What is a synthesis reaction?*

Decomposition The opposite of synthesis is decomposition. A **decomposition reaction** is a reaction in which a compound breaks down into two or more simpler substances. The reactant in a decomposition reaction must be a compound. The products may be elements or compounds. The general equation for a decomposition reaction is

$$AB \longrightarrow A + B$$

When electricity passes through water, the water decomposes into hydrogen gas and oxygen gas. You can describe the decomposition of water by writing the following equation.

$$2H_2O \longrightarrow 2H_2 + O_2$$

Notice that this reaction is the opposite of the synthesis of water.

Another example of decomposition occurs in the making of cement. Cement factories use a giant kiln, or oven, to heat a mixture of clay and limestone. The heat causes the calcium carbonate in the limestone to decompose into lime, CaO, and carbon dioxide.

$$CaCO_3 \longrightarrow CaO + CO_2$$

The carbon dioxide escapes the kiln through a smokestack. The clay-and-lime mixture is cooled and ground into cement powder.

The How It Works box on page 201 describes a decomposition reaction that is used to make automobiles safer.

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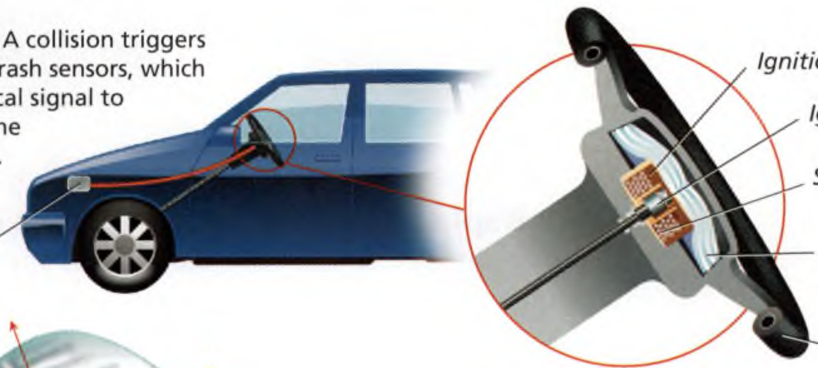
HOW It Works

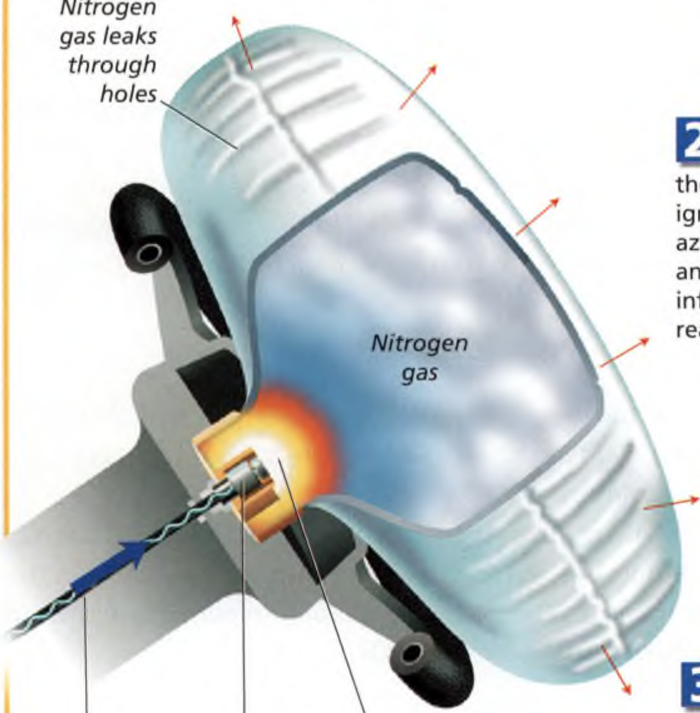
Automobile Safety: Air Bags

Air bags are inflatable cushions built into a car's steering wheel or dashboard. In a crash, the bags inflate, protecting both the driver and the passenger. The whole process takes 0.04 second. **Interpreting Diagrams** ***What is the source of the gas that fills an air bag?***

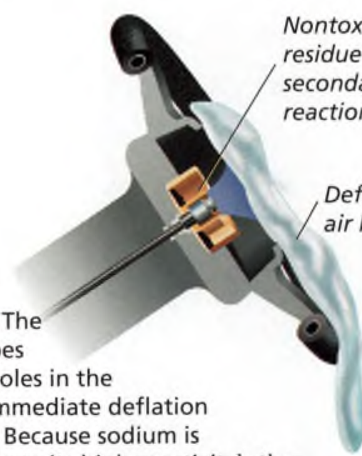
Testing air bags
Air bags have been shown to reduce the risk of serious injury in a head-on collision by 30 percent. New cars have air bags on both the driver and passenger sides.

1 Collision A collision triggers the car's crash sensors, which send an electrical signal to the igniter in the steering wheel.





2 Air bag inflates The igniter sets off a combustion reaction that heats up the sodium azide (NaN_3) contained in the ignition unit. When it is heated, the sodium azide decomposes into metallic sodium (Na) and nitrogen gas (N_2). The nitrogen gas inflates the air bag. The decomposition reaction is $2NaN_3 \rightarrow 2Na + 3N_2$.



3 Air bag deflates The nitrogen escapes through tiny holes in the bag, causing immediate deflation of the air bag. Because sodium is dangerous (due to its high reactivity), the ignition unit also contains other chemicals that react with sodium to form a nontoxic material.

Figure 11 A single-replacement reaction occurs when copper wire is submerged in a solution of silver nitrate. As the copper replaces the silver in the silver nitrate solution, the solution turns blue, and silver crystals form on the wire.

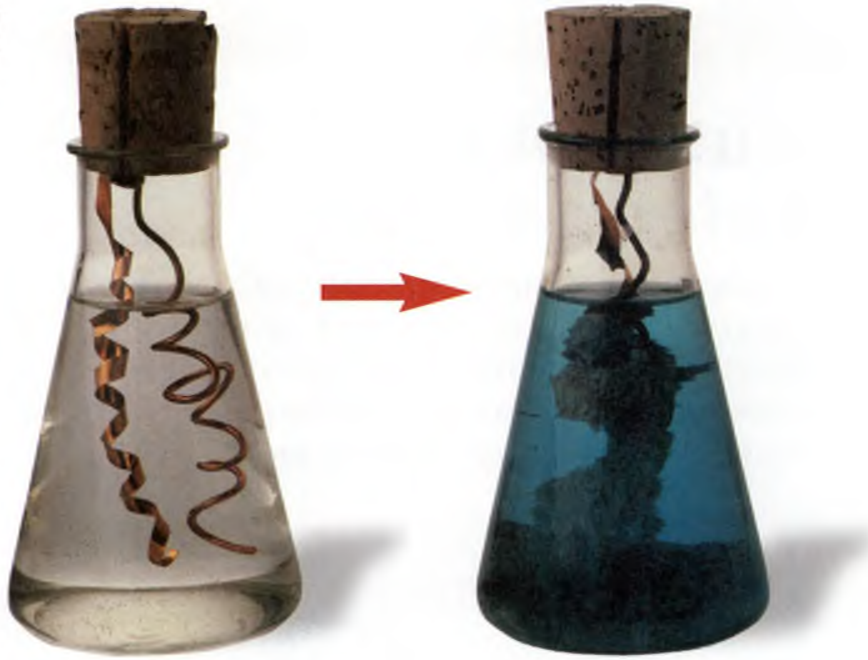


Figure 12 Potassium reacts with water in a single-replacement reaction that produces hydrogen gas and potassium hydroxide.

Single Replacement A **single-replacement reaction** is a reaction in which one element takes the place of another element in a compound. Single-replacement reactions have the general form

$$A + BC \longrightarrow B + AC$$

Suppose you dip a coil of copper wire into a solution of silver nitrate and water, as shown in Figure 11. A vivid chemical reaction takes place as the solution turns blue and the submerged part of the wire becomes coated with a silvery metal. In this reaction, the copper replaces the silver in silver nitrate to form copper(II) nitrate. The equation for this reaction is

$$Cu + 2AgNO_3 \longrightarrow 2Ag + Cu(NO_3)_2$$

Notice that one of the products is silver, which you can see adhering to the wire in Figure 12. The other product, copper(II) nitrate, gives the solution its blue color.

Recall that alkali metals are very reactive elements. Figure 12 shows potassium reacting with water. This is another example of a single-replacement reaction, as the element potassium replaces hydrogen in water to form potassium hydroxide, KOH.

$$2K + 2H_2O \longrightarrow H_2 + 2KOH$$

The heat produced by this chemical reaction causes the hydrogen gas to ignite explosively.

Reading Checkpoint *What is a single-replacement reaction?*

Double Replacement A **double-replacement reaction** is one in which two different compounds exchange positive ions and form two new compounds. The general form of a double replacement reaction is

$$AB + CD \longrightarrow AD + CB$$

Notice that two replacements take place in this reaction. Not only is A replacing C, but C is also replacing A.

Solutions of lead(II) nitrate, $Pb(NO_3)_2$, and potassium iodide, KI, are both colorless. However, when these solutions are mixed, as shown in Figure 13, a yellow precipitate forms as a result of a double-replacement reaction. The equation for this reaction is

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

The lead ions in $Pb(NO_3)_2$, trade places with the potassium ions in KI. The products are lead(II) iodide, PbI_2, which precipitates out of solution, and potassium nitrate, KNO_3, which remains in solution.

When geologists test the calcium carbonate content in a rock, they make use of the following double-replacement reaction.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2CO_3$$

One of the products of this reaction is calcium chloride, $CaCl_2$. The other product is carbonic acid, H_2CO_3, which decomposes into carbon dioxide gas and water.

$$H_2CO_3 \longrightarrow CO_2 + H_2O$$

Figure 13 When potassium iodide solution is poured into a solution of lead(II) nitrate, a double-replacement reaction takes place. Lead(II) iodide forms as a yellow precipitate. **Comparing and Contrasting** ***How does a double-replacement reaction differ from a single-replacement reaction?***

Quick Lab

Identifying a Type of Reaction

Materials

piece of zinc, copper(II) sulfate ($CuSO_4$) solution, 250-mL beaker, tongs, paper towel

Procedure

1. Place the zinc in the beaker and add enough $CuSO_4$ solution to cover the zinc as shown. **CAUTION** *Be careful when using chemicals. Copper sulfate is toxic.*
2. After 5 minutes, carefully remove the zinc from the solution using the tongs and place the zinc on the paper towel to dry. Observe any changes that have occurred to the zinc and the solution of $CuSO_4$. **CAUTION** *Follow your teacher's instructions for disposal of used chemicals. Wash your hands with soap or detergent before leaving the laboratory.*

Analyze and Conclude

1. **Observing** What clues indicate that a chemical reaction has taken place?
2. **Applying Concepts** What were the reactants in this reaction? What were the products? Write a balanced chemical equation for the reaction.
3. **Classifying** Is this a single-replacement or double-replacement reaction? Explain.

Figure 14 A Bunsen burner generates heat and light by the combustion of natural gas. **Interpreting Photos** ***What reactants or products are visible in the reaction shown above?***

Combustion A **combustion reaction** is one in which a substance reacts rapidly with oxygen, often producing heat and light. The burning of natural gas, shown in Figure 14, is an example of combustion. The main component of natural gas is methane, CH_4. When methane burns in an unlimited supply of oxygen, the following reaction occurs.

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

The products of the reaction are carbon dioxide and water. The combustion of methane also generates both heat and light.

By now you know the chemical equation for the combustion of hydrogen.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Notice that you could also classify this reaction as the synthesis of water. The classifications for chemical reactions sometimes overlap.

Reactions as Electron Transfers

So far, you have learned that chemical reactions can be identified by the type of reactant or by the number of reactants and products. For example, in a combustion reaction one of the reactants must be oxygen. In a synthesis reaction, two or more reactants combine to form a single product.

As scientists learned more about the structure of the atom, they found different ways to describe how reactions take place. **The discovery of subatomic particles enabled scientists to classify certain chemical reactions as transfers of electrons between atoms.** A reaction in which electrons are transferred from one reactant to another is called an **oxidation-reduction reaction,** or redox reaction.

Oxidation For a long time, people have known that metals react with oxygen. Calcium, for instance, reacts with oxygen and forms calcium oxide (CaO), shown in Figure 15. Iron reacts with oxygen and forms rust, or iron(III) oxide (Fe_2O_3). These types of synthesis reactions, in which a metal combines with oxygen, traditionally have been classified as oxidations.

When calcium reacts with oxygen, the following reaction takes place.

$$2Ca + O_2 \longrightarrow 2CaO$$

Notice that while the atoms of both reactants (Ca and O_2) are neutral, the product of the reaction is a compound composed of ions (Ca^{2+} and O^{2-}). When calcium reacts with oxygen, each neutral calcium atom loses two electrons and becomes a calcium ion with a charge of 2+.

$$Ca \longrightarrow Ca^{2+} + 2e^-$$

Figure 15 Calcium oxide, or lime, is produced when calcium burns in the presence of oxygen. In this reaction, the calcium is oxidized and the oxygen is reduced.

Any process in which an element loses electrons during a chemical reaction is called oxidation. A reactant is oxidized if it loses electrons. Note that the modern definition of oxidation is much broader than the original meaning. Oxygen doesn't always have to be present in order for an element to lose electrons. For example, when sodium reacts with chlorine, each neutral sodium atom loses one electron and becomes a sodium ion, Na^+.

Reduction As calcium atoms lose electrons during the synthesis of calcium oxide, the oxygen atoms gain electrons. As each neutral oxygen atom gains two electrons, it becomes an ion with a charge of 2−.

$$O + 2e^- \longrightarrow O^{2-}$$

The process in which an element gains electrons during a chemical reaction is called reduction. A reactant is said to be reduced if it gains electrons.

Oxidation and reduction always occur together. When one element loses electrons, another element must gain electrons. Note that oxidation-reduction reactions do not always involve complete transfers of electrons. For example, in the synthesis of water, hydrogen is oxidized as it partially loses electrons. Oxygen is reduced as it partially gains electrons.

Go Online NSTA SCILINKS

For: Links on oxidation and reduction
Visit: www.SciLinks.org
Web Code: ccn-1072

Section 7.2 Assessment

Reviewing Concepts

1. What are five general types of reactions?
2. How did the discovery of subatomic particles affect the classification of reactions?
3. The synthesis of water is described by the equation $2H_2 + O_2 \longrightarrow 2H_2O$. How is the decomposition of water related to this reaction? Explain, using a chemical equation.
4. Explain the difference between a single-replacement reaction and a double-replacement reaction.
5. Propane, C_3H_8, is frequently used in camping stoves. When propane undergoes combustion, what are the products formed?
6. Is the reaction represented by the following equation a redox reaction? Explain your answer.

$$2Hg + O_2 \longrightarrow 2HgO$$

Critical Thinking

7. **Predicting** What is the product of the synthesis reaction between magnesium and iodine? Explain your answer.
8. **Classifying** Identify these reactions as synthesis, decomposition, single replacement, double replacement, or combustion.
   a. $Pb(NO_3)_2 + 2HCl \longrightarrow PbCl_2 + 2HNO_3$
   b. $2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$
   c. $Ca + 2HCl \longrightarrow CaCl_2 + H_2$
   d. $2SO_2 + O_2 \longrightarrow 2SO_3$

Writing in Science

Explanatory Paragraph Write a paragraph explaining why the formation of water can be classified as a synthesis, combustion, or oxidation-reduction reaction.

7.3 Energy Changes in Reactions

Reading Focus

Key Concepts

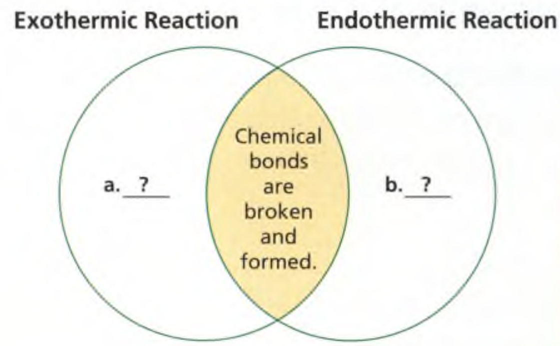
- What happens to chemical bonds during a chemical reaction?
- What happens to energy during a chemical reaction?

Vocabulary

- chemical energy
- exothermic reaction
- endothermic reaction

Reading Strategy

Comparing and Contrasting Copy the Venn diagram. As you read, complete it to show the differences between exothermic and endothermic reactions.



If you've ever had a barbecue, you may have used a gas grill like the one shown in Figure 16. Many types of gas grills use propane, C_3H_8. You can think of a propane grill as the scene of a chemical reaction—specifically, a combustion reaction. The reactants are propane and oxygen, and the products are carbon dioxide and water. However, the description of this reaction is incomplete unless you consider the heat and light produced. Heat, after all, is the reason for using a propane grill.

Chemical Bonds and Energy

The heat produced by a propane grill is a form of energy. When you write the chemical equation for the combustion of propane, you can include "heat" on the right side of the equation.

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O + \text{Heat}$$

This equation states that the heat released in the reaction came from the reactants. **Chemical energy** is the energy stored in the chemical bonds of a substance. A propane molecule has ten single covalent bonds (eight C—H bonds and two C—C bonds). The chemical energy of a propane molecule is the energy stored in these bonds. Likewise, oxygen, carbon dioxide, and water molecules all have energy stored in their chemical bonds.

Figure 16 Many portable barbecue grills burn propane gas.

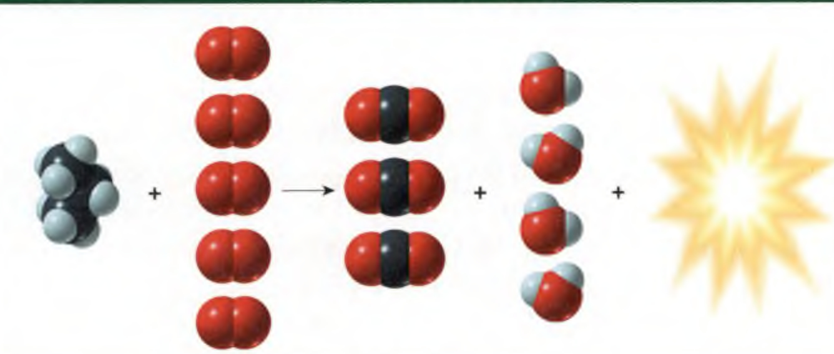
Energy changes in chemical reactions are determined by changes that occur in chemical bonding. **Chemical reactions involve the breaking of chemical bonds in the reactants and the formation of chemical bonds in the products.** In the combustion of propane, the bonds in propane and oxygen molecules are broken, while the bonds in carbon dioxide and water molecules are formed.

Breaking Bonds As Figure 17 illustrates, each propane molecule reacts with five oxygen molecules. In order for the reaction to occur, eight C—H single bonds, two C—C single bonds, and five O=O double bonds must be broken. Breaking chemical bonds requires energy. This is why propane grills have an igniter, a device that produces a spark. The spark provides enough energy to break the bonds of reacting molecules and get the reaction started.

Forming Bonds Figure 17 also shows you that for each molecule of propane burned, three molecules of carbon dioxide and four molecules of water are formed. This means that six C=O double bonds and eight O—H single bonds are formed in the reaction. The formation of chemical bonds releases energy. The heat and light given off by a propane stove result from the formation of new chemical bonds. The bonds form as the carbon, hydrogen, and oxygen atoms in the propane and oxygen molecules are rearranged into molecules of carbon dioxide and water.

Reading Checkpoint *Does breaking chemical bonds require energy or release energy?*

Propane Combustion



C_3H_8 *propane*	+	$5O_2$ *oxygen*	→	$3CO_2$ *carbon dioxide*	+	$4H_2O$ *water*	+	Heat
8 C–H, 2 C–C *bonds broken*		5 O=O *bonds broken*	→	6 C=O *bonds formed*		8 O–H *bonds formed*		

Figure 17 In order for the combustion of propane to occur, all the chemical bonds in the reactants (propane and oxygen) must be broken. The formation of the chemical bonds in the products completes the reaction. **Inferring** ***How does the chemical energy of the reactants compare to the chemical energy of the products in this reaction?***

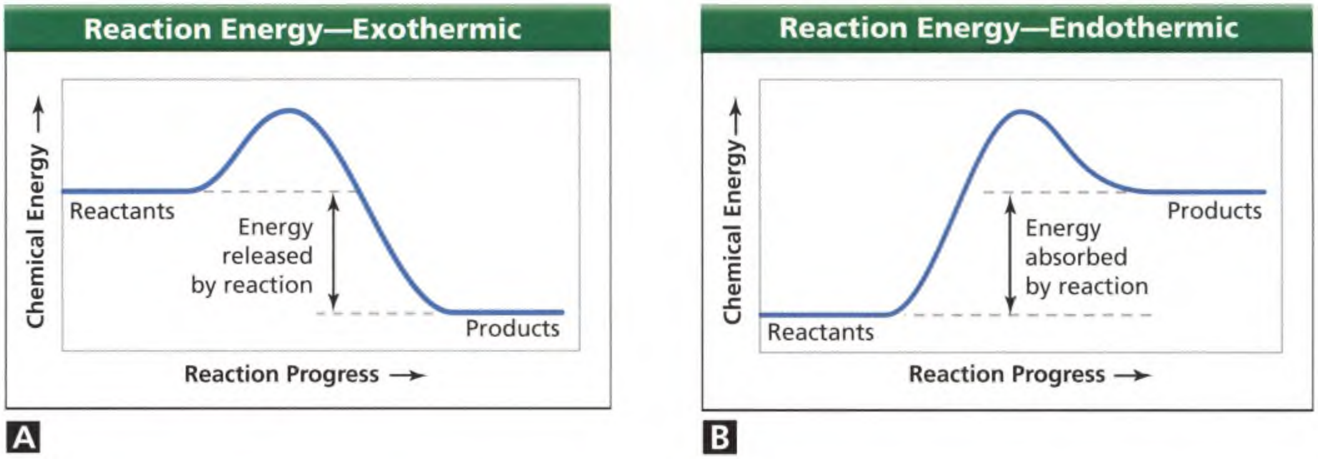


Figure 18 In chemical reactions, energy is either released or absorbed. **A** In an exothermic reaction, energy is released to the surroundings. **B** In an endothermic reaction, energy is absorbed from the surroundings. **Using Graphs** ***How do the energy diagrams show that energy is conserved in chemical reactions?***

Exothermic and Endothermic Reactions

Recall that physical changes can release or absorb energy. During an exothermic change, such as freezing, energy is released to the surroundings. During an endothermic change, such as melting, energy is absorbed from the surroundings. Energy also flows into and out of chemical changes. **During a chemical reaction, energy is either released or absorbed.**

Exothermic Reactions A chemical reaction that releases energy to its surroundings is called an **exothermic reaction.** In exothermic reactions, the energy released as the products form is greater than the energy required to break the bonds in the reactants.

Combustion is an example of an extremely exothermic reaction. When 1 mole of propane reacts with 5 moles of oxygen, 2220 kJ (kilojoules) of heat is released. You can use this value to replace "heat" in the combustion equation written earlier.

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O + 2220\text{ kJ}$$

Figure 18A shows how chemical energy changes during an exothermic reaction. Notice that the chemical energy of the reactants is greater than the chemical energy of the products. The difference between these amounts of energy equals the amount of heat given off by the reaction.

In any reaction, the chemical energy reaches a peak before the reactants change into products. This peak represents the amount of energy required to break the chemical bonds of the reactants. Unless reacting particles collide with enough energy to break these bonds, the reaction will not occur. For example, at room temperature, the collisions between propane and oxygen molecules are not energetic enough to result in combustion. However, if you increase the temperature by adding a spark, some of the molecules around the spark move faster and are able to collide with enough energy to react.

Endothermic Reactions A chemical reaction that absorbs energy from its surroundings is called an **endothermic reaction.** In an endothermic reaction, more energy is required to break the bonds in the reactants than is released by the formation of the products.

Figure 18B shows the energy diagram for an endothermic reaction. Notice that the energy of the products is greater than the energy of the reactants. The difference between these amounts of energy equals the amount of heat that must be absorbed from the surroundings.

When mercury(II) oxide is heated to a temperature of about 450°C, it breaks down into mercury and oxygen, as shown in Figure 19. The decomposition of mercury(II) oxide is an endothermic reaction that can be described by the following equation.

$$2HgO + 181.7\ kJ \longrightarrow 2Hg + O_2$$

Because heat is absorbed, the energy term appears on the left side of the equation. For every 2 moles of HgO that decomposes, 181.7 kJ of heat must be absorbed.

Figure 19 The orange-red powder in the bottom of the test tube is mercury(II) oxide. At about 450°C, mercury(II) oxide decomposes into oxygen gas (which escapes from the test tube) and mercury (droplets of which can be seen collecting on the sides of the test tube).

Conservation of Energy

In an exothermic reaction, the chemical energy of the reactants is converted into heat plus the chemical energy of the products. In an endothermic reaction, heat plus the chemical energy of the reactants is converted into the chemical energy of the products. In both cases, the total amount of energy before and after the reaction is the same. This principle is known as the law of conservation of energy. You will read more about how energy is conserved later.

Section 7.3 Assessment

Reviewing Concepts

1. What happens to chemical bonds as a chemical reaction occurs?
2. How do chemical reactions involve energy?
3. Is the combustion of propane endothermic or exothermic?
4. When propane reacts with oxygen, does the surrounding area become warmer or cooler?
5. Is energy created during an exothermic reaction? Explain.

Critical Thinking

6. **Inferring** Explain why methane does not react with oxygen at room temperature.
7. **Calculating** Methane reacts with oxygen in the following combustion reaction.

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

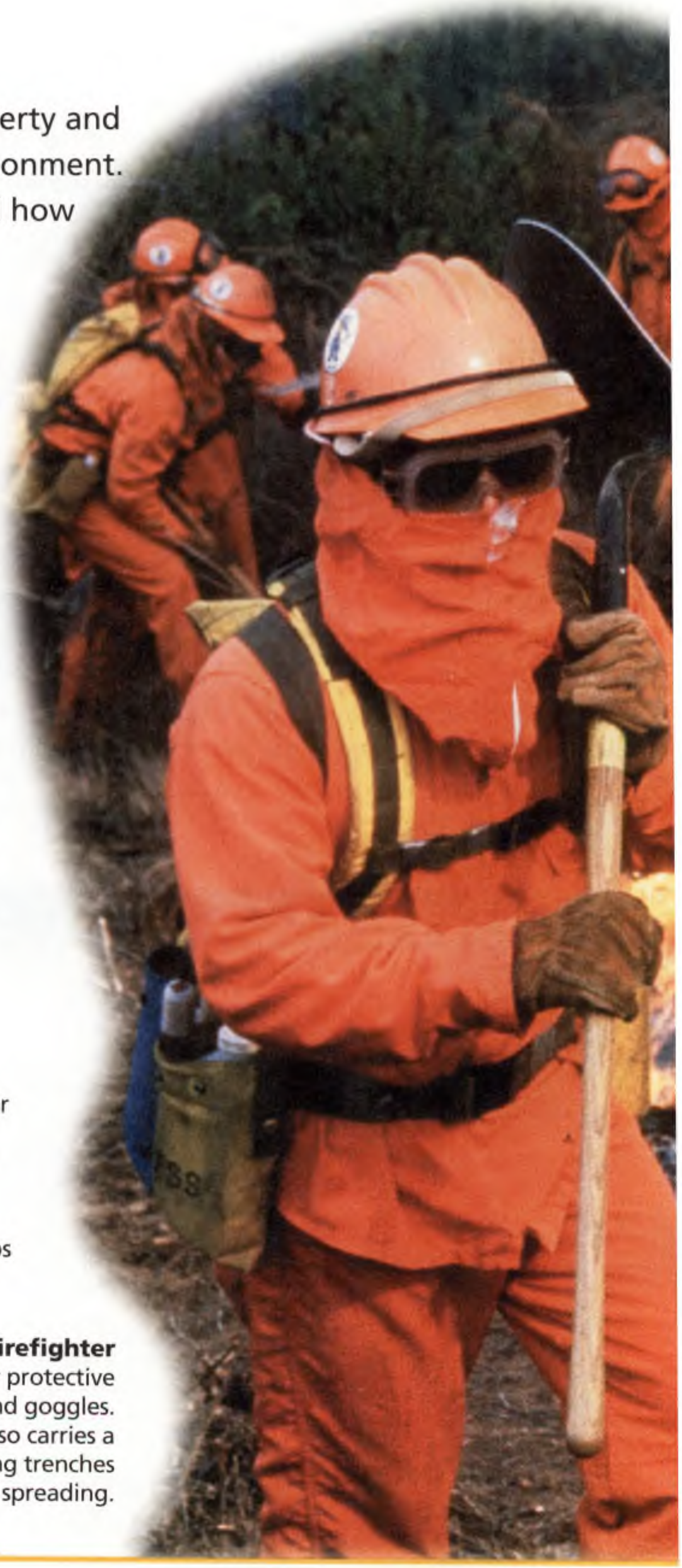
What bonds are broken when one molecule of methane reacts with two molecules of oxygen?

Connecting Concepts

Chemical Bonds Reread the descriptions of chemical bonds in Sections 6.1 and 6.2. Then, describe the decomposition of mercury(II) oxide. Specify which bonds are ionic and which bonds are covalent.

Firefighting

Uncontrolled fires threaten lives and property and can have a devastating effect on the environment. To fight fires, it is necessary to understand how they start, and what sustains them.



Fire is the result of combustion, a rapid reaction between oxygen and fuel. During combustion, fuel and oxygen react to form carbon dioxide, water, and heat. Most fires in rural areas, called wildfires, are caused by people being careless with campfires or cigarettes. Arson and lightning are also common causes. Usually the fire starts in an area of dry grass, which will ignite at a temperature of 150–200°C. Burning grass can create enough heat to ignite bushes, and these, in turn, may be tall enough to carry the flames into trees. (Wood has a higher ignition temperature, around 260°C.) Environmental conditions, such as drought that has left vegetation dry, and strong winds, can cause a small fire to spread. Wind also carries fire forward into new areas. Burning twigs and branches that become detached from trees can be blown into new areas of vegetation.

Dousing
Planes or helicopters are used to drop water or a fire-retardant slurry on a large fire. The slurry contains ammonium sulfate, $(NH_4)_2SO_4$, which helps smother the fire.

Firefighter
Firefighters wear protective clothing, helmets, and goggles. This firefighter also carries a shovel for digging trenches to stop a fire spreading.

Backburning
One way to stop a big fire is to start a smaller fire. This technique, called backburning, burns off vegetation between the fireline and the main fire. When the two fires meet, the blaze stops spreading because the land on either side has already been burned.

Firebreaks
In parts of the world where wildfires are common, strips of land are cleared of combustible vegetation to create firebreaks. Firefighters also dig firebreaks during fires to prevent fires from spreading.

Wildfire in Washington State
This satellite picture shows a forest fire that raged for two months in 2000 at the Department of Energy's site at Hanford. The fire was started by a vehicle fire and reached more than 500 square kilometers.

Going Further

- Write a paragraph explaining how dousing, backburning, and firebreaks affect the chemical reactions involved in wildfires.
- Take a Discovery Channel Video Field Trip by watching "Taming the Flames."

Discovery Channel School
Video Field Trip

7.4 Reaction Rates

Reading Focus

Key Concept

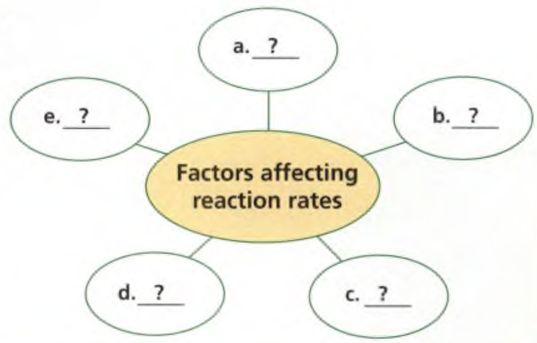
- What does a reaction rate tell you?
- What factors cause reaction rates to change?

Vocabulary

- reaction rate
- catalyst

Reading Strategy

Building Vocabulary Copy the partially completed web diagram at the right. Then, as you read, complete it with key terms from this section.



You may have heard of athletes "burning Calories" when they exercise. A Calorie is a unit of energy used in the field of nutrition. The average banana, for instance, contains about 100 Calories. The cyclist in Figure 20 can use up, or burn, as many as 10,000 Calories during the course of a race. That adds up to a lot of bananas!

If you eat a banana, you provide your body with about 100 Calories to burn. This energy is released in a series of reactions that take place inside your body. A much faster way of releasing the energy contained in a banana is to burn it—outside the body—in a combustion reaction. In both cases, the total amount of energy released is the same. However, the time it takes for the energy to be released is different in each case.

Figure 20 A cyclist burns the Calories in a banana faster than a person walking would. But burning the banana outside the body would release the energy of the banana even faster.

Reactions Over Time

The progress of any chemical reaction can be measured over time. Different reactions have different durations. Some reactions, such as the explosion of TNT, happen almost instantaneously. Other reactions, such as tree leaves changing color during autumn, happen gradually.

Any change that happens over a period of time can be expressed as a rate. For example, speed is the rate that distance changes over time. A **reaction rate** is the rate at which reactants change into products over time. **Reaction rates tell you how fast a reaction is going.** That is, how fast the reactants are being consumed, how fast the products are being formed, or how fast energy is being absorbed or released.

Factors Affecting Reaction Rates

Recall that chemical reactions involve collisions between particles of reactants. The reaction rate depends on how often these particles collide. If the collisions occur more frequently, then the reaction rate increases. If the collisions occur less frequently, then the reaction rate decreases. Almost any reaction rate can be changed by varying the conditions under which the reaction takes place. **Factors that affect reaction rates include temperature, surface area, concentration, stirring, and catalysts.**

Go Online NSTA SCILINKS

For: Links on factors affecting reaction rate
Visit: www.SciLinks.org
Web Code: ccn-1074

Temperature Suppose you are frying an egg in a frying pan. What happens if you increase the heat under the pan? The hotter the pan, the faster the egg will cook. Generally, an increase in temperature will increase the reaction rate, while a decrease in temperature will decrease the reaction rate. For instance, you store milk in a refrigerator to slow down the reactions that cause the milk to spoil. These reactions don't stop completely. Even milk stored in a refrigerator will eventually spoil. But the rate of spoiling decreases if the milk is kept cold.

Increasing the temperature of a substance causes its particles to move faster, on average. Particles that move faster are both more likely to collide and more likely to react. If the number of collisions that produce reactions increases, then the reaction rate increases.

Reading Checkpoint *How does temperature affect reaction rates?*

Surface Area Grain may not strike you as a dangerous material, but it can be explosive under the right conditions. The cause of the fire in Figure 21 was a combustion reaction between grain dust (suspended in the air) and oxygen. The rate of combustion was very rapid due to the small particle size of the grain dust.

The smaller the particle size of a given mass, the larger is its surface area. Imagine using a newspaper to cover the floor of a room. If you keep all the sections folded together, you can only cover a small area. However, if you separate the newspaper into pages and lay them out like tiles, you can cover a much larger area with the same mass of paper.

An increase in surface area increases the exposure of reactants to one another. The greater this exposure, the more collisions there are that involve reacting particles. With more collisions, more particles will react. This is why increasing the surface area of a reactant tends to increase the reaction rate.

Figure 21 This grain elevator in Potlatch, Idaho, exploded when grain dust reacted with oxygen in the air.
Applying Concepts *How does surface area affect reaction rates?*

Quick Lab

Observing the Action of Catalysts

Materials

5 test tubes, test-tube rack, marking pencil, dropper pipet, wood splint, platinum wire, 0.1 g manganese dioxide (MnO_2), 5 drops of copper(II) chloride ($CuCl_2$) solution, 0.1 g raw potato, graduated cylinder, 25 mL hydrogen peroxide (H_2O_2)

Procedure

1. Label the 5 test tubes from A to E with the marking pencil.
2. Put a small piece of platinum wire in test tube A. Add a tiny amount (about the tip of the wood splint) of MnO_2 to test tube B. Use the dropper pipet to put 5 drops of $CuCl_2$ in test tube C. Put a piece of potato in test tube D. Test tube E should remain empty for now. **CAUTION** *MnO_2 and $CuCl_2$ are toxic.*
3. Carefully add 5 mL of hydrogen peroxide to test tube A. **CAUTION** *Be careful when using chemicals.* Observe how fast the bubbles are produced.
4. Repeat Step 3 with test tubes B through E.

Analyze and Conclude

1. **Observing** What effect did the platinum wire, MnO_2, $CuCl_2$, and the potato have on the rate at which the bubbles were produced in the hydrogen peroxide?
2. **Comparing and Contrasting** Which catalyst(s) caused the reaction to go the fastest? The slowest?
3. **Inferring** Why did you put only hydrogen peroxide in test tube E?

Stirring You can also increase the exposure of reactants to each other by stirring them. For example, when you wash your clothes in a washing machine, particles of detergent react with particles of the stains on your clothes. This reaction would go slowly if you just left your clothes soaking in a tub of water and detergent. A washing machine speeds up the reaction by stirring the contents back and forth. Collisions between the particles of the reactants are more likely to happen. Stirring the reactants will generally increase the reaction rate.

Figure 22 The dye solution in the left beaker is more concentrated than the solution in the right. Increasing the concentration of the dye increases the rate of color change in the material.

Concentration Another way you can change the reaction rate is to change the concentration of the reactants. Concentration refers to the number of particles in a given volume. The more reacting particles that are present in a given volume, the more opportunities there are for collisions involving those particles. The reaction rate is faster.

Both of the beakers in Figure 22 contain a piece of material dipped in dye solution. Dyeing is a chemical reaction in which dye particles react with the particles of the material being dyed. The material dipped in the more concentrated dye becomes colored more quickly.

For gases, concentration changes with pressure. The greater the pressure of a gaseous reactant, the greater is its concentration, and the faster is the reaction rate.

Catalysts Sometimes you can change a reaction rate by using catalysts. A **catalyst** is a substance that affects the reaction rate without being used up in the reaction. Chemists often use catalysts to speed up a reaction or enable a reaction to occur at a lower temperature. In the making of sulfuric acid, one of the steps involved is the reaction of sulfur dioxide with oxygen to form sulfur trioxide. This reaction happens very slowly without a catalyst such as vanadium(V) oxide.

$$2SO_2 + O_2 \xrightarrow{V_2O_5} 2SO_3$$

Since the catalyst is neither a reactant nor a product, it is written over the arrow. Because the catalyst is not consumed, it can be used to speed up the same reaction over and over again.

Recall that in order for a reaction to take place, the reacting particles must collide with enough energy to break the chemical bonds of those particles. As shown in Figure 23, a catalyst lowers this energy barrier. One way that a catalyst can do this is by providing a surface on which the reacting particles can come together. Imagine that you go to a party and make several new friends. By bringing people together, the party has made it easier for you to form those friendships. Similarly, a catalyst can "invite" reacting particles together so that they are more likely to react.

Reaction Energy—Catalyzed and Uncatalyzed

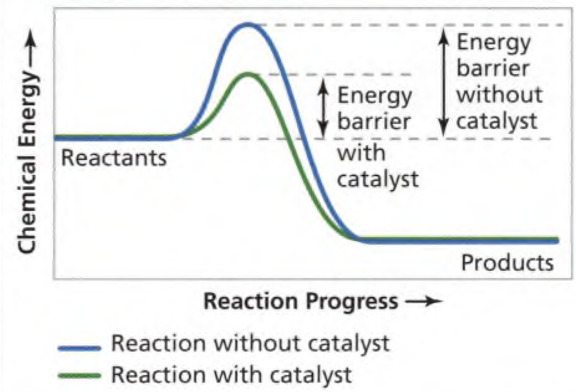


Figure 23 The graph above shows how a catalyst lowers the amount of energy required for effective collisions between reacting particles.
Using Graphs ***In an exothermic reaction, how does a catalyst affect the amount of energy released?***

Section 7.4 Assessment

Reviewing Concepts

1. What does a reaction rate tell you?
2. What five factors affect reaction rates?
3. Explain why reactions take place faster at higher temperatures.
4. When you add baking soda to vinegar, the mixture fizzes as carbon dioxide gas is produced. Suppose you added water to the vinegar before you mixed it with the baking soda. What do you think would happen to the rate of carbon dioxide production?
5. How does a catalyst make a reaction go faster?
6. Platinum is a catalyst for the decomposition of hydrogen peroxide into water and oxygen.

$$2H_2O_2 \xrightarrow{Pt} 2H_2O + O_2$$

What would you expect to see if platinum were added to hydrogen peroxide solution?

Critical Thinking

7. **Applying Concepts** Explain why, if you want to store uncooked hamburger meat for a month, you put it in a freezer rather than a refrigerator.
8. **Evaluating** The reaction between magnesium and hydrochloric acid produces hydrogen. If you increase the concentration of HCl, the reaction takes place faster. Could HCl be considered a catalyst for this reaction? Explain your answer.

Writing in Science

Compare and Contrast Paragraph Write a paragraph explaining how temperature, concentration, surface area, and catalysts affect reaction rates.

7.5 Equilibrium

Reading Focus

Key Concepts

- Under what conditions do physical and chemical equilibria occur?
- How do equilibrium systems respond to change?

Vocabulary

- equilibrium
- reversible reaction

Reading Strategy

Outlining As you read, make an outline of the most important ideas in this section.

I. Equilibrium
  A. Types of Equilibria
    1. ____________
    2. ____________
  B. ____________
    1. Temperature
    2. Pressure
    3. ____________

Suppose you're waiting in line for a toll booth at a bridge, like some of the cars shown in Figure 24. You notice that every time a car passes by a toll booth in the direction you are traveling, another car passes through the toll plaza in the opposite direction. The rate of cars entering equals the rate of cars exiting. As a result, the number of cars on either side of the toll plaza remains constant, although cars are continually entering and exiting the bridge.

Types of Equilibria

The traffic at a toll bridge is similar to a system in equilibrium. **Equilibrium** (plural *equilibria*) is a state in which the forward and reverse paths of a change take place at the same rate.

Recall that changes to matter are either physical or chemical. When opposing physical changes take place at the same rate, a physical equilibrium is reached. When opposing chemical changes take place at the same rate, a chemical equilibrium is reached.

Figure 24 About 190,000 vehicles pass through the toll plaza of New York City's Verrazano-Narrows Bridge every day.

Physical Equilibrium What happens when you pour some water into a jar and then close the lid? You might think that nothing happens at all. But in fact, some of the water undergoes a physical change by evaporating. As more water evaporates, some of the water vapor condenses. Eventually, the rate of evaporation equals the rate of condensation, and the system reaches equilibrium as shown in Figure 25.

When liquid water is in equilibrium with water vapor, you can describe the system by writing this equation.

$$H_2O(l) \rightleftharpoons H_2O(g)$$

Here, *l* stands for liquid and *g* stands for gas. The pair of arrows in this equation indicates that the forward change (evaporation) and the reverse change (condensation) are happening simultaneously and at the same rate. Both the forward and reverse changes are physical changes, so this equation represents a physical equilibrium. **When a physical change does not go to completion, a physical equilibrium is established between the forward and reverse changes.**

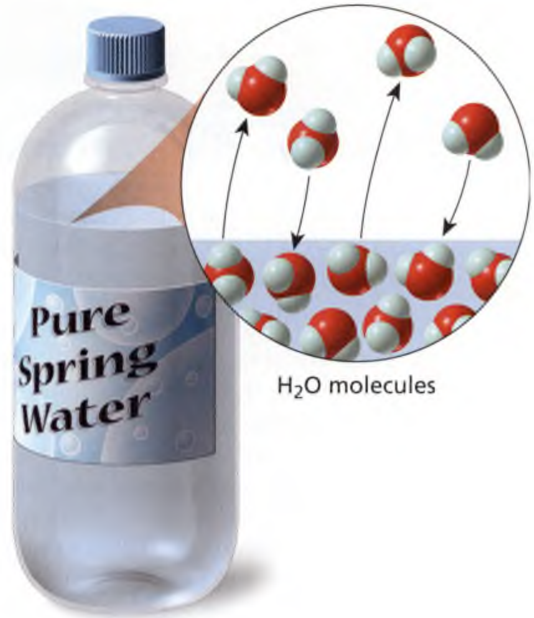


Figure 25 Liquid water left in a closed container eventually reaches equilibrium with its vapor. **Interpreting Diagrams** ***What do the arrows represent in the diagram above?***

Chemical Equilibrium All the chemical equations you have seen so far have been written with single arrows, which suggest that all reactions go to completion in one direction. In reality, however, most reactions are reversible to some extent. A **reversible reaction** is a reaction in which the conversion of reactants into products and the conversion of products into reactants can happen simultaneously.

In the previous section, you read about the synthesis of sulfur trioxide from sulfur dioxide and oxygen. This is actually a reversible reaction that can be expressed as

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

If sulfur dioxide and oxygen are mixed in a closed container, the forward reaction will start to produce sulfur trioxide. However, once molecules of sulfur trioxide form, some of them will change back into the reactants by the reverse reaction. Eventually, the rate of the forward reaction (synthesis) will equal the rate of the reverse reaction (decomposition), and the system will reach equilibrium. **When a chemical reaction does not go to completion, a chemical equilibrium is established between the forward and reverse reactions.** During chemical equilibrium, the reactants change into products just as fast as the products change back into reactants.

Reading Checkpoint ***What happens during chemical equilibrium?***

Go Online
NSTA SCILINKS

For: Links on factors affecting equilibrium
Visit: www.SciLinks.org
Web Code: ccn-1075

Problem-Solving Activity

Recreating High Altitudes

An important chemical equilibrium in your blood involves the reaction of hemoglobin (Hb) with oxygen (O_2) to form oxyhemoglobin (HbO_2).

$$Hb + O_2 \rightleftharpoons HbO_2$$

This equilibrium changes with altitude. As you move from lower to higher elevations, the concentration of oxygen in the air decreases, and the equilibrium shifts in the direction that produces less oxyhemoglobin. Your body responds to the shift by producing more hemoglobin. Studies have shown that athletes can improve their performance at sea level by living or training at high altitudes. Some training facilities are designed to recreate high altitudes. Imagine that you are asked to build such a facility.

Defining the Problem In your own words, state the problem you face.

Organizing Information Use Le Châtelier's principle to determine how high altitudes affect this equilibrium system.

Creating a Solution The physical properties of the air inside the training facility include temperature, pressure, and composition. Figure out how to shift the equilibrium in the direction you want by changing one of these properties.

Presenting Your Plan Write a proposal to an athletic team that could benefit from using your training facility. Explain how your facility recreates a high-altitude environment.

Factors Affecting Chemical Equilibrium

Like reaction rates, chemical equilibria can change depending on the conditions of the reaction. While a reaction rate either increases or decreases in response to a change, an equilibrium shifts. That is, the equilibrium favors either the forward or the reverse reaction. **When a change is introduced to a system in equilibrium, the equilibrium shifts in the direction that relieves the change.** This rule was first observed by Henri Le Châtelier, shown in Figure 26. Today, the rule is known as Le Châtelier's principle.

The making of ammonia is an example of a process in which chemists apply Le Châtelier's principle. Ammonia is an important industrial chemical used to make fertilizers, cleaning agents, dyes, and plastics. The following equation describes the synthesis of ammonia.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{Heat}$$

Suppose you have a system that contains nitrogen, hydrogen, and ammonia in equilibrium. By applying Le Châtelier's principle, you can predict how this system will be affected by changes in temperature, pressure, and concentration. In the ammonia plant shown in Figure 27, chemists must consider these same factors.

Temperature In the equation for the synthesis of ammonia, heat is written as a product. This tells you that the forward reaction is exothermic. In the reverse reaction, heat is a reactant. So the decomposition of ammonia is endothermic.

Figure 26 French Chemist Henri-Louis Le Châtelier (1850–1936) published the first version of his principle of chemical equilibrium in 1884.

What would happen if you increased the temperature of a system that contained nitrogen, hydrogen, and ammonia? According to Le Châtelier's principle, if you added heat to the system, the equilibrium would shift in the direction that removes heat from the system. The system would favor the reverse reaction, which is endothermic. So by increasing the temperature, you would decrease the amount of ammonia.

Pressure Suppose you increased the pressure of the system. According to Le Châtelier's principle, if you increased the pressure, the equilibrium would shift in the direction that decreases the pressure of the system. In order to decrease pressure, the system would favor the reaction that produces fewer gas molecules. You can see that the left side of the equation has four gas molecules, while the right side has two. So by increasing the pressure, you would shift the equilibrium to the right, producing more ammonia.

Concentration A change in concentration of the reactants or products can also affect equilibrium. Suppose you removed ammonia from the nitrogen-hydrogen-ammonia system. Le Châtelier's principle tells you that the equilibrium would shift in the direction that produces ammonia. In order to produce ammonia, the system would favor the forward reaction.

Figure 27 Operating an ammonia plant at relatively low temperature, high pressure, and low ammonia concentration maximizes the amount of ammonia produced.

Section 7.5 Assessment

Reviewing Concepts

1. What happens when a physical change does not go to completion? What happens when a reaction does not go to completion?
2. Once a chemical reaction has reached equilibrium, how does the system respond to change?
3. What does the double-ended arrow indicate in the following chemical equation?

$$CaCO_3 \rightleftharpoons Ca^{2+} + CO_3^{2-}$$

4. For which of the following reactions are both reactants and products likely to be found when the reaction appears to be complete? Explain.

$$Mg + O_2 \longrightarrow 2MgO$$

$$HF + H_2O \rightleftharpoons H_3O^+ + F^-$$

Critical Thinking

5. **Inferring** Suppose the following reaction is allowed to come to equilibrium.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

How will increasing the pressure on this system affect the amount of N_2O_4 formed?

6. **Using Models** At 0°C, liquid water is in equilibrium with ice. Make a drawing of water molecules at this temperature, and describe what is happening.

Connecting Concepts

Phase Changes Write an equation for a system in which the sublimation and deposition of water have reached equilibrium. Use what you studied in Section 3.3 to explain what changes are happening.

Design Your Own Lab

Manipulating Chemical Equilibrium

Chemical reactions tend to go to equilibrium. It is possible to shift the equilibrium by changing the conditions under which the reaction occurs. Factors that can affect chemical equilibrium include the concentration of reactants and products, temperature, and pressure. In this lab, you will observe a chemical reaction and use your observations to predict how one factor will shift the equilibrium of the reaction. Then, you will perform an experiment to test your prediction.

Problem

How can you change the equilibrium of a chemical reaction?

Materials

- iodine-starch solution
- 150-mL beaker
- 4 dropper pipets
- spot plate
- ascorbic acid (vitamin C) solution
- chlorine bleach (sodium hypochlorite, NaOCl) solution

Skills

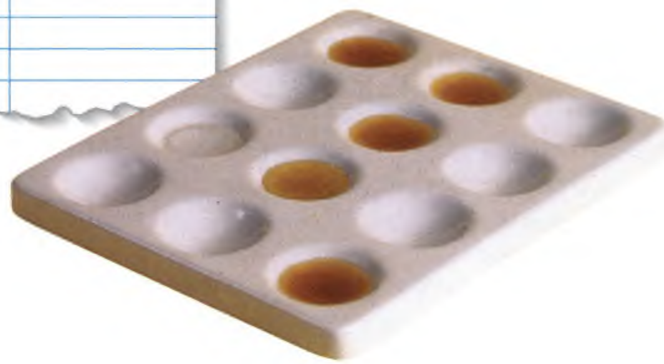
Formulating Hypotheses, Designing Experiments, Observing

Procedure

Part A: Observing a Reversible Reaction

1. Pour 50 mL of iodine-starch solution into the 150-mL beaker. The dark color of this solution is due to the presence of iodine molecules (I_2) within the grains of starch. **CAUTION** *Handle iodine solutions with care. Iodine is toxic.*
2. Use a dropper pipet to transfer 3 drops of iodine-starch solution from the beaker to one well on the spot plate.
3. Use another clean dropper pipet to add one drop of ascorbic acid solution to the iodine-starch solution on the spot plate. Continue to add ascorbic acid solution to the mixture on the spot plate, one drop at a time, until the mixture becomes clear. When an iodine molecule reacts with ascorbic acid, the iodine molecule is reduced and breaks down into two colorless iodide ions ($2I^-$).

Sample Data Table			
Initial Solution	Solution Added	Quantity Added (mL)	Color of Resulting Mixture



4. Use the third clean dropper pipet to transfer one drop of colorless iodide solution to a second well on the spot plate.
5. Use the last clean dropper pipet to add bleach solution to the drop of colorless iodide solution, one drop at a time. Continue until the dark color of the iodine-starch solution reappears. **CAUTION** *Bleach can damage skin and clothing.* The chlorine bleach (NaOCl) oxidizes iodide ions (I^-), converting them to iodine molecules (I_2).
6. Write a chemical equation showing the equilibrium between iodine molecules and iodide ions. This equation does not need to be balanced. Label the two sides of your equation to indicate which substance appears dark and which appears colorless.

Part B: Design Your Experiment

7. **Predicting** Select one of the solutions used earlier that affects the equilibrium between iodine molecules and iodide ions. Record your prediction of the change you will observe in an iodine-starch solution as you add the solution that you selected.
8. **Designing Experiments** Design an experiment to test your prediction. Your experimental plan should describe in detail how you will perform your experiment.
9. Construct a data table like the sample data table shown, in which to record your observations. (*Note:* Your data table may not be exactly like the sample data table.)
10. Perform your experiment only after your teacher has approved your plan. Record your observations in your data table. **CAUTION** *Wash your hands with soap or detergent before leaving the laboratory.*

Analyze and Conclude

1. **Analyzing Data** What factor did you investigate? How did it affect the equilibrium between iodine molecules and iodide ions?
2. **Predicting** How would you expect the equilibrium to change if you added more iodide ions to the mixture? Explain your answer.
3. **Calculating** When chlorine bleach (sodium hypochlorite, NaOCl) oxidizes iodide ions to iodine molecules, sodium hypochlorite is reduced to sodium chloride (NaCl) and water (H_2O). Write a balanced chemical equation for this reaction, beginning with the reactants sodium hypochlorite, iodide ions, and hydrogen ions (H^+).
4. **Drawing Conclusions** How does the addition of more product affect the chemical equilibrium of a reaction?

Go Further Design an experiment to determine whether other substances that are easily oxidized or reduced, such as iron ions, can reduce iodine to iodide, or oxidize iodide to iodine. Then, with your teacher's approval and supervision, perform your experiment.

CHAPTER 7

Study Guide

7.1 Describing Reactions

Key Concepts

- The law of conservation of mass states that mass is neither created nor destroyed.
- In order to show that mass is conserved during a reaction, a chemical equation must be balanced.
- Because chemical reactions often involve large numbers of small particles, chemists use a unit called the mole to measure amounts of a substance.
- In chemical reactions, the mass of a reactant or product can be calculated by using a balanced chemical equation and molar masses.

Vocabulary

reactants, *p. 192;* products, *p. 192;* chemical equation, *p. 193;* coefficients, *p. 194;* mole, *p. 195;* molar mass, *p. 196*

7.2 Types of Reactions

Key Concepts

- The general types of chemical reactions are synthesis reactions, decomposition reactions, single-replacement reactions, double-replacement reactions, and combustion reactions.
- Scientists classify certain chemical reactions as transfers of electrons between atoms.

Vocabulary

synthesis reaction, *p. 200;* decomposition reaction, *p. 200;* single-replacement reaction, *p. 202;* double-replacement reaction, *p. 203;* combustion reaction, *p. 204;* oxidation-reduction reaction, *p. 204*

7.3 Energy Changes in Reactions

Key Concepts

- Chemical reactions involve the breaking of chemical bonds in the reactants and the formation of chemical bonds in the products.
- During a chemical reaction, energy is either released or absorbed.

Vocabulary

chemical energy, *p. 206;* exothermic reaction, *p. 208;* endothermic reaction, *p. 209*

7.4 Reaction Rates

Key Concepts

- Reaction rates tell you how fast a reaction is going.
- Factors that affect reaction rates include temperature, surface area, concentration, stirring, and catalysts.

Vocabulary

reaction rate, *p. 212;* catalyst, *p. 215*

7.5 Equilibrium

Key Concepts

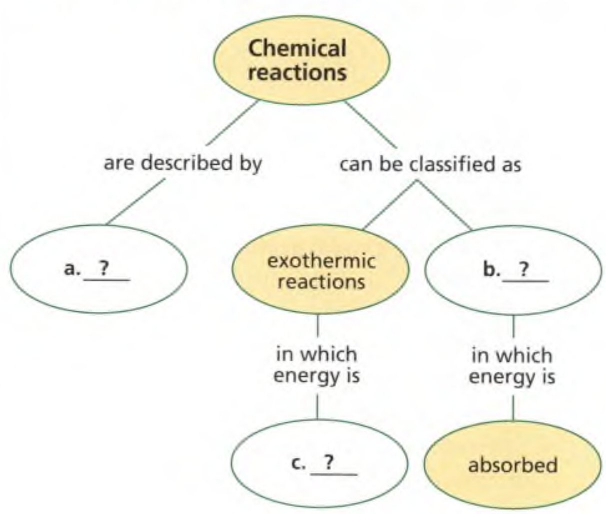
- When a physical change does not go to completion, a physical equilibrium is established between the forward and reverse changes. When a chemical reaction does not go to completion, a chemical equilibrium is established between the forward and reverse reactions.
- When a change is introduced to a system in equilibrium, the equilibrium shifts in the direction that relieves the change.

Vocabulary

equilibrium, *p. 216;* reversible reaction, *p. 217*

Thinking Visually

Concept Map Use information from the chapter to complete the concept map below.



CHAPTER 7

Assessment

Interactive Textbook with assessment at PHSchool.com

Reviewing Content

Choose the letter that best answers the question or completes the statement.

1. In the following equation, what are the reactants?
   $NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$
   a. $NaHCO_3$ and HCl b. $NaHCO_3$ and NaCl
   c. HCl and NaCl d. NaCl, H_2O, and CO_2

2. Which of the following is a statement of the law of conservation of mass?
   a. Mass is created but not destroyed.
   b. Mass is destroyed but not created.
   c. Mass is neither created nor destroyed.
   d. Mass is both created and destroyed, depending on the chemical reaction.

3. The mass of a hydrogen atom is 1.0 amu, and the mass of a carbon atom is 12.0 amu. What is the molar mass of methane, CH_4?
   a. 13.0 amu b. 13.0 g
   c. 16.0 amu d. 16.0 g

4. In what type of reaction does one reactant form two or more products?
   a. synthesis b. decomposition
   c. single replacement d. double replacement

5. What particle is transferred from one atom to another in a redox reaction?
   a. electron b. neutron
   c. proton d. nucleus

6. Which of the following is a single replacement?
   a. $KOH + HCl \longrightarrow KCl + H_2O$
   b. $2Na + 2H_2O \longrightarrow H_2 + 2NaOH$
   c. $2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$
   d. $H_2O + CO_2 \longrightarrow H_2CO_3$

7. How are reactions related to chemical bonds?
   a. Bonds in the reactants are broken, and bonds in the products are formed.
   b. Bonds in the products are broken, and bonds in the reactants are formed.
   c. Bonds in both the reactants and products are broken.
   d. Bonds are formed in both the reactants and the products.

8. What type of reaction always releases energy?
   a. endothermic b. exothermic
   c. decomposition d. oxidation-reduction

9. In general, an increase in temperature
   a. increases reaction rate.
   b. decreases reaction rate.
   c. does not affect reaction rate.
   d. acts as a catalyst.

10. What takes place at chemical equilibrium?
   a. Reactants form more quickly than products.
   b. Products form more quickly than reactants.
   c. Reactants and products stop forming.
   d. Reactants and products form at the same rate.

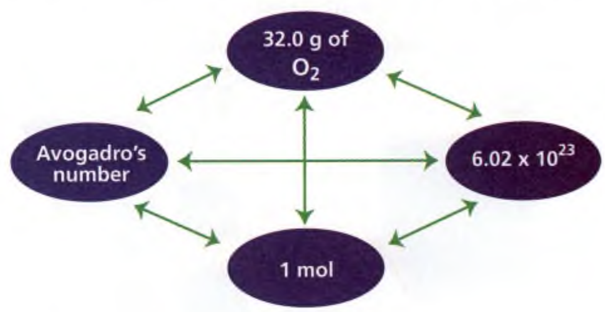
Understanding Concepts

11. Write the following chemical equation in words.

$$CaCO_3 + \text{Heat} \longrightarrow CaO + CO_2$$

12. Explain how a balanced chemical equation shows that mass is conserved.

13. Explain the following diagram in your own words.



14. Compare oxidation and reduction.

15. Paper burns by combining with oxygen. Why doesn't paper burn every time it contacts oxygen?

16. Give an example of a chemical reaction that occurs slowly. Give another example of a chemical reaction that occurs quickly.

Answer Questions 17–18 based on the equation below.

$$C(s) + H_2O(g) + \text{Heat} \rightleftharpoons CO(g) + H_2(g)$$

17. How would you adjust the temperature to increase the amount of product?

18. Does the removal of hydrogen gas as it is produced shift the reaction to the left or the right?