CHAPTER 20
Chemical Reactions and Energy

Chapter Preview

Sections

20.1 Energy Changes in Chemical Reactions

MiniLab 20.1 Dissolving—Exothermic or Endothermic?

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ChemLab Energy Content of Some Common Foods

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Wow, That's Hot Stuff!

This chemical reaction produces a lot of thermal energy. In fact, this reaction is used to fuse metals together by creating a metal seam between the two pieces. Two common gases used for this process are acetylene and oxygen.

Start-up Activities

Launch Lab

Speeding Reactions

Many chemical reactions occur so slowly that you don't even know they are happening. For some reactions, it is possible to alter the reaction speed using another substance.

Safety Precautions

Materials

- hydrogen peroxide
- beaker or cup
- baker's yeast
- toothpicks

Procedure

1. Create a "before and after" table and record your observations.
2. Pour about 10 mL of hydrogen peroxide into a small beaker or cup. Observe the hydrogen peroxide.
3. Add a "pinch" (1/8 tsp) of yeast to the hydrogen peroxide. Stir gently with a toothpick and observe the mixture again.

Analysis

Into what two products does hydrogen peroxide decompose? Why aren't bubbles produced in step 1? What is the function of the yeast?

What I Already Know

Review the following concepts before studying this chapter.

Chapter 1: energy in chemical changes

Chapter 6: reasons why reactions go forward; equilibrium; and reaction rates

Chapter 10: temperature and particle motion

Reading Chemistry

Analyze some of the charts, graphs and tables that appear throughout the chapter. Write down any questions or new vocabulary words that are used in the figures. Note the page number of figures you have questions about, and try to find the answers in the text as you read the chapters.

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SECTION 20.1

Energy Changes in Chemical Reactions

SECTION PREVIEW

Objectives

- ✓ **Compare** and **contrast** exothermic and endothermic chemical reactions.
- ✓ **Analyze** the energetics of typical chemical reactions.
- ✓ **Illustrate** the meaning of *entropy*, and **trace** its role in various processes.

Review Vocabulary

Electron transport chain: the controlled release of energy from glucose by the step-by-step movement of electrons to lower energy levels.

New Vocabulary

heat
law of conservation of energy
fossil fuel
entropy

"Smile," says the photographer, as she pushes a button. The camera shutter opens and an electric current from a small lithium battery sparks across a gap in the flash unit. This spark ionizes xenon gas, creating a bright flash of light. The energy from the chemical reaction in the lithium battery has been successfully put to use.

Exothermic and Endothermic Reactions

As you learned in Chapter 1, chemical reactions can be exothermic or endothermic. Recall that an exothermic reaction releases heat, and an endothermic reaction absorbs heat. **Figure 20.1** pictures an exothermic process as a reaction that's going downhill energetically and an endothermic process as a reaction that's going uphill.

Exothermic Reactions

If you have ever started a campfire or built a fire in a fireplace, you know that the burning of wood is an example of an exothermic process. Once you have ignited the wood, the reaction generates enough heat to keep itself going. A net release of heat occurs, which is what makes the reaction exothermic.

Figure 20.1
Exothermic and Endothermic Reactions
The exothermic reaction (left) gives off heat because the products are at a lower energy level than the reactants. The endothermic reaction (right) absorbs heat because the products are at a higher energy level than the reactants.

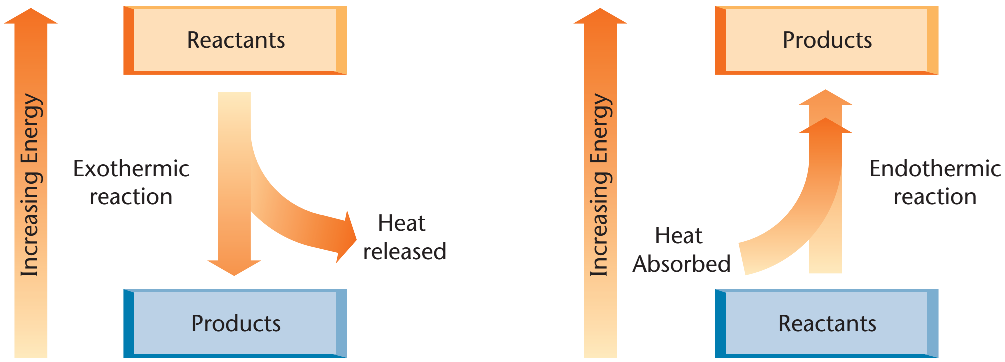
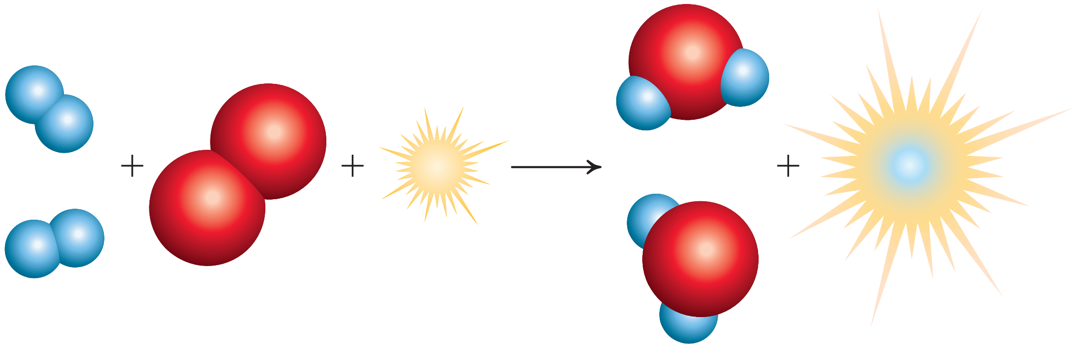


Figure 20.2

The Exothermic Formation of Water

It takes only a small amount of energy to start the reaction between hydrogen and oxygen to form water. The energy released by the reaction is much greater, so the reaction is exothermic. The energy from this reaction has been used to power car and truck engines.



$2\ H_2$ + O_2 + a little energy $\longrightarrow$ $2\ H_2O$ + a lot of energy

The reaction between hydrogen gas and oxygen gas to form water, shown in **Figure 20.2,** is another example of an exothermic reaction. Once a small amount of energy—often just a spark—is added to the mixture of gases, the reaction continues to completion, usually explosively. No additional input of energy from outside is needed to keep it going. Once energy has been supplied to break the covalent bonds in the first few molecules of hydrogen and oxygen, the atoms combine to form water and release enough energy to break the bonds in additional hydrogen and oxygen molecules. The net energy is released as heat.

Endothermic Reactions

Consider the reverse of the reaction just discussed. Just as water can be formed from hydrogen and oxygen, it can also be decomposed to re-form hydrogen and oxygen. In the process of electrolysis, electrical energy is used to break the covalent bonds that unite the hydrogen atoms and the oxygen atoms in the water molecules. The hydrogen atoms pair up to form hydrogen molecules, and the oxygen atoms pair up to form oxygen molecules. The formation of the new bonds releases energy, but not as much as the amount required during the bond breaking. Additional energy must be added continuously during the electrolysis. The reaction absorbs heat energy and is, therefore, endothermic.

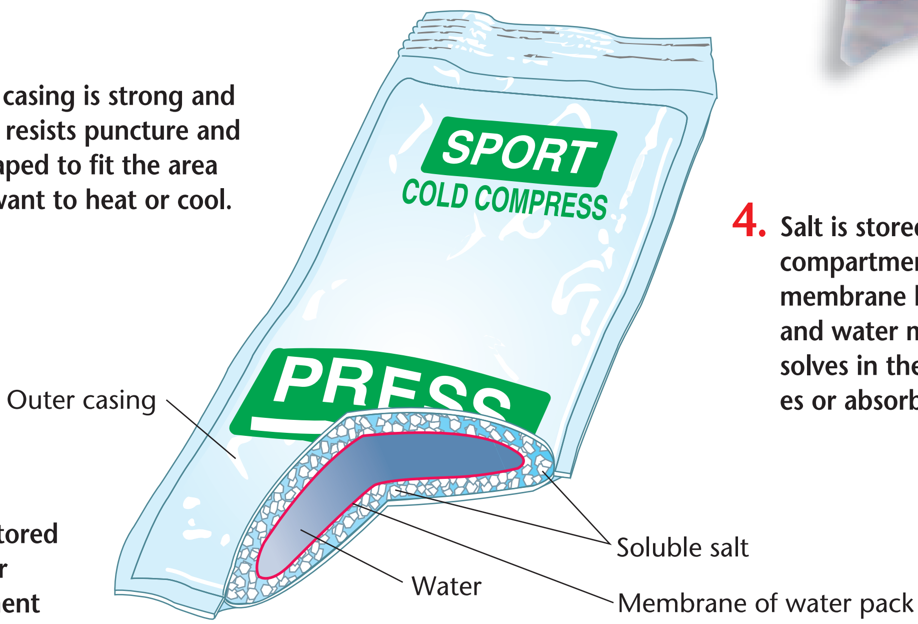
All endothermic reactions are characterized by a net absorption of energy. In the History Connection on page 58, you read about another example of an endothermic reaction: the decomposition of orange mercuric oxide into the elements mercury and oxygen. As long as heat is applied, the compound continues to decompose, but if the heat source is removed, the reaction stops. The net absorption of heat energy that is required is what makes the reaction endothermic.

How it Works

Hot and Cold Packs

Instant hot and cold packs create aqueous solutions that form exothermically or endothermically and therefore release or absorb heat. A hot pack generates heat when a salt such as calcium chloride dissolves in water that is stored in the pack. The calcium chloride dissolves exothermically. A cold pack absorbs heat when a salt such as ammonium nitrate dissolves in water. The ammonium nitrate dissolves endothermically. In both cases, the salt and water are separated by a thin membrane. All you have to do is squeeze the pack to mix the components and you have instant heat or cold at your fingertips.

1. The outer casing is strong and flexible. It resists puncture and can be shaped to fit the area that you want to heat or cool.



2. Water is stored in an inner compartment separate from the solid salt.

3. The inner membrane breaks easily when you knead or squeeze the pack or strike it sharply.

4. Salt is stored in the outer compartment. When the inner membrane breaks, the salt and water mix. The salt dissolves in the water and releases or absorbs energy.

Thinking Critically

1. Another type of hand warmer contains fine iron powder and chemicals that cause the iron to rust. The rusting of iron lets the hand warmer maintain temperatures above 60°C for several hours. Explain how that is possible.
2. When a solid in a cold pack dissolves in water, the process takes place spontaneously. What causes the solution process to be spontaneous in spite of the fact that it is endothermic?

Heat

The energy that is involved in exothermic and endothermic reactions is usually in the form of heat. **Heat** is defined as the energy transferred from an object at high temperature to an object at lower temperature. Recall that energy is measured in joules; the symbol for joules is *J*. The symbol for a kilojoule, which is equal to 1000 J, is *kJ*.

WORD ORIGIN

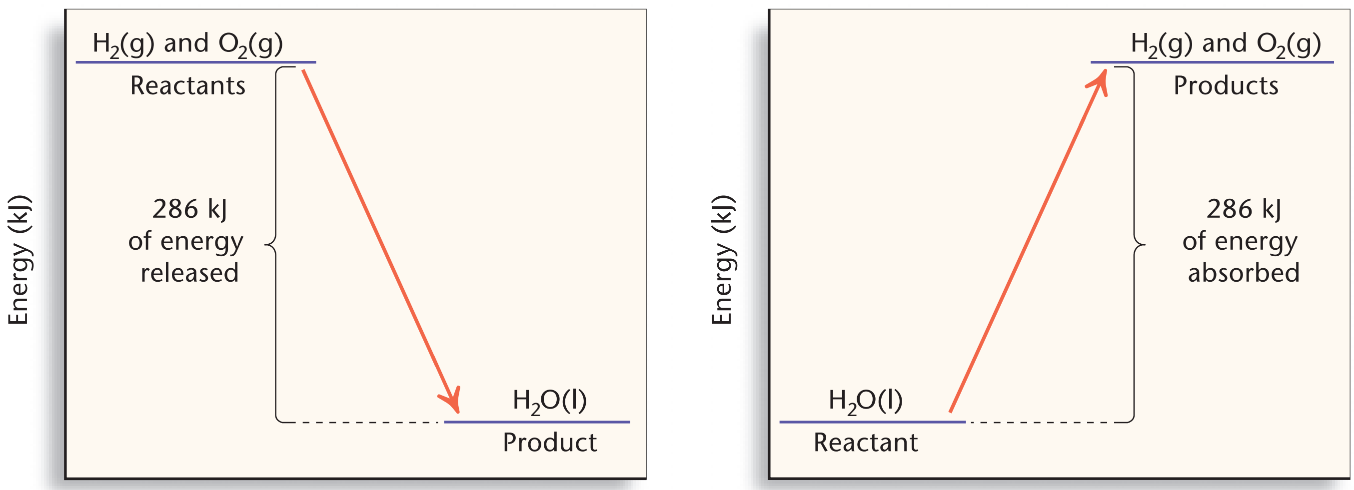
energy:
en (GK) in
ergon (GK) work
A person who has a great deal of energy can work hard all day.

Using Symbols to Show Energy Changes

Energy changes are frequently included in the equation for a chemical reaction. The amount of heat absorbed or evolved during a reaction is a measure of the energy change that accompanies the reaction. When 1 mol (18.0 g) of liquid water is produced from hydrogen and oxygen gas, 286 kJ of energy are given off. This means that the energy of the uncombined hydrogen gas and oxygen is greater than the energy of the water. When 1 mol (18.0 g) of liquid water decomposes to form hydrogen gas and oxygen gas, 286 kJ of energy are absorbed. This also shows that the energy of the uncombined hydrogen and oxygen is greater than the energy of the water. The graphs in **Figure 20.3** illustrate this relationship.

Scientists have observed that the energy released in the formation of a compound from its elements is always identical to the energy required to decompose that compound into its elements. This observation is an illustration of an important scientific principle known as the **law of conservation of energy.** That law states that energy is neither created nor destroyed in a chemical change, but is simply changed from one form to another. In an exothermic reaction, the heat released comes from the change from reactants at higher energy to products at lower energy. In an endothermic reaction, the heat absorbed comes from the opposite change.

Figure 20.3
Formation and Decomposition of Water
As the graphs show, the energy produced when 1 mol of liquid water forms from the elements hydrogen and oxygen is equal in magnitude to the energy absorbed when the water decomposes.



Formation of $H_2O(l)$

Decomposition of $H_2O(l)$

miniLAB 1

Dissolving—Exothermic or Endothermic?

The dissolving of a solid in water, like most other processes, may liberate energy or absorb energy. If the dissolving process is exothermic, the liberated energy raises the temperature of the solution. If the process is endothermic, it absorbs energy from the solution, lowering the temperature. Examine the dissolution of several common solids in water.

Procedure

1. Measure 100 mL of water into a 250-mL beaker. Set a Celsius thermometer in the water and allow it to come to the water's temperature. Record that temperature. Remove the thermometer from the water.
2. Add to the water approximately 1 tablespoon of the solid to be tested, and stir with a stirring rod for 20 seconds. Put the thermometer back into the solution. Record the temperature.
3. Pour the solution down the drain with large amounts of tap water.
4. Repeat steps 1 through 3 for each of the solids to be tested.

Analysis

1. Which of the solids that you tested dissolve exothermically? Which dissolve endothermically?
2. Which of the solids that you tested could be mixed with water inside a flexible plastic container to produce a cold pack that might be used by medical personnel?

The difference in energy between products and reactants in a chemical change is symbolized ΔH (delta H), where the symbol Δ means a difference or change and the letter H represents the energy. The energy absorbed or released in a reaction ($\Delta H_{reaction}$) is related to the energy of the products and the reactants by the following equation.

$$\Delta H_{reaction} = H_{products} - H_{reactants}$$

For exothermic reactions, ΔH is negative because the energy stored in the products is less than that in the reactants. For endothermic reactions, ΔH is positive because the energy of the products is greater than that of the reactants. The value of ΔH is often shown at the end of a chemical equation. For example, the exothermic formation of 2 mol of liquid water from hydrogen and oxygen gas would be written like this.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad \Delta H = -572 \text{ kJ}$$

The equation for the endothermic decomposition of 2 mol of liquid water would be written like this.

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g) \quad \Delta H = +572 \text{ kJ}$$

The value 572 kJ in these equations is 2×286 kJ, which is the amount of energy released when 1 mol of liquid water forms. Note the use of the symbols (s), (l), and (g). When energy values are included with an equation for a reaction, it is especially important to show the states of reactants and products because the energy change in a reaction can depend greatly upon physical states.

Activation Energy

The hydrocarbons found in petroleum and natural gas are the remains of plants and other organisms that lived millions of years ago. Oil and natural gas are called **fossil fuels** for this reason. Fossil fuels are a rich source of energy because when they react with oxygen to produce carbon dioxide and water, a great deal of energy is released in the form of heat.

However, fossil fuels do not burn automatically. Energy, usually in the form of heat or light, is needed to get the chemical reaction started. The combustion of hydrocarbon fuels requires this input of energy, called activation energy, to begin the reaction. For example, the butane gas in a disposable lighter requires a spark to start the combustion of the gas.

In Chapter 6, you learned that activation energy is needed to cause particles to collide with enough force to make them react. Activation energy is required in both exothermic and endothermic reactions. The fact that a fuel requires an input of energy—such as from a spark—to begin burning does not mean that the combustion reaction is endothermic. The reaction releases a net amount of heat, and so it is exothermic.

Look more closely at activation energy and heat of reaction, using as an example the burning of methane, which is the main component in natural gas, to yield carbon dioxide and water vapor. The equation for this reaction is as follows.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \quad \Delta H = -802 \text{ kJ}$$

A graph of the energy change during the progress of this reaction is shown in **Figure 20.4.** Notice how the energy curve rises, then falls. The rise represents the activation energy, which is the energy difference between the reactants and the maximum energy stage in the reaction. The fall represents the energy liberated by the formation of new chemical substances. When 1 mol of methane burns, 802 kJ of heat are given off. The reaction is exothermic, as is shown by the negative sign of ΔH. The energy stored in the products is less than that stored in the reactants, so a net amount of energy is released. Some of the released energy provides the activation energy needed to keep the reaction going.

WORD ORIGIN

combustion:
combustus **(L)**
burned

The internal combustion engine in a car burns the hydrocarbons in gasoline.

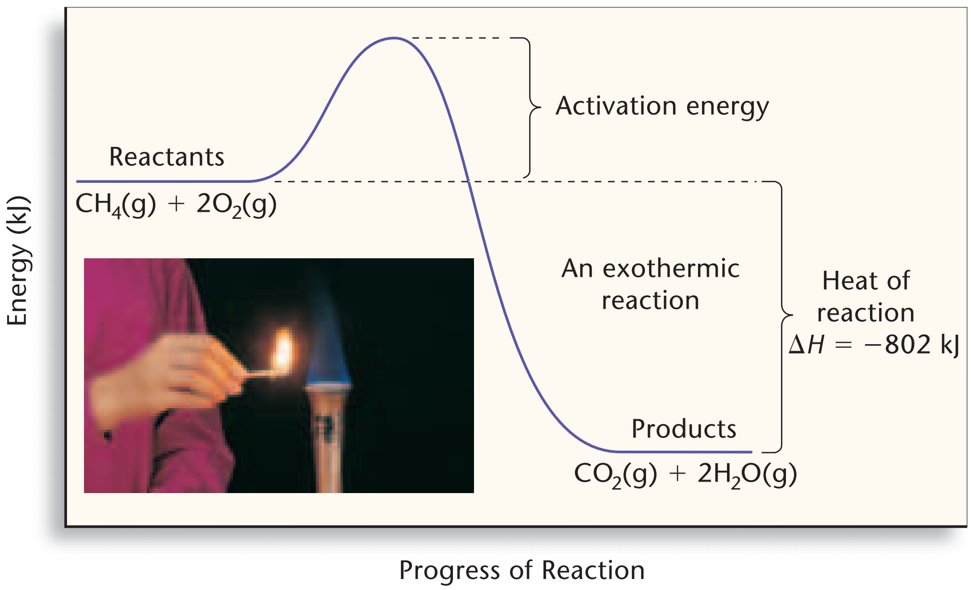
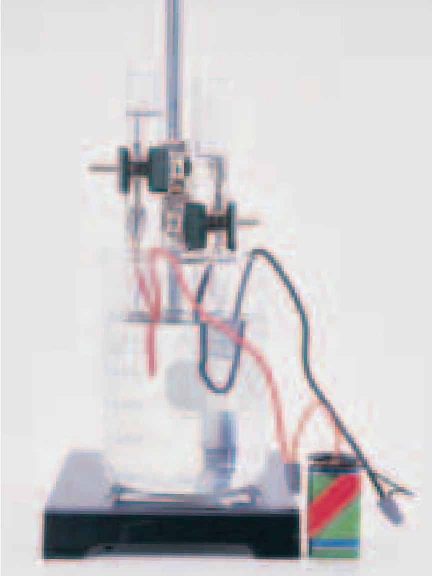


Figure 20.4
Energy in an Exothermic Reaction
In order to occur, the combustion of methane, illustrated in the photo at the left, requires an input of activation energy—in this case, provided by a match or sparking device in the stove. Overall, the reaction releases 802 kJ of energy per mole of methane. Notice from the graph that the products are in a lower energy state than the reactants. The negative value of ΔH reflects this fact.



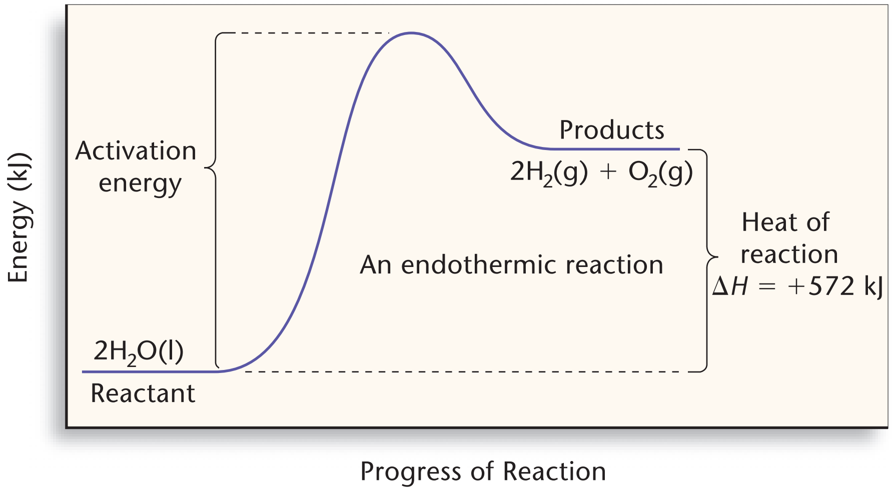


Figure 20.5
Energy in an Endothermic Reaction
The decomposition of water requires a continuous input of energy, such as electrical energy from a battery. Overall, the reaction absorbs 572 kJ of energy for every 2 mol of liquid water. The products are in a higher energy state than the reactant. The positive value of ΔH reflects this fact.

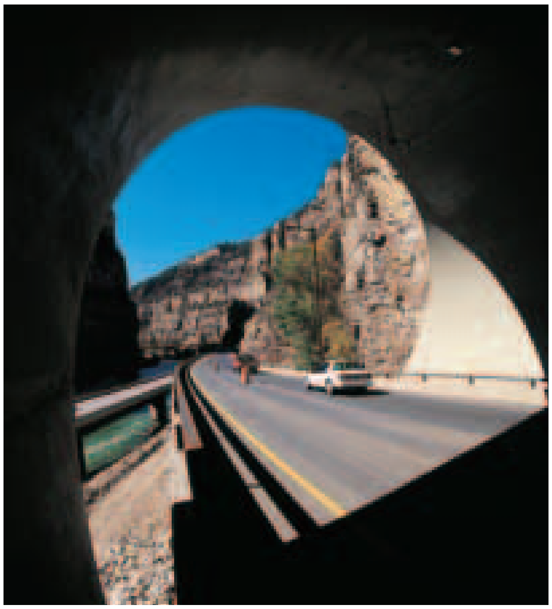
Now, consider an example of an endothermic reaction—one that you have already read about—the decomposition of water. The equation for the reaction is as follows.

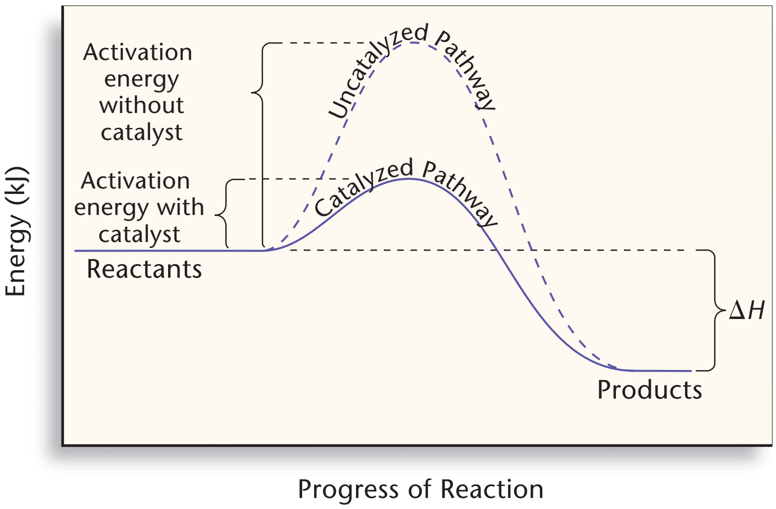
$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g) \quad \Delta H = +572 \text{ kJ}$$

A graph of the energy changes during the progress of this reaction is shown in **Figure 20.5.** Notice how the energy curve rises above the level of the reactant, then falls only slightly when the products form. Therefore, the products have more stored energy than the reactants, and the graph clearly shows a net gain of energy. Because of this net gain, ΔH is positive. Energy must be added continuously to keep the reaction going.

Recall from Chapter 6 that a catalyst can be used to speed up a reaction. The catalyst provides a different reaction path—one for which the activation energy is lower. The catalyst thus creates a shortcut or tunnel through the energy hill between the reactants and products. More of the collisions will now be effective because less energy is required. The energy diagram in **Figure 20.6** shows the effect of a catalyst on activation energy.

Figure 20.6
Effect of a Catalyst
It is easier for cars to go through the tunnel than for them to climb the mountain. ▼





▲ As the graph of a catalyzed reaction shows, the activation energy is lowered, providing an easier reaction pathway. As a result, the reaction goes faster than would be the case without the catalyst. Note that the heat of reaction, ΔH, is the same in both cases.

Everyday Chemistry

Catalytic Converters

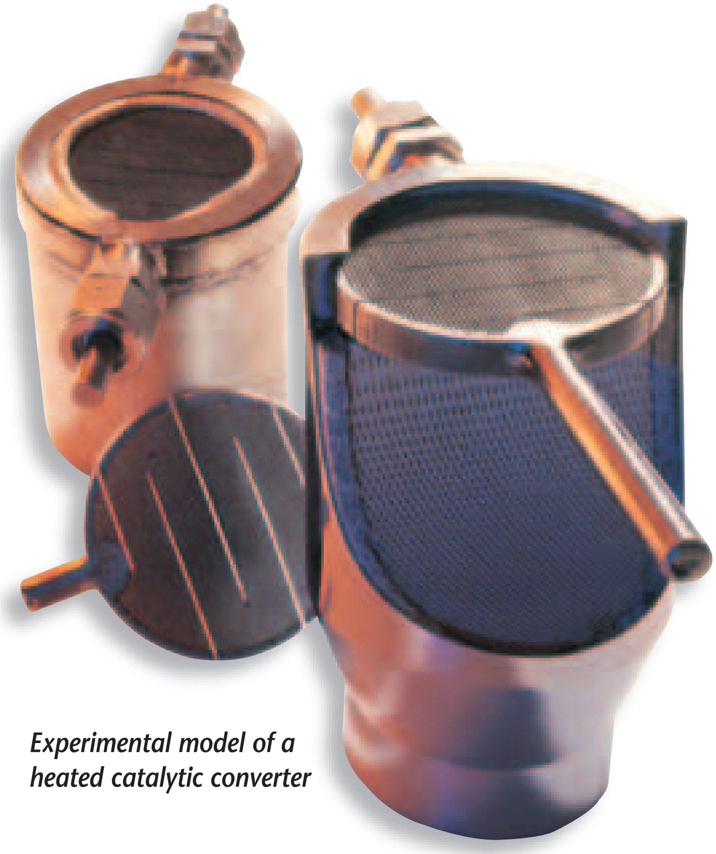
Since 1975, every new car sold in the United States has a catalytic converter installed in the exhaust system. This device contains porous, heat-resistant material coated with a catalyst. The purpose of the catalytic converter is to reduce air pollution.

How a catalytic converter works A typical catalytic converter consists of particles of platinum and rhodium deposited on a ceramic structure that is like a honeycomb. The platinum and rhodium catalyze reactions that remove pollutants such as nitrogen monoxide (NO), carbon monoxide (CO), and unburned hydrocarbons. When nitrogen monoxide binds to the rhodium surface, it breaks down to oxygen and nitrogen. The bound oxygen reacts with carbon monoxide, which has also become bound to the rhodium surface. The reaction produces carbon dioxide. The oxidation of unburned hydrocarbons produces carbon dioxide and water.

The honeycomb arrangement of catalytic materials in the converter provides a large surface area for the reactions to take place. Such an arrangement increases the rate at which the pollutants are removed. Catalytic converters have reduced the pollutants released into the air by cars by as much as 90 percent.

Improving the catalytic converter The operating temperature of a catalytic converter is between 316° and 649°C, which is in the range of standard exhaust-gas temperatures for a vehicle being driven on the road. Below 316°C, the device does nothing. The pollutants emitted during the beginning of the driving period, when the converter is just heating up, slip through the catalytic converter without being changed.

In an effort to stop air pollution during the warm-up period, some catalytic converters heat up to 400°C within 5 seconds. The result is that almost as soon as the car is started, air pollutants in the exhaust are broken down. Heated catalytic converters are proving to be highly efficient.



Experimental model of a heated catalytic converter

Someday, perhaps even those 5 seconds of dirty emissions before the converter is heated will be eliminated.

Exploring Further

1. **Acquiring Information** Find out why cars with catalytic converters cannot use leaded gasoline.
2. **Applying** Instead of the honeycomb structure of the catalytic converter described above, some converters contain pellets coated with a platinum-palladium mixture. Why is the use of pellets effective?

Chemistry Online

To learn more about how catalytic converters cut down pollution, visit the Chemistry Web site at chemistryca.com

Forces That Drive Chemical Reactions

When pure aluminum metal is exposed to chlorine gas, aluminum chloride is formed. This chemical change is spontaneous; that is, it just happens on its own. The reaction is highly exothermic. The following equation summarizes the process.

$$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s) + 1408\ kJ$$

What allows such reactions to occur spontaneously? The answer has to do with the general forces that drive all reactions.

Order, Disorder, and Entropy

Scientists have observed two tendencies in nature that explain why chemical reactions occur. The first tendency is for systems to go from a state of high energy to a state of low energy. The state of low energy is more stable. Thus, for example, exothermic reactions are more likely to occur than endothermic ones, all other things being equal. The second tendency is for systems to become more disordered. Spontaneous reactions tend to occur if energy decreases and if disorder increases.

Figure 20.7 shows an everyday spontaneous change that increases disorder. The first photo shows a full glass of milk on the breakfast table. The second photo shows the same glass of milk after someone has bumped into it and caused it to fall to the floor. Notice the disorder in the second picture. The glass is broken into seemingly random shards, and the milk is scattered into pools and droplets.

You have already learned a good deal about energy changes, but the concept of disorder in chemical changes may not be as familiar. Scientists use the term **entropy** to describe and measure the degree of disorder. Unlike energy, which is conserved in chemical changes and in the universe as a whole, entropy is not conserved. The natural tendency of entropy is to increase. The fallen glass illustrates how entropy increases in natural, spontaneous changes because when disorder increases, entropy increases. In most spontaneous, naturally occurring processes, entropy increases.

Figure 20.7

Disorder and Spontaneity

When the glass full of milk falls off a table, it breaks and the contents spill. The broken glass is in a state of greater disorder. This breaking and scattering is a spontaneous change. The reverse change—the glass reassembling itself and refilling with the milk—is extremely unlikely to occur. Changes that are spontaneous in one direction are not spontaneous in the other direction under the same conditions.

Some kinds of changes are apt to increase entropy, **Figure 20.8.** These changes include melting and evaporation. Increases in numbers of molecules also tend to increase entropy. Entropy increases during reactions that result in increases in the number of molecules. Such changes result in greater disorder at the submicroscopic level.

Figure 20.8

Increase in Entropy

Humpty-Dumpty's problem was too great an increase in entropy. An input of work can sometimes counter entropy increases. You know that living things require a constant supply of energy for life functions. One of the ways they use that energy is to maintain the strict organization of molecules required for life. The energy is used to do work to overcome entropy increases at the molecular level.

The Direction of a Chemical Reaction

At room temperature, most exothermic reactions tend to proceed spontaneously forward. In other words, they favor the formation of products. In an exothermic reaction, the released energy, usually in the form of heat, raises the temperature of the products and many more atoms and molecules in the surroundings. The energy is distributed more randomly than it was before the reaction. The motion of a greater number of atoms and molecules increases. Therefore, disorder increases.

You have read that both heat and entropy play a role in determining spontaneity. In general, the direction of a chemical reaction is determined by the magnitude and direction of the heat energy and entropy changes. For example, a reaction will proceed in the forward direction, toward formation of products, if that direction results in both a release of heat and an increase in entropy. As an example, consider the combustion of butane, C_4H_{10}.

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g) + \text{heat}$$

This reaction occurs spontaneously in the forward direction because both energy and entropy changes are favorable. The energy decreases because the reaction is exothermic—a favorable change. The entropy increases partially because the total number of molecules increases, from 15 to 18—also a favorable change. The release of heat and increase in entropy combine to drive this reaction forward.

Look at another example of a reaction favored by both energy and entropy changes—the one between calcium carbonate and hydrochloric acid.

Supplemental Problems

For more practice with solving problems, see Supplemental Practice Problems, Appendix B.

Try at Home Lab

See page 872 in Appendix F for **Observing Entropy**

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g) + \text{heat}$$

This reaction favors formation of products because it produces gases and liquids. They are more disordered than the solid $CaCO_3$, so entropy increases. Heat is given off, so the products are at a lower energy state than the reactants. That is also favorable. Once again, the decreased energy and increased entropy both drive the reaction in the forward direction.

What happens if only one of the changes—either energy or entropy—is favorable? If the favorable change is great enough to outweigh the unfavorable one, the reaction will still be spontaneous. Thus, even some endothermic reactions are spontaneous if disorder increases greatly. Also, some reactions that increase order are spontaneous if they are exothermic enough. Spontaneity depends on the balance between energy and entropy factors.

Table 20.1 summarizes these factors. Note that temperature plays a role in determining spontaneity when one factor is favorable and the other is not.

Table 20.1 Predicting Whether a Reaction Is Spontaneous in the Forward Direction

Energy Change	Entropy Change	Spontaneous?
decrease (exothermic)	increase	yes
decrease (exothermic)	decrease	yes at low temperature; no at high temperature
increase (endothermic)	increase	no at low temperature; yes at high temperature
increase (endothermic)	decrease	no

SECTION REVIEW

Understanding Concepts

1. Are reactions that occur spontaneously at room temperature generally exothermic or endothermic? Explain.
2. Describe the energy changes that occur when a match lights.
3. ΔH for a reaction is negative; compare the energy of the products and the reactants. Is the reaction endothermic or exothermic?

Thinking Critically

4. **Relating Cause and Effect** Describe each of the following processes as involving an increase or decrease in entropy.
   - **a)** Water in an ice-cube tray freezes.
   - **b)** You pick up scattered trash along a highway and pack it into a bag.
   - **c)** Your campfire burns, leaving gray ashes.

Applying Chemistry

5. **Fermentation** Yeast can spontaneously ferment the sugar in grapes or apples and form ethanol. When the process occurs in grapes, it results in wine. In apples, it produces a beverage called hard cider. The first stage in the fermentation process is exothermic, and the balanced equation for the reaction is as follows.

   $$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

   Determine whether entropy increases in this reaction, and relate the change in heat energy and entropy to the spontaneity of the process.

SECTION 20.2

Measuring Energy Changes

These days, many people are trying to avoid fattening foods and looking for things to eat that taste good and provide adequate nutrition without increasing body weight. Have you ever been on a diet and had to count Calories? If so, you may have wondered how food manufacturers determine the amount of energy available in a certain amount of food.

SECTION PREVIEW

Objectives

✓ **Sequence** the technique of calorimetry and **illustrate** its use.

✓ **Compare** the heat generated by some common fuels and by some foods.

✓ **Analyze** the efficiency of industrial processes and the need to conserve resources.

Review Vocabulary

Entropy: term used to describe and measure the degree of disorder in progress.

New Vocabulary

calorie
kilocalorie
Calorie

Calorimetry

The heat generated in chemical reactions can be measured by a technique called calorimetry and a device called a calorimeter. The device and its use are illustrated in **Figure 20.9,** which deals with an exothermic reaction. A calorimeter can also be used to study the heat absorbed in endothermic reactions. In the case of endothermic reactions, the surrounding water in the calorimeter supplies the heat and decreases in temperature.

In measurements involving a calorimeter, you calculate the heat lost or gained by the surrounding water. This is done by means of the following equation.

$$q_w = (m)(\Delta T)(C_w)$$

In this equation, the symbol q_w stands for heat absorbed by water, m is the mass of the water, ΔT is the temperature change of the water, and C_w is the specific heat of water, which equals 4.184 J/g•°C.

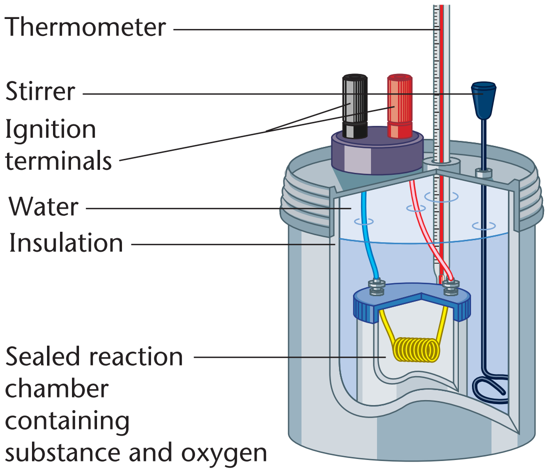


Figure 20.9
Calorimetry
A weighed sample of a substance is burned in pure oxygen inside a container called a reaction chamber. The heat that is released flows into the surrounding water and raises its temperature. Given the temperature change and mass of the water, you can calculate the heat given off in the reaction.

By the law of conservation of energy, there is no net creation or destruction of energy. Any heat absorbed or released by the water has been released or absorbed by the reaction being studied. The symbol q_{reaction} stands for the heat change of the reaction. Heat loss by the reaction means heat gain for the water, and heat gain for the reaction means heat loss for the water. Therefore, the heat of reaction is equal to the negative of the heat change of the water.

$$q_{\text{reaction}} = -q_{\text{w}}$$

The following sample problem shows how calorimetry can be used to measure energy change in a chemical reaction.

SAMPLE PROBLEM 1 Calculating Heat of Reaction for Combustion

The burning of 1.60 g of methane in oxygen, to yield carbon dioxide gas and liquid water, causes the surrounding 1.52 kg of water in a calorimeter to change in temperature from 20.0°C to 34.0°C. What is the heat of reaction for this combustion of 1 mol of methane?

Analyze

- You have all the information you need to find the heat change for the water, q_{w}. This can be set equal to $-q_{\text{reaction}}$ for the burning of the 1.60 g of methane. What you then need to do is convert that value to find q_{reaction} for 1.0 mol of methane. That will equal ΔH.

Set Up

- Find the temperature change of the water, using the equation $\Delta T = T_{\text{final}} - T_{\text{initial}}$. Then, use the equation $q_{\text{w}} = (m)(\Delta T)(C_{\text{w}})$. The negative of the value of q_{w} will equal q_{reaction} for the burning of 1.60 g of methane. Use the molar mass of methane, 16.0 g methane/1.0 mol methane, as a conversion factor to find the q_{reaction} for 1.0 mol of methane.

Solve

- First, calculate ΔT.

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$
$$= 34.0°\text{C} - 20.0°\text{C}$$
$$= 14.0°\text{C}$$

Then, calculate q_{w}.

$$q_{\text{w}} = (m)(\Delta T)(C_{\text{w}})$$

$$= \frac{1.52 \times 10^3\ \cancel{\text{g}}}{} \left| \frac{14.0\cancel{°\text{C}}}{} \right| \frac{4.184\ \text{J}}{\cancel{\text{g}}\cdot\cancel{°\text{C}}} = 8.90 \times 10^4\ \text{J} = 89.0\ \text{kJ}$$

Now, find q_{reaction} for the 1.60 g of methane.

$q_{\text{reaction}} = -q_{\text{w}} = -89.0$ kJ = heat released in burning 1.60 g methane

Now, use a molar-mass conversion factor to find q_{reaction}, or ΔH, for 1 mol of methane.

$$q_{\text{reaction}} = \frac{-89.0\ \text{kJ}}{1.60\ \cancel{\text{g methane}}} \left| \frac{16.0\ \cancel{\text{g methane}}}{1\ \text{mol methane}} \right. = -890\ \text{kJ/mol methane}$$

$$q_{\text{reaction}} = \Delta H = -890\ \text{kJ/mol methane}$$

Check

- Check to be sure that all units were correctly handled, and review the calculation to be sure that it was sound. The result does check out.

Problem-Solving HINT

Be careful with signs in the calculation of ΔT. If the reaction is endothermic, the sign of ΔT will be negative.

Supplemental Problems

For more practice with solving problems, see Supplemental Practice Problems, Appendix B.

PRACTICE PROBLEMS

1. How much heat is absorbed by a reaction that lowers the temperature of 500.0 g of water in a calorimeter by 1.10°C?
2. Aluminum reacts with iron(III) oxide to yield aluminum oxide and iron. Calculate the heat given off in the reaction if the temperature of the 1.00 kg of water in the calorimeter increases by 3.00°C.
3. When 1.00 g of a certain fuel gas is burned in a calorimeter, the temperature of the surrounding 1.000 kg of water increases from 20.00°C to 28.05°C. All products and reactants in the process are gases. Calculate the heat given off in this reaction. How much heat would 1.00 mol of the fuel give off, assuming a molar mass of 65.8 g/mol?

Energy Value of Food

Many years ago, chemists measured heat in calories instead of joules. A **calorie** is the heat required to raise the temperature of 1 g of liquid water by 1°C. A **kilocalorie** is a unit equal to 1000 calories. One calorie is equal to 4.184 J. One joule is equal to 0.239 calorie. Some nutritionists and dietitians still use Calories but are switching to kilojoules. The energy value of foods is measured in units called Calories. Note the capital *C*. One food **Calorie** is the same as 1 kilocalorie. One food Calorie is also equal to 4.184 kJ.

Chemical compounds in the food you eat provide you with energy. The compounds undergo slow combustion, combining with oxygen to produce the waste products carbon dioxide and water, along with compounds needed for body growth and development. Fats provide about 9 Calories per gram. In contrast, 1 g of carbohydrate or protein provides about 4 Calories. **Figure 20.10** illustrates some types of nutrients. **Table 20.2** shows how many Calories you get from typical servings of some foods.

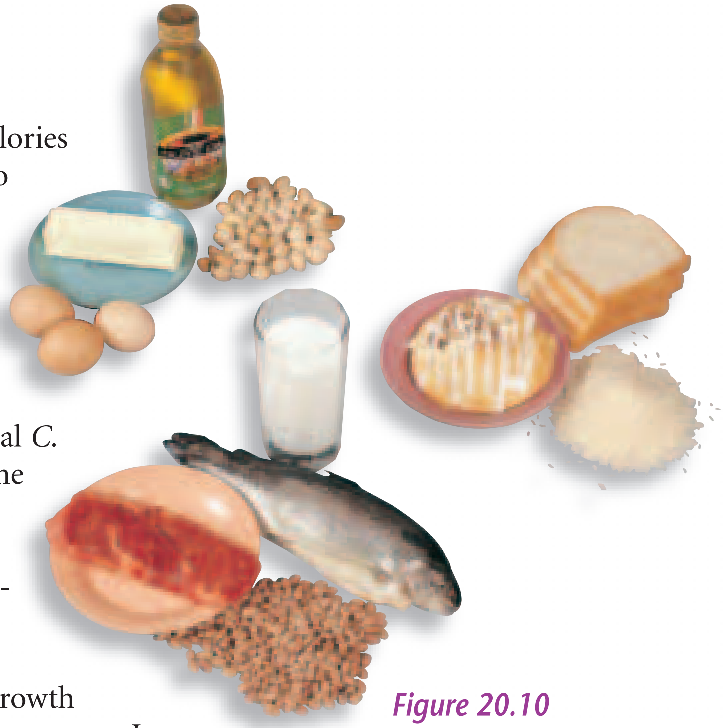


Figure 20.10
Types of Nutrients
Foods such as those shown here contain a variety of chemical compounds. They can be grouped into three general types—protein, carbohydrate, and fat—that have different energy contents. Your body processes transform some of the energy released when these foods are broken down into work and heat. However, when you take in more food than you need, your body stores the extra energy by producing fat for later use.

Table 20.2 The Caloric Value of Some Foods

Food	Quantity	Kilojoules	Calories
butter	1 tbsp = 14 g	418	100
peanut butter	1 tbsp = 16 g	418	100
spaghetti	0.5 cup = 55 g	836	200
apple	1	283	70
chicken (broiled)	3 oz = 84 g	502	120
beef (broiled)	3 oz = 84 g	1000	241

ChemLab

SMALL SCALE

Energy Content of Some Common Foods

Foods supply the energy and nutrients you require to build and sustain your body and maintain your levels of activity. The energy is released within cells during the process of respiration. In that process, oxygen combines with energy-storage substances, such as glucose, to produce carbon dioxide, water, and heat. The reaction is essentially a slow combustion process. In this ChemLab, you will compare the amounts of energy liberated from the combustion of three food items.

Problem

How much energy is released during the combustion of some common food items?

Objectives

- **Interpret data** to calculate the energy released during the combustion of pecans, marshmallows, and a food item of your choice.
- **Compare** the energies obtained from the food items.
- **Infer,** based upon the chemical compositions of the foods and upon the amounts of energy obtained, which types of foods contain the greatest amount of energy.

PREPARATION

Materials

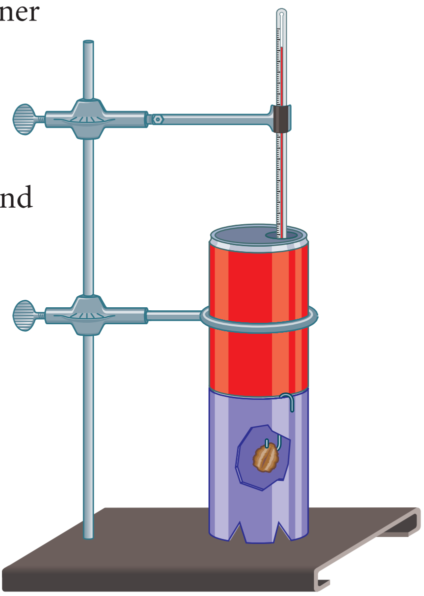
oven mitt
bottle opener with can-piercing end
empty, clean soft-drink can with a tab closure
empty clean can, minus the top and bottom lids, of approximately the same diameter as the soft-drink can
ring stand
small iron ring
beaker tongs
Celsius thermometer
100-mL graduated cylinder
glass stirring rod
balance
large paper clip
matches
pecan half
2 small marshmallows
food sample of your choice

Safety Precautions

Be careful in using the match and in handling the cans after each experiment because they may be hot. Perform the experiment in a well-ventilated room because acrid fumes may be produced during the combustion process. Wear an apron and goggles.

PROCEDURE

1. Use a can opener to open holes near the bottom of the open-ended can. Bend a large paper clip to fashion the food support as shown in the figure.



2. Use the graduated cylinder to measure 100 mL of tap water, and pour the water into the soft-drink can. Support this can by the ring stand as shown in the figure. Put the thermometer through the opening in the top of the can and allow it to reach the temperature of the water. Record the initial water temperature to the nearest 0.1°C.
3. Weigh a pecan half, record its mass, and impale it on the paper-clip food support. Hang the support on the edge of the open-ended can with the food sample inside, and rest the can on the ring stand.
4. Use a match to light the pecan. Quickly swing the water-filled soft-drink can assembly directly over the combustion can, allowing a small separation between the two cans.
5. Allow the pecan to burn as completely as possible. If the flame goes out during the process and the pecan is still mostly unburned, you must empty the water in the can and start over with fresh water and a new pecan half. After the pecan has burned out, allow the thermometer to reach the highest temperature. Record this temperature.
6. Carefully disassemble the apparatus, empty the water from the soft-drink can, and dispose of the burned food item.
7. Repeat steps 2-6 with two small marshmallows.
8. Repeat steps 2-6 with a small food item of your choice.

ANALYZE AND CONCLUDE

1. **Interpreting Data** Use your data to calculate the number of kilojoules of energy liberated per gram of each food item. Assume that the water in the soft-drink can has a density of 1.00 g/mL and that the specific heat of the water is 4.19 J/g°C.
2. **Comparing and Contrasting** Compare and rank the food items according to the amounts of energy liberated per gram.
3. **Inferring** Based upon the chemical compositions of the foods you tested (protein, carbohydrate, fat, etc.), what type of food provides the greatest amount of energy per gram?

APPLY AND ASSESS

1. Would it be desirable to eat mainly the type of food that provides the greatest amount of energy per unit mass? Explain.
2. What aspects of your procedure may have caused error in your calculated results? How would each of these sources of error affect your result?

DATA AND OBSERVATIONS

Food Item	Mass (g)	Initial H_2O Temperature (°C)	Final H_2O Temperature (°C)
Pecan			
Marshmallow			
Other food			

You can use calorimetry to measure the energy content of food. The food is burned rapidly in oxygen, rather than slowly, as in the body, but the amount of energy released is the same. The following sample problem shows you how to calculate food Calories from calorimetry data.

SAMPLE PROBLEM 2 Measuring Food Calories

A 1.00-g sample of nuts reacts with excess oxygen in a calorimeter. The calorimeter contains 1.00 kg of water that has an initial temperature of 15.40°C and a final temperature of 20.20°C. Find the energy content of the nuts. Express your answer in kJ/g and Calories/g.

Analyze

- You first need to find the heat change for the water, q_w, which is equal to $-q_{reaction}$ for the burning of the 1.00 g of nuts.

Set Up

- Find the temperature change of the water. Then, use the equation $q_w = (m)(\Delta T)(C_w)$ after finding the temperature change of the water. The negative of the value of q_w will equal $q_{reaction}$ in kilojoules. A conversion factor can then be used to convert to Calories.

Solve

- $$\Delta T = T_{final} - T_{initial}$$
  $$= 20.20°C - 15.40°C = 4.80°C$$
  $$q_w = (m)(\Delta T)(C_w)$$
  $$= \frac{1.00 \times 10^3\ \cancel{g} \,\big|\, 4.80\cancel{°C} \,\big|\, 4.184\ J}{\quad\big|\quad\big|\, \cancel{g}\cdot\cancel{°C}}$$
  $$= 2.01 \times 10^4\ J = 20.1\ kJ$$
  Now, apply a conversion factor to convert to Calories.
  $$= \frac{20.1\ kJ \,\big|\, 1\ Calorie}{\big|\, 4.184\ kJ} = 4.80\ Calories$$

Check

- Check to be sure that all units were correctly handled, and repeat the calculation to be sure that it was sound. The result does check out.

PRACTICE PROBLEMS

Supplemental Problems

For more practice with solving problems, see Supplemental Practice Problems, Appendix B.

4. A group of students decides to measure the energy content of certain foods. They heat 50.0 g of water in an aluminum can by burning a sample of the food beneath the can. When they use 1.00 g of popcorn as their test food, the temperature of the water rises by 24°C. Calculate the heat released by the popcorn, and express your answer in both kilojoules and Calories per gram of popcorn.

5. Another student comes along and tells the group in problem 4 that she has read the label on a popcorn bag that states that 30 g of popcorn yields 110 Calories. What is that value in Calories/gram? How can you account for the difference?

6. A 3.00-g sample of a new snack food is burned in a calorimeter. The 2.00 kg of surrounding water change in temperature from 25.0°C to 32.4°C. What is the food value in Calories per gram?

Energy Economics

Have you ever wondered why recycling has become so important? Part of the answer has to do with energy. For many materials, the energy required for recycling used objects is less than the energy of simply throwing the used objects away and making new ones from fresh raw materials.

Aluminum Production and Recycling

Aluminum is the metal of choice for soft-drink cans, as shown in **Figure 20.11.** It is lightweight, has a low specific heat (0.902 J/g°C), and conducts heat rapidly. As a result of its low specific heat, the aluminum itself absorbs little heat, so the beverage inside the container can be cooled quickly. In Chapter 17, you learned that aluminum is produced by electrolysis of its principal ore, bauxite. The production of aluminum requires large amounts of energy, particularly electrical energy. Therefore, one reason to recycle aluminum is to conserve that energy and reduce the cost of producing aluminum products.

Figure 20.11
Aluminum: A Useful Metal
The soft-drink cans and other objects shown here are all made from aluminum.

A comparison of the energy required to produce new aluminum cans and the energy required to produce cans from recycled aluminum reveals that 15-20 cans can be made from recycled aluminum for every one that can be made from new aluminum ore. Thus, the cost of using recycled aluminum is much less than that of using new aluminum. In addition to saving money on energy costs, recycling aluminum conserves valuable raw materials and uses less fossil fuel, which is the source of most of the energy required to produce the aluminum. To understand where the energy savings come from in reducing use of fossil fuels, it is important to look at the process used to change chemical energy into electrical energy.

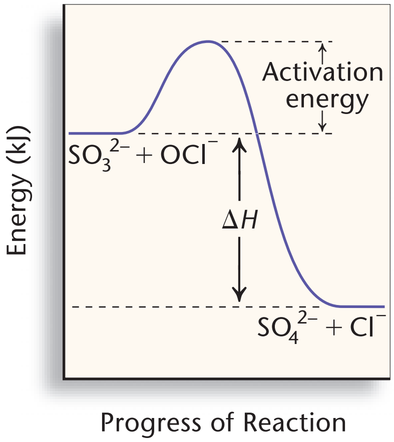
Fact of the MATTER

Recycling aluminum consumes only about seven percent of the energy required to make the same amount of new aluminum from ore. Recycling 25 aluminum cans saves an amount of energy equivalent to that available from burning a gallon of gasoline.

miniLAB 2

Heat In, Heat Out

If a chemical reaction is endothermic, the reverse reaction is exothermic. Examine the reaction between sulfite ion (SO_3^{2-}) and hypochlorite ion (OCl^-), which yields sulfate ion (SO_4^{2-}) and chloride ion (Cl^-).



Procedure

1. Pour approximately 40 mL of a commercial chlorine bleach solution into a 150-mL beaker. Put a thermometer into the solution and record the temperature. Be careful handling the bleach and avoid inhaling the fumes. Work in a well-ventilated room.
2. Pour approximately 40 mL of a 0.5*M* sodium sulfite solution into the beaker that contains the bleach, and stir gently for a few seconds.
3. Record the final temperature of the mixture.

Analysis

1. Based on the temperature change, was the reaction that occurred exothermic or endothermic? Would the reverse reaction be exothermic or endothermic? Refer to the graph for help.
2. Write the balanced equation for the reaction. Identify the oxidizing agent and the reducing agent in the reaction.

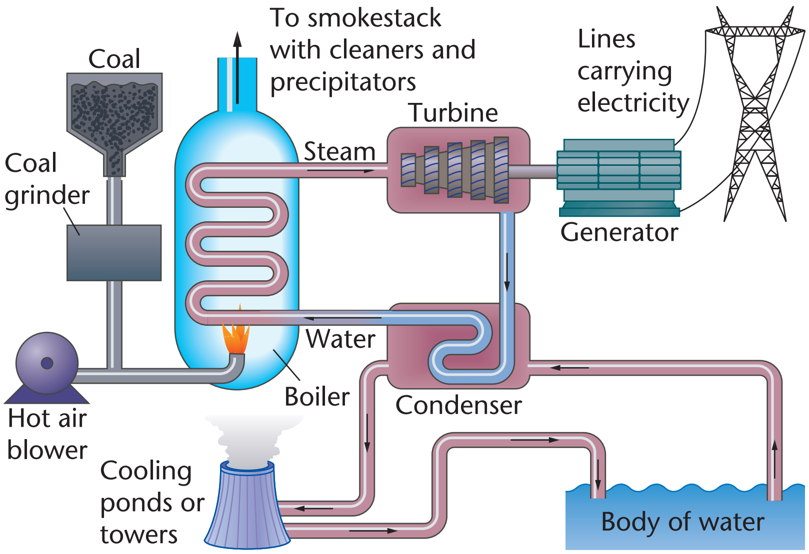
Converting Chemical Energy to Electricity

The most convenient form of energy available at the present time is electricity. It provides light, heat, hot water, and the ability to run machines of all kinds. Televisions, air conditioners, home appliances, and computers are but a few of the conveniences you get to enjoy because of electricity. In the United States, most electricity must be produced by burning fossil fuels, typically coal. **Figure 20.12** illustrates the production of electricity in a coal-fired power plant.

Figure 20.12

Energy Processes in a Coal-Fired Power Plant

In a typical electrical power plant, coal burns, producing heat that boils liquid water into steam. The kinetic energy of the moving steam drives a turbine, which drives a generator that produces the electrical energy.



EARTH SCIENCE CONNECTION

Bacterial Refining of Ores

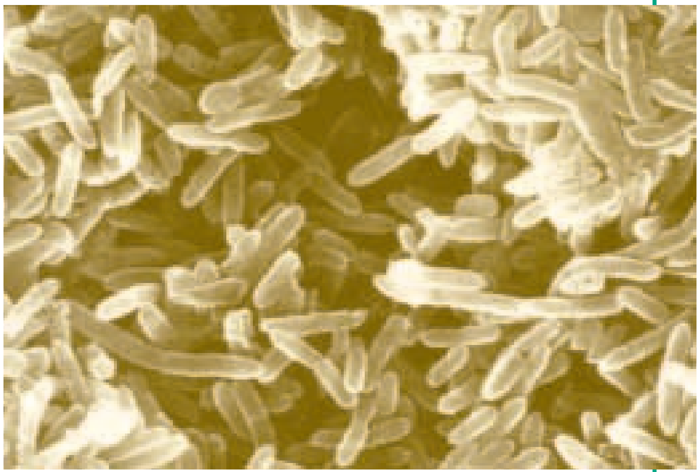
Roughly 2000 years ago, Roman miners began to investigate a blue liquid that they often noticed around piles of discarded, low-grade copper ore. They associated the blue color with copper, probably based on their experience with blue minerals, such as turquoise, which were often found in the vicinity of that metal. When they heated those minerals in a charcoal fire, they obtained copper.

The blue color of the liquid made them suspect that copper was present in it. They heated the liquid to see whether it really did contain copper. Their efforts yielded success. Pure copper was produced from copper salts in the liquid. What the Romans did not realize was that the copper in the liquid had been leached from the low-grade copper ore by bacteria.

Why bacteria are used to extract metals After crude ores are mined, they are crushed and then the metals are extracted by chemical treatment or by heating to high temperatures. Another way to extract metals from ore is to use bacteria to do part of the job. This method is less damaging to the environment, costs less in terms of energy input, and offers a way to use low-grade ores productively. Today, 24 percent of the copper produced is the result of bioprocessing by bacteria.

The bacterium *Thiobacillus ferrooxidans* obtains its energy by means of reactions involving various minerals. As a result, it produces acid and an oxidizing solution of Fe^{3+} ions, which can react with metals in crude ore.

The process is simple and requires little energy. Low-grade ore is first treated with sulfuric acid to stimulate the growth of bacteria. The microbes process the ore and release copper ions into solution. The metal is then extracted from the solution.



There's gold in them thar' bacteria Gold is another prospect for biorefining. As the sources of high-grade gold ore are disappearing, miners are being forced to mine low-grade ores. The gold in these lesser ores is present in the form of sulfides. To burn off the sulfur, the ores are traditionally treated by roasting or pressure oxidation. Only then can the gold be extracted with cyanide ions.

In one bio-oxidation process, low-grade gold ore is mixed with a brew of bacteria in huge, stirred reactors. A newer process places open piles of gold ore on an impermeable base. Bacterial cultures and the fertilizers needed to nurture them are poured onto the gold ore. In both methods, *T. ferrooxidans* treats the gold sulfides at a lower cost and with an efficiency of recovery increased from 70 to 95 percent.

Connecting to Chemistry

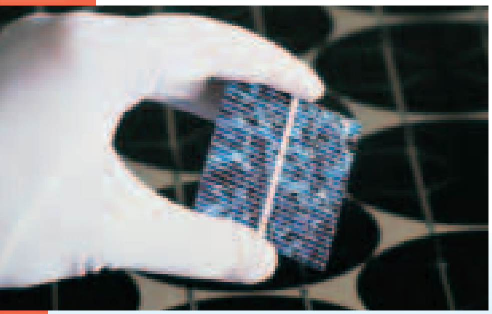
1. **Acquiring Information** Thermophilic bacteria, which thrive at temperatures around 100°C or higher, are being considered as possible candidates for biorefining. Find out what chemical advantage these bacteria have over other bacteria.
2. **Thinking Critically** What are the energy advantages of biorefining?

Alternative Energy Sources

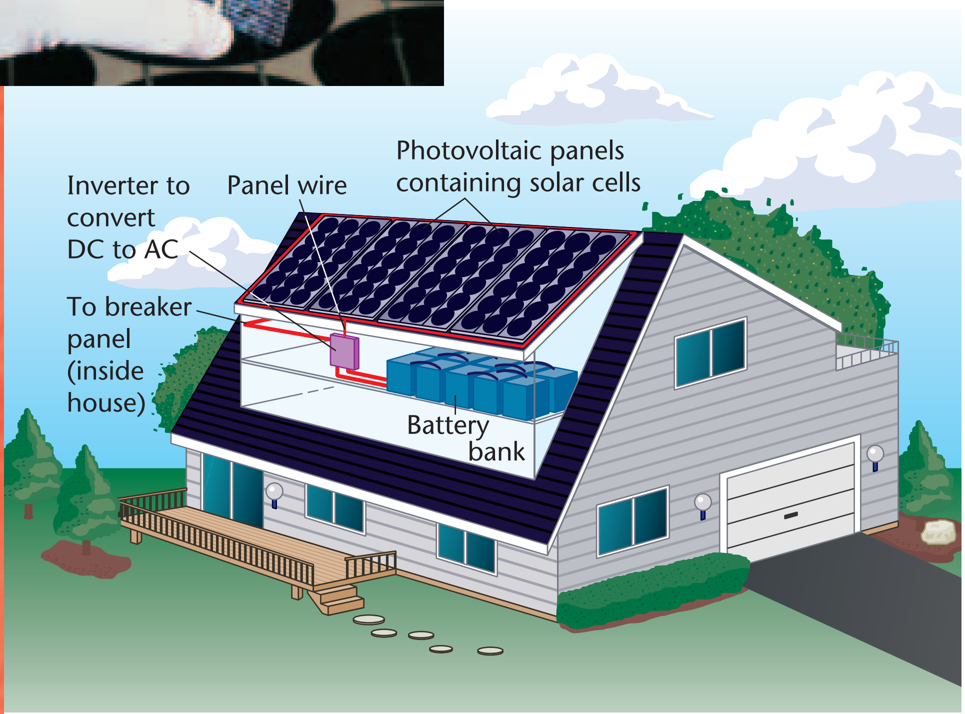
Energy can be neither created nor destroyed. Does that mean that people can keep using the same energy sources at the present rate forever? Fossil fuels are a finite source of energy that should be conserved. Using alternative, renewable sources of energy can prolong the supply of fossil fuels.

Solar Energy

One alternative source of energy is solar energy. The sun will be dispensing energy for the next 5 billion years. More than 300 000 homes in the United States are using solar energy to heat their living quarters. The solar home in the illustration was built to capture sunlight and convert it to heat that warms the rooms. Adobe walls, clay tile floors, triple windows, and heavily insulated walls and ceilings store the sun's warmth in winter. An overhang keeps the high-angle solar rays from entering the house in summer but does not block the lower-angle sun rays in winter.



A single photovoltaic cell



Photovoltaic cells use solar energy to produce electricity. These solar cells consist of layers of silicon with trace amounts of gallium or phosphorus and are capable of emitting electrons when struck by sunlight. Two main problems must be overcome before solar cells are a reliable source of electricity for everyday use. The first is the efficiency of conversion of solar energy to electrical energy. The second problem is finding efficient ways to store the electricity for use during the dark hours or during periods of cloudy weather.

Geothermal Energy

Magma, which is molten rock, can heat solid rock surrounding the magma chamber. When this occurs, water in the porous rock above the heated solid rock turns to steam. If there are cracks in the solid rock above, the steam escapes to the surface in the form of a geyser or hot spring. This is a natural source of geothermal energy. Many natural geothermal regions lie in the earthquake and volcano belts along Earth's crustal plate edges.

Geysers in Yellowstone National Park are an indication of the heat that is stored beneath the ground. This energy can be tapped and used to run power plants to produce electricity. The largest geothermal power plant in the world is at The Geysers in California. It generates 1000 megawatts of electric power—enough to serve the needs of 1 million people.

Natural geothermal energy is available in only a few spots around the world. To create new sites requires deep drilling. Although a geothermal power plant generates electricity at one-fourth the cost of power from a new nuclear plant and one-half the cost of power from a new coal plant, it is expensive to drill down to a magma chamber. Another disadvantage of these power plants is that they produce air pollution in the form of hydrogen sulfide, ammonia, and radioactive materials released from deep within Earth.

Power plant at The Geysers

Wind Energy

In some parts of the country, wind power is an almost unlimited source of energy. Wind can turn turbines that produce mechanical energy. A generator converts the mechanical energy into electrical energy. Usually, hundreds of wind turbines are needed to run a power plant.

Wind power emits no air pollutants. It requires no water for cooling. However, wind power can be used only in areas that have reliable winds. When the winds die down, a backup system for producing electricity must also be available. At present, wind energy is sometimes more expensive than other energy sources. Cheaper wind turbines could change this.

DISCUSSING THE TECHNOLOGY

1. **Hypothesizing** Which form of alternative energy might be feasible in your state? Explain.
2. **Acquiring Information** Investigate and describe how a solar furnace works.
3. **Thinking Critically** The three kinds of alternative energy sources discussed in this feature are not effective in all places. Despite that fact, how can they help solve the nation's energy problems?

Saving Energy with Catalysts

You have learned in Chapter 18 that polyethylene is a plastic used in many household products. It is made from ethylene, a simple hydrocarbon whose formula is C_2H_4. Ethylene is produced industrially by removing hydrogen from ethane (C_2H_6), another hydrocarbon, according to the following equation.

$$C_2H_6(g) + \text{heat} \rightarrow C_2H_4(g) + H_2(g)$$

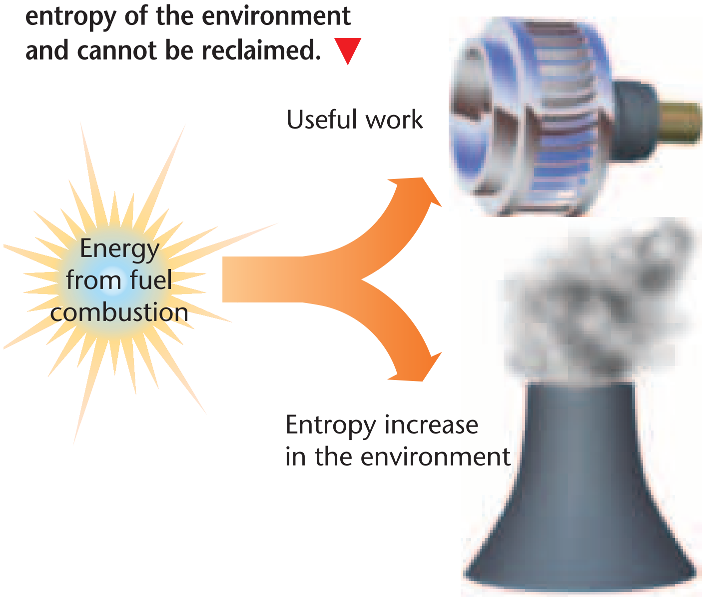
In theory, 1 mol of ethylene could be made with the addition of only 137 kJ of heat. In practice, the reaction above uses nearly four times that amount of energy to produce 1 mol of ethylene because of inefficiency in transferring energy. Over many years, chemists have introduced refinements into the production process. The development of better catalysts has lowered the energy requirements for this and other chemical processes. Research on energy-saving catalysts continues as energy becomes more expensive.

Entropy: One of the Costs of Using Energy

Any time energy is produced or is converted from one form to another, as it is in power plants and industrial processes, some of the energy is lost. This may seem surprising because of the law of conservation of energy. However, in this case, the energy is not destroyed but is simply not usable to do work. What happens to it? It is wasted as heat. In many processes, the amount of energy lost as heat exceeds the amount of energy that can be harnessed to do useful work.

The waste heat generated in systems that use or produce energy cannot be reclaimed, reused, or recycled because it has increased the random motion, and therefore the disorder, of molecules in the environment. As a result of this increased disorder, entropy increases as shown in **Figure 20.13.** The environment cannot spontaneously reorganize itself to give back the energy that increased its entropy. This increased disorganization of the environment is part of the price of using energy.

Figure 20.13
Using Energy Increases Entropy
Every time energy is used to do useful work such as spinning turbines to generate electricity, some energy is lost to the environment, usually as heat. This heat increases the entropy of the environment and cannot be reclaimed. ▼



In some cases, the heat is useful in situations where an increase in temperature is desirable, such as heating greenhouses in cold climates. ▼

A process, such as recycling, that saves energy by conserving electricity does more than reduce the total consumption of electricity. It also saves the chemical energy of original fossil fuel that would otherwise be lost forever as waste heat and increased entropy. This sometimes amounts to a savings of 70 percent of the fuel, making its energy available for some other use.

Energy and Efficiency

Because of entropy increases, no power or industrial plant, no matter how well designed, can completely convert heat from chemical reactions into useful work. Waste heat is inevitable. This kind of loss is easiest to understand in terms of efficiency, which is the amount of work that can be obtained from a process compared to the amount that must go into it.

For example, a modern fossil-fuel-burning plant has a maximum theoretical efficiency of only 63 percent. This means that, at most, 63 percent of the chemical energy from the original fuel can be converted to useful work. The remainder is lost as waste heat.

Other inefficiencies further reduce the percentage of energy that is converted to useful work. **Table 20.3** illustrates some of these factors in regard to a modern electrical power plant. The data in the table allow you to compute the plant's overall efficiency, which is simply the product of the efficiencies of all the individual steps multiplied by the maximum theoretical efficiency. If the maximum theoretical efficiency is 63 percent, then the efficiency of generating electricity by the process detailed in **Table 20.3** can be calculated as follows.

$$(0.63)(0.90)(0.75)(0.95)(0.90) = 0.36 = 36\%$$

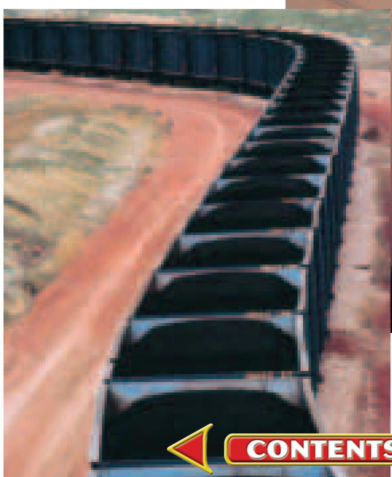
This low value is typical of the energy efficiency of power and industrial plants.

Fact of the Matter

At one time, scientists thought that heat was a colorless, odorless, weightless fluid. This fluid, which they called *caloric,* was believed to cause other substances to expand when it was added to them.

Table 20.3 Typical Efficiencies in Power Production

Maximum theoretical efficiency	63%
Efficiency of boiler	90%
Mechanical efficiency of turbine	75%
Efficiency of electrical generator	95%
Efficiency of power transmission	90%
Overall actual efficiency	36%



In a coal-fired power plant, a maximum of 36 percent of the energy available from coal is converted to electrical energy. ▶

Figure 20.14

Cooling Tower

A cooling tower is a common way to get rid of waste heat. Inside the tower, hot water is sprayed into the air while large fans draw air through the droplets. Although this process only transfers the heat to the atmosphere, it is an improvement over pumping the hot water back into streams, an action that killed organisms in the stream.

Improving the Efficiency of Industrial Processes

The efficiency of industrial processes can be increased and valuable energy resources can be conserved if plants that use improved energy-saving devices are designed and built. Also, if uses are found for waste heat—such as heating buildings and growing food in cold climates—energy can be conserved. Cooling towers, **Figure 20.14,** are one way of transferring waste heat. The development of more energy-efficient refrigerators, air conditioners, and lightbulbs can also improve the efficiency of electricity use.

Finally, the energy conversion processes of organisms can be studied and perhaps eventually applied to industrial methods. Biological processes convert energy from one form to another and at the same time create highly ordered—that is, low-entropy—systems. These natural processes are far more efficient than most industrial processes.

SECTION REVIEW

Understanding Concepts

1. Explain why it is impossible to convert 100 percent of the chemical energy in a fossil fuel to electrical energy.
2. The specific heat of aluminum is 0.902 J/g°C. The specific heat of copper is 0.389 J/g°C. If the same amount of heat is applied to equal masses of the two metals, which metal will increase more in temperature? Explain.
3. How much heat, in kilojoules, is given off by a chemical reaction that raises the temperature of 700 g of water in a calorimeter by 1.40°C?

Thinking Critically

4. **Using a Table** Using **Table 20.3** as a guide, determine the overall efficiency of a power plant that has a maximum theoretical efficiency of 50 percent, given that the other efficiencies are the same as those in the table.

Applying Chemistry

5. **Calorie Counting** Using **Table 20.2** as a guide, compare the Caloric content of 1 g of butter to 1 g of spaghetti. Which food gives you more Calories per gram?

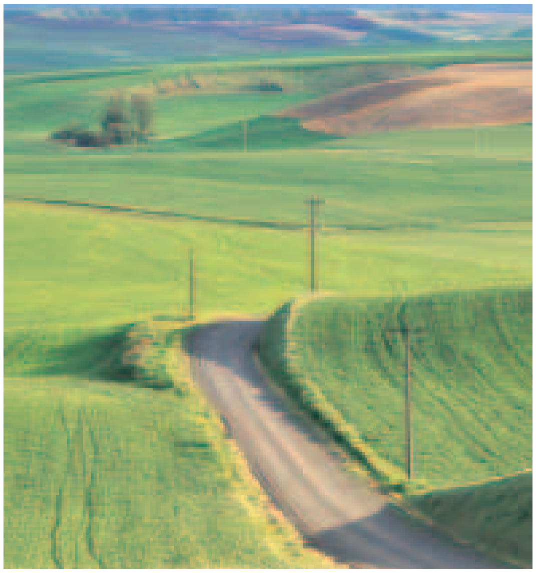
chemistryca.com/self_check_quiz

CONTENTS

SECTION 20.3

Photosynthesis

All organisms need energy to survive, and the main source of energy for Earth is the sun. Somehow, the sun's energy must be captured as chemical energy so that living things can use it. This process is called photosynthesis.



The Basis of Photosynthesis

Plants and other photosynthetic organisms have at least one obvious characteristic in common: they are green, the color of chlorophyll. The chlorophyll in cells of plants and algae is found in organelles called chloroplasts, as shown in **Figure 20.15.** The chlorophyll in photosynthetic bacteria is found on membranes spread throughout the cells. However, in both cases, the process works essentially the same way. Chlorophyll absorbs light and changes it to chemical energy, triggering a complex series of changes that together make up photosynthesis. Let's examine the chemical process in more detail. Note, as you read, that most of the changes absorb energy.

SECTION PREVIEW

Objectives

✓ Analyze the process and importance of photosynthesis.

✓ Compare the energy efficiency of photosynthesis and processes that produce electricity.

✓ Trace how energy from the sun passes through a food web.

Review Vocabulary

Calorie: heat required to raise the temperature of 1 gram of water by 1°C.

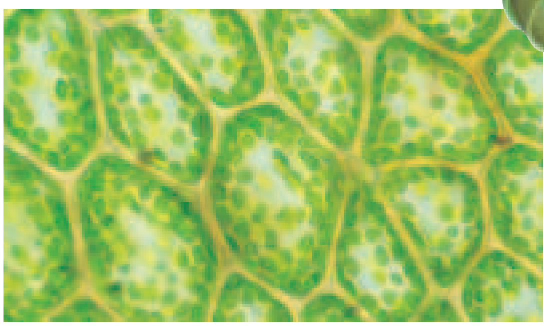
New Vocabulary

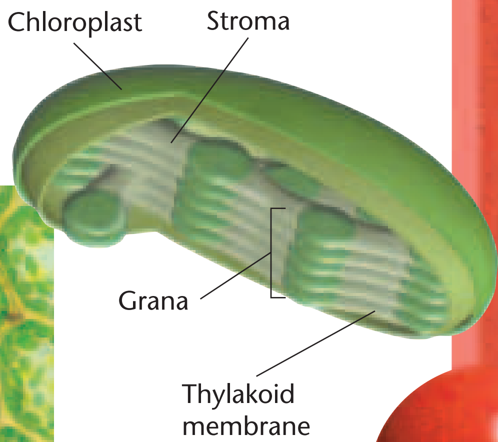
photosynthesis

Figure 20.15

Chloroplasts

The photo (left) shows some plant cells as seen through a microscope. The green structures inside are chloroplasts, which contain the light-absorbing green pigment chlorophyll. The diagram (right) shows a single chloroplast. Notice the grana, which are stacks of membranes on which the chlorophyll is located. The liquid surrounding the membranes is called stroma and is the place where sugar molecules are synthesized.





The Chemistry of Photosynthesis

The process of **photosynthesis** involves a series of reactions in which green plants and some other organisms manufacture carbohydrates from carbon dioxide and water using energy from sunlight. The net reaction that is carried out in photosynthesis is one that produces simple sugars and oxygen gas from carbon dioxide and water. The balanced equation is generally written as follows.

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

The formula $C_6H_{12}O_6$ represents glucose, a simple sugar.

Photosynthesis is divided into two basic series of reactions—the light reactions and the Calvin cycle.

WORD ORIGIN

chlorophyll:
chloros (GR) green
phyllon (GR) leaf
Chlorophyll is the substance that gives a green color to the leaves of a plant.

The Light Reactions

When light reaches a chlorophyll molecule, the energy of the light is absorbed. The energy excites electrons in the molecule, as described in **Figure 20.16.** A high-energy electron is released from the chlorophyll and transfers some of its energy to a molecule of ADP, adenosine diphosphate. That causes the ADP to bond to a third phosphate to form ATP, adenosine triphosphate. As you may recall from Chapter 19, ATP is the main energy storehouse in cells and also plays a critical role in respiration.

As you can see by following the process shown in **Figure 20.16,** electrons lost by the chlorophyll in the first step are restored to it by a reaction in which water is split into elemental oxygen and hydrogen ions. The oxygen is released into the atmosphere and can later be used by organisms, including you, for respiration.

Figure 20.16
The Light Reactions

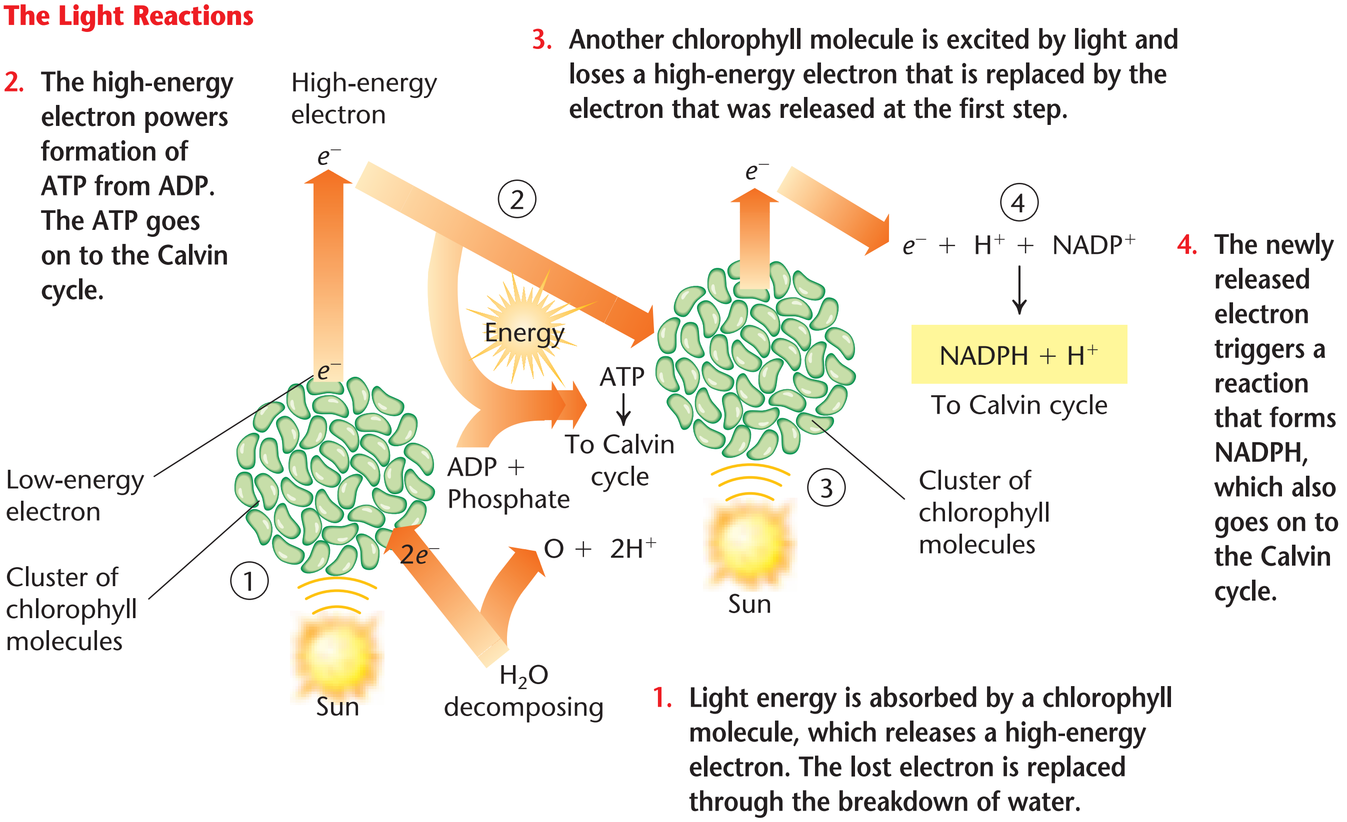
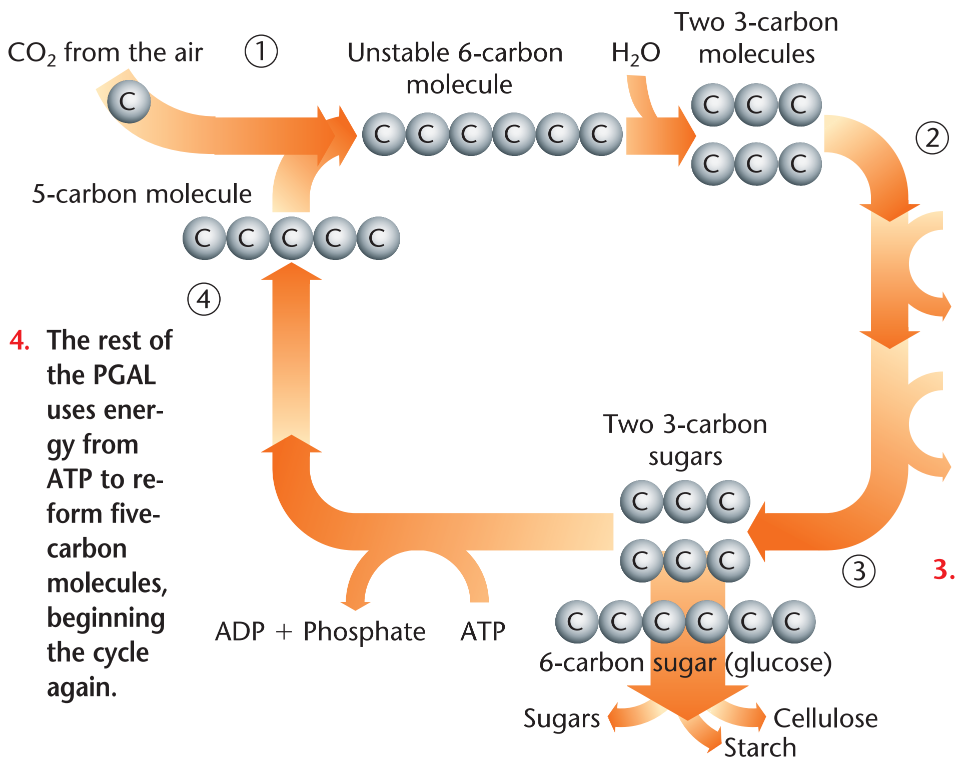


Figure 20.17
The Calvin Cycle

1. **In the Calvin cycle, atmospheric carbon dioxide reacts with a five-carbon molecule to form an unstable six-carbon molecule. That six-carbon molecule breaks down into two three-carbon molecules.**

2. **The two three-carbon molecules interact with ATP, NADPH, and H^+ from the light reactions, which provide energy and hydrogen atoms. As a result, the two three-carbon molecules are converted to two PGAL molecules, which are three-carbon sugars.**



4. **The rest of the PGAL uses energy from ATP to re-form five-carbon molecules, beginning the cycle again.**

3. **Some of the PGAL that is made in this way can combine to form the six-carbon sugar glucose. It can also be used directly as an energy source or to make fats or nucleic acids.**

The electron from the original chlorophyll has now lost its energy and moves on to join a different chlorophyll molecule—one that also absorbs light energy but is involved in a different reaction. The transferred electron replaces an electron that is energized by light absorption and released. This newly released electron is involved in another reaction, which forms an important coenzyme called NADPH. The NADPH is formed from NADP and from hydrogen ions that were freed in the earlier breakdown of water. Both the NADPH and the ATP formed as a result of the first absorption of energy continue to the second series of photosynthesis reactions, the Calvin cycle.

The Calvin Cycle

Once the light reactions have produced ATP and NADPH, the second series of photosynthesis reactions can occur. As shown in **Figure 20.17,** the Calvin cycle, which takes place in the stroma of chloroplasts, makes use of NADPH from the light reactions, as well as carbon dioxide taken in from the surroundings. Simple sugars, such as glucose, result.

In the Calvin cycle, ATP turns back into ADP and provides energy for a series of reactions. In the first of these, NADPH loses a hydrogen ion and becomes NADP. The hydrogen ion from the NADPH, together with other H^+ ions produced during the light reactions, provides hydrogen for the synthesis of the sugars. The carbon and oxygen for the sugars are provided mainly by the carbon dioxide.

Energy and the Role of Photosynthesis

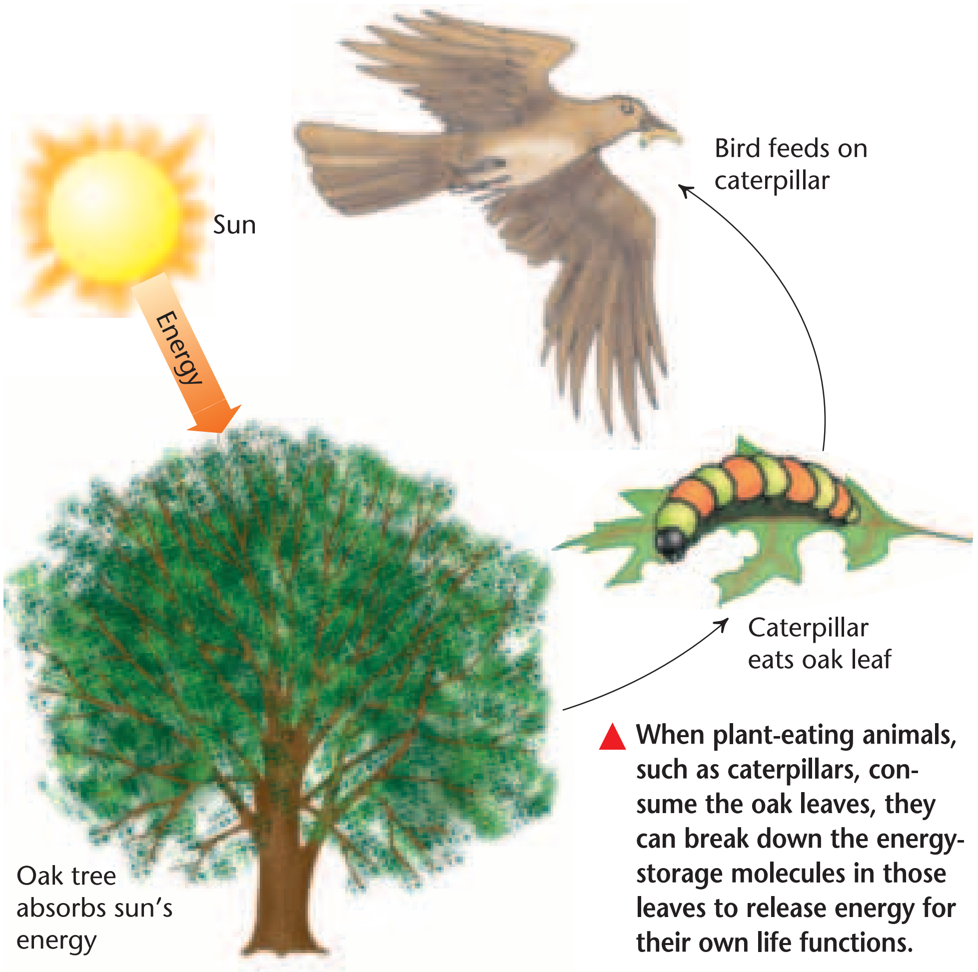
As you read earlier, most of the reactions that occur during photosynthesis are endothermic. The absorbed energy is stored by using it to synthesize high-energy molecules. Only chlorophyll-containing organisms can carry out photosynthesis and make these high-energy molecules that they need in order to survive. The animals that eat those organisms get the energy they need from them, as shown in **Figure 20.18.** So, photosynthesis is the ultimate process for chemically storing the energy that all organisms in the food web need. You could thus describe photosynthesis as the entry point of energy for life on Earth. See **Figure 20.19** on the next page.

The organization and maintenance of your body, like that of every organism, depend upon energy that is used to power chemical changes. However, energy is not the only factor involved when systems undergo changes. There is also entropy. Natural systems have a tendency toward increased disorder. However, living things are highly ordered, and many of the processes that take place in them, such as building more complex molecules, involve decreases in entropy rather than increases.

When meat-eating animals, such as some kinds of birds, consume the plant eaters, they make use of the energy that was stored in their prey.

Figure 20.18

Energy and the Food Web
Ultimately, the sun is the source of the energy, and photosynthesis is the process that captures it and makes all life possible.



The oak tree absorbs energy from the sun and stores it in the molecules it makes during photosynthesis. The stored energy allows it to carry out essential functions.

When plant-eating animals, such as caterpillars, consume the oak leaves, they can break down the energy-storage molecules in those leaves to release energy for their own life functions.

How is it possible, then, for organisms to overcome the tendency to disorder and maintain their complex structures and life functions? The secret of their success is their ability to absorb energy from outside sources, such as light and food. These outside sources continue pumping in the energy needed to keep the otherwise-nonspontaneous processes going.

When the body converts stored chemical energy to other uses during respiration, heat is produced. The heat, if it built up, could lead to destructive entropy increases and a breakdown of structures and functions. Some of the heat produced is radiated to the environment. The heat given off raises the entropy of the surroundings rather than that of the body. The low entropy of living things is thus maintained partially at the expense of the entropy of the surroundings.

In warm-blooded organisms, such as humans, some of the heat produced during respiration is retained by the body, which uses it to maintain a constant temperature. This temperature allows various biochemical processes to take place at the proper rates.

Supplemental Problems

For more practice with solving problems, see Supplemental Practice Problems, Appendix B.

Figure 20.19

Photosynthesis and Respiration in Balance

This British chemist spent 15 days in a sealed chamber where his oxygen was entirely supplied by 30 000 wheat plants carrying out photosynthesis. All carbon dioxide for photosynthesis came from the man's respiration. This 1995 experiment was the first in a series to determine whether humans could live in a similar but larger chamber on the moon or another planet.

SECTION REVIEW

Understanding Concepts

1. Plants appear green because their leaves reflect green and yellow light. Explain what happens to the energy in the wavelengths of light that are not reflected.
2. How does chlorophyll function in photosynthesis?
3. Explain why a meat-eating animal such as an eagle depends on photosynthesis for its energy.

Thinking Critically

4. **Inferring** What would be the effect on Earth's atmosphere if all plants were temporarily unable to carry out the process of photosynthesis?

Applying Chemistry

5. **Opposite Processes** Compare the process of photosynthesis with the process of respiration in Chapter 19. Explain why they are often considered opposite processes.

CHAPTER 20 ASSESSMENT

REVIEWING MAIN IDEAS

20.1 Energy Changes in Chemical Reactions

- Exothermic reactions give off energy; endothermic reactions absorb energy.
- Energy can be converted from one form to another, but cannot be created or destroyed.
- Activation energy is the energy needed to get a reaction started. A catalyst lowers the activation energy for a given reaction.
- Entropy is a measure of disorder. Spontaneous processes tend to proceed from a state of high energy to a state of low energy. They also tend to proceed from a state of less disorder to a state of greater disorder.

20.2 Measuring Energy Changes

- Chemical reactions are a source of energy because the energy of chemical bonds can be converted to heat, light, or electrical energy. Heat changes can be measured by calorimetry.
- Whenever energy is converted from one form to another, some energy is lost as heat. The lost energy generally cannot be used again or converted back to a useful form of energy.

20.3 Photosynthesis

- Photosynthesis is a process by which plants use light energy to manufacture carbohydrates from carbon dioxide and water. Because such substances are the primary source of energy for all organisms, the results of photosynthesis provide energy for life on Earth.
- The reactions that occur during photosynthesis include the light reactions and the Calvin cycle. Most of the reactions that occur during this process are endothermic.

Vocabulary

For each of the following terms, write a sentence that shows your understanding of its meaning.

calorie	kilocalorie
Calorie	law of conservation of energy
entropy	photosynthesis
fossil fuel	
heat	

UNDERSTANDING CONCEPTS

1. Explain how your car converts energy from one form to another.
2. How does the process by which plants obtain sugars differ from that by which most other organisms obtain sugars?
3. The efficiency of an automobile engine is improved by heating the interior of the car. Why is this the case?
4. Why is energy needed to sustain an endothermic reaction?
5. What is meant by *entropy?*
6. Suppose that a coal-fired power plant converts energy from coal into electrical energy with 36 percent efficiency. How many kilojoules of energy from coal are needed to produce 1.0 kJ of electrical energy?
7. Hydrogen is used as a fuel for the space shuttle because it provides more energy per gram than many other fuels. The combustion of hydrogen is described by the following equation.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$
$$\Delta H = -484 \text{ kJ}$$

a) Is the reaction exothermic or endothermic?
b) How much energy does complete combustion of 1.00 g of hydrogen provide?

8. What is the difference between a positive ΔH and a negative ΔH in terms of energy and the kind of reaction involved?
9. Why is outside energy needed to start most exothermic reactions but not to sustain them?
10. Why does it take more energy to decompose water than it does to boil water?
11. Describe the general trends of spontaneous reactions in terms of energy and entropy change.

APPLYING CONCEPTS

Chemistry and Technology

12. Explain why garbage could be an excellent fuel for the production of electricity.

How It Works

13. Another type of pack that can be used to supply heat to injuries is reusable. It contains a supersaturated solution of a salt and a disc of metal. When the metal disc is bent, the solute begins to crystallize and releases heat. The pack can be reset by heating it in boiling water, which causes the salt to dissolve again. How can you account for the heat given off by this kind of pack?

Everyday Chemistry

14. Write balanced equations for the following reactions, which occur in catalytic converters: a) the breakdown of nitrogen monoxide into oxygen and nitrogen, and b) the reaction of carbon monoxide with oxygen to produce carbon dioxide.

Earth Science Connection

15. The specific heat of lead is 0.128 J/g • °C. How much heat is required to raise the temperature of 10.0 g of lead from 20.0°C to 30.0°C? Compare this amount of heat to the amount needed to raise the temperature of 10.0 g of water by the same amount.
16. List two factors that would make an endothermic reaction more likely to occur.
17. Why will the butane gas in a disposable lighter fail to ignite if the flint is worn out and does not generate a spark?
18. Cite three common examples that illustrate the natural tendency for entropy to increase.
19. Heat from a burning match is necessary to ignite a candle. Why is it incorrect to say that the burning of a candle is, therefore, an endothermic reaction?

THINKING CRITICALLY

Interpreting Data

20. Account for the fact that there is an energy difference between the following reactions.

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
$\Delta H = -802$ kJ

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
$\Delta H = -890$ kJ

Applying Concepts

21. The kinetic model of matter and the concept of entropy both explain the process of diffusion, which is the spontaneous spreading of liquid or gas particles throughout a volume. Explain, in terms of both kinetics and disorder, why a drop of a dye placed into a beaker of water will eventually color the entire volume of water.

Making Predictions

22. The heats of reaction for the formation of the hydrogen halides from their elements are listed in the table below. On the basis of this information, predict which compound is most stable, in terms of not breaking down into its elements. Which is least stable? Explain.

Compound	ΔH (kJ/mol)
HF	−268
HCl	−92
HBr	−36
HI	+25

Interpreting Data

23. **ChemLab** Two different foods are burned in a calorimeter. Sample 1 has a mass of 6.0 g and releases 25 Calories of heat. Sample 2 has a

mass of 2.1 g and releases 9.0 Calories of heat. Which food releases more heat per gram?

Relating Cause and Effect

24. MiniLab 1 Which factor—entropy, added heat, or both—promotes the dissolution in water of a solid that spontaneously dissolves endothermically? Which factor promotes dissolution in water of a solid that spontaneously dissolves exothermically?

Inferring

25. MiniLab 2 A certain reaction proceeds in the forward direction with a ΔH of -16 kJ. Will the reverse reaction be endothermic or exothermic? What will be the value of ΔH for the reverse reaction?

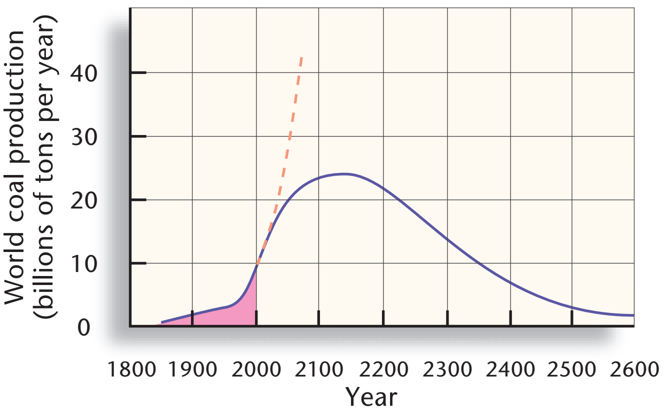
CUMULATIVE REVIEW

26. Distinguish between atomic number and mass number. How do each of these two numbers compare for isotopes of an element? (Chapter 2)

27. Why does the second period of the periodic table contain eight elements? (Chapter 7)

SKILL REVIEW

28. Using a Graph The graph below projects the world production of coal (solid line) until reserves are depleted and also projects the demand for coal (dashed line). Use the graph to answer the following questions.



a) How much longer can the present growth rate in coal consumption continue?

b) In what year will maximum coal production be reached, and how much coal will be produced that year?

c) Compare the amount of coal produced in 1900 to the production in the peak year.

29. Using a Data Table The chemical formulas, heats of combustion, and formula masses of three hydrocarbons are given in the following table.

Hydrocarbon	Heat of Combustion (kJ/mol)	Formula Mass (g/mol)
Methane, CH_4	−890	16
Butane, C_4H_{10}	−2859	58
Hexane, C_6H_{14}	−4163	86

a) According to the data, which hydrocarbon yields the greatest amount of energy per gram?

b) Using the data table, draw a graph of heat of combustion in kilojoules per mole versus formula weight. Use your graph to estimate the heat of combustion of propane (C_3H_8).

c) Conduct library research to find uses for these four hydrocarbons. Search out advertisements or magazine photos illustrating the uses. Prepare a display or bulletin-board presentation to illustrate your findings.

PROBLEM SOLVING

30. The compound B_5H_9 was once proposed as a rocket fuel because it has a high heat of combustion. In a combustion experiment, a scientist observes that 1.00 g of B_5H_9 burned in excess oxygen in a calorimeter raises the temperature of 800 g of water from 24.0°C to 44.3°C. How much heat is generated in the reaction?

31. An organic fuel is burned in a calorimeter. The reaction releases 194 kJ of heat. If the initial temperature of the water was 21.0°C, and the final temperature was 51.9°C, how much water was in the calorimeter?

Standardized Test Practice

Chemical Reaction	Heat of Reaction (kJ/mol)
$4Fe(s) + 3O_2(g) \rightarrow 2FeO_3(s)$	−1625
$NH_4NO_3(s) \rightarrow NH_4{+}(aq) + NO_3^-(aq)$	+27
$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$	−2808
$H_2O(l) \rightarrow H_2O(g)$	+40.7

Use the table above to answer questions 1–3.

1. The rusting of an iron nail is a(n)

a) decomposition reaction.
b) displacement reaction.
c) endothermic reaction.
d) exothermic reaction.

2. How much energy is required to boil 0.5 L water into steam?

a) 1.1 kJ
b) 20.35 kJ
c) 732.6 kJ
d) 1130.6 kJ

3. What is the heat of reaction for the production of 360 g of glucose during the process of photosynthesis ?

a) 2808 kJ
b) −2808 kJ
c) 5616 kJ
d) −5616 kJ

4. During a catalyzed reaction, the catalyst

a) provides additional energy to increase the rate of the reaction.
b) reduces the required activation energy to increase the rate of the reaction.
c) transforms an endothermic reaction into an exothermic reaction.
d) transforms an exothermic reaction into an endothermic reaction.

5. Which of the following is NOT an example of entropy?

a) A perfume bottle is opened releasing evaporated alcohol molecules that carry the perfume scents to every corner of a room.
b) Solar energy bathes the Earth with light and heat providing energy for photosynthesizing organisms to convert water and carbon dioxide into glucose.
c) Over thousands of years, ocean waves erode rock and coral into small particles that are carried and deposited to form beaches on shorelines.
d) A large pile of snow resting on an asphalt surface experiences a drastic increase in temperature and melts into water which runs off the surface into the soil.

6. The specific heat of ethanol is 2.44 J/g °C. How many kilojoules of energy are required to heat 50.0 g of ethanol from −20.0 C to 68.0 °C?

a) 10.7 kJ
b) 8.30 kJ
c) 2.44 kJ
d) 1.22 kJ

Unknown Metal Characteristics
Mass of metal = 4.68 g
Quantity of heat absorbed (q) = 256 J
Change in Temperature = 182 °C

Use the information in the table above to answer question 7.

7. Calculate the specific heat for the unknown metal.

a) 0.15 J/g °C
b) 0.301 J/g °C
c) 3.3 J/g °C
d) 6.58 J/g °C

Test Taking Tip

Your Mistakes Can Teach You The mistakes you make before the test are helpful because they show you the areas in which you need more work.