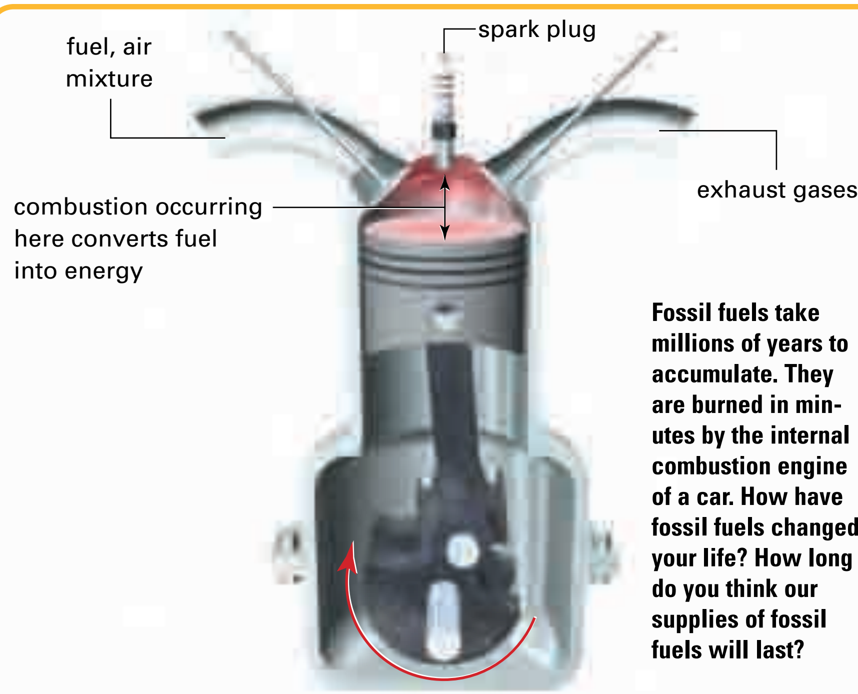
CHAPTER 14

Energy Trapped in Hydrocarbons

The whole world runs on energy—and so do you! Fossil fuels provide energy to power cars and heat buildings. Food provides energy to keep your body alive and warm. Both sources of energy come from organic compounds, such as hydrocarbons, sugars, and proteins.

Green plants, algae, and plankton trap the Sun's energy through the process of photosynthesis. After these organisms die, they are broken down by natural processes. Their remains accumulate on Earth's surface. In some areas, these remains build up in thick layers, which are eventually covered by rock and soil. Under certain conditions, over billions of years, pressure changes these layers into something new: fossil fuels. **Fossil fuels** (such as coal, natural gas, and petroleum) are fuels that are made from fossilized organic materials. The trapped energy from the Sun is still present in fossil fuels. To use this energy, we need to extract it. **Combustion**, or burning, is the most common way to extract energy from fossil fuels.

In this chapter, you will explore the ways in which our society obtains energy from fossil fuels. You will get a chance to measure exactly how much energy is obtained from an organic substance by doing your own combustion reaction. As well, you will learn how dangerous incomplete combustion reactions can be.



Fossil fuels take millions of years to accumulate. They are burned in minutes by the internal combustion engine of a car. How have fossil fuels changed your life? How long do you think our supplies of fossil fuels will last?

Chapter Preview

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- explaining bonding in molecular compounds (Chapter 3, section 3.3)
- writing and balancing chemical equations for different reactions (Chapter 4, section 4.1)
- significant digits (Chapter 1, section 1.2)
- problem solving in gas laws (Chapter 12, section 12.3)
- naming and drawing aliphatic compounds (Chapter 13, section 13.2)

14.1 Formation and Combustion Reactions

Section Preview/ Specific Expectations

In this section, you will

- **write** balanced chemical equations for the complete and incomplete combustion of hydrocarbons
- **perform** an experiment to produce and burn a hydrocarbon
- **recognize** the importance of hydrocarbons as fuels and as precursors for the production of petrochemicals
- **identify** the risks and benefits of the uses of hydrocarbons, for society and the environment
- **communicate** your understanding of the following terms: *complete combustion*, *incomplete combustion*

One of our most common uses of hydrocarbons is as fuel. (See Figure 14.1.) The combustion of fossil fuels gives us the energy we need to travel and to keep warm in cold climates. Fossil fuel combustion is also an important source of energy in the construction and manufacturing industries. As well, many power plants burn natural gas when generating electricity. Even goldsmiths use hydrocarbons, such as butane, as a heat source when crafting gold jewellery. At home, we often burn fossil fuels, such as natural gas, to cook our food.

How do we get energy from these compounds? In this section, you will learn how complete and incomplete combustion can be expressed as chemical equations. Combustion in the presence of oxygen is a *chemical* property of all hydrocarbons. (In Chapter 13, you learned about some *physical* properties of hydrocarbons, such as boiling point and solubility.)

Figure 14.1 How are fossil fuels being used in these photographs?

Chemistry Bulletin

Science | Technology | Society | Environment

Lamp Oil and the Petroleum Age

Abraham Gesner was born in 1797 near Cornwallis, Nova Scotia. Although Gesner became a medical doctor, he was much more interested in fossils. Gesner was fascinated by hydrocarbon substances, such as coal, asphaltum (asphalt), and bitumen. These substances were formed long ago from fossilized plants, algae, fish, and animals.

When Gesner was a young man, the main light sources available were fire, candles, and whale oil lamps. Gesner had made several trips to Trinidad. He began to experiment with asphaltum, a semisolid hydrocarbon from Trinidad's famous "pitch lake." In 1846, while giving a lecture in Prince Edward Island, he startled his audience by lighting a lamp that was filled with a fuel he had distilled from asphaltum. Gesner's lamp fuel gave more light and produced less smoke than any other lamp fuel the audience had ever seen used.

Gesner needed a more easily obtainable raw material to make his new lamp fuel. He tried a solid, black, coal-like bitumen from Albert County, New Brunswick. This substance, called albertite, worked better than any other substance that Gesner had tested.

Making Kerosene

One residue from Gesner's distillation process was a type of wax. Therefore, he called his lamp fuel *kerosolain*, from the Greek word for "wax oil." He soon shortened the name to *kerosene*. To produce kerosene, Gesner heated chunks of albertite in a retort (a distilling vessel with a long downward-bending neck). As the albertite was heated, it gave off vapours. The vapours passed into the neck of the retort, condensed into liquids, and trickled down into a holding tank. Once Gesner had finished the first distillation, he let the tank's contents stand for several hours. This allowed water and solid to settle to the bottom. Then he drew off the oil that remained on top.

Gesner distilled this oil again, and then treated it with sulfuric acid and calcium oxide. Finally he distilled the oil once more.

By 1853, Gesner had perfected his process. In New York, he helped to start the North American Kerosene Gas Light Company. Gesner distinguished between three grades of kerosene: grades A, B, and C. Grade C, he said, was the best lamp oil. Grades A and B could also be burned in lamps, but they were dangerous because they could cause explosions and fires.

Although Gesner never knew, his grades A and B kerosene became even more useful than the purer grade C. These grades were later produced from crude oil, or petroleum, and given a new name: gasoline!

Gesner laid the groundwork for the entire petroleum industry. All the basics of later petroleum refining can be found in his technology.

Making Connections

1. In the early nineteenth century, whales were hunted extensively for their oil, which was used mainly as lamp fuel. When kerosene became widely available, the demand for whale oil decreased. Find out what effect this had on whalers and whales.
2. How do you think the introduction of kerosene as a lamp oil changed people's lives at the time? What conclusions can you draw about the possible impact of technology?

Figure 14.2 Sour gas, $H_2S_{(g)}$, is sometimes "flared off" (burned) from an oil well. This combustion produces sulfur dioxide gas, which reacts with the water in the atmosphere to produce acid rain. Oil companies are now making an effort to reduce this type of pollution.

Complete and Incomplete Combustion

During a typical combustion reaction, an element or a compound reacts with oxygen to produce oxides of the element (or elements) found in the compound. Figure 14.2 shows an example of a combustion reactions.

Hydrocarbon compounds will burn in the presence of air to produce oxides. This is a chemical property of all hydrocarbons. **Complete combustion** occurs if enough oxygen is present. A hydrocarbon that undergoes complete combustion produces carbon dioxide and water vapour. The following equation shows the complete combustion of propane. (See also Figure 14.3.)

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

If you burn a fuel, such as propane, in a barbecue, you want complete combustion to occur. Complete combustion ensures that you are getting maximum efficiency from the barbecue. More importantly, toxic gases can result from **incomplete combustion**: combustion that occurs when not enough oxygen is present. During incomplete combustion, other products (besides carbon dioxide and water) can form. The equation below shows the incomplete combustion of propane. Note that unburned carbon, $C_{(s)}$, and carbon monoxide, $CO_{(g)}$, are produced as well as carbon dioxide and water.

$$2C_3H_{8(g)} + 7O_{2(g)} \rightarrow 2C_{(s)} + 2CO_{(g)} + 2CO_{2(g)} + 8H_2O_{(g)}$$

Figure 14.4 shows another example of incomplete combustion.
Go back to the equation for the complete combustion of propane. Notice that the mole ratio of oxygen to propane for the complete combustion (5 mol oxygen to 1 mol propane) is higher than the mole ratio for the incomplete combustion (7 mol oxygen to 2 mol propane, or 3.5 mol oxygen to 1 mol propane). These ratios show that the complete combustion of propane used up more oxygen than the incomplete combustion. In fact, the incomplete combustion probably occurred because not enough oxygen was present. You just learned that incomplete combustion produces poisonous carbon monoxide. This is why you should never operate a gas barbecue or gas heater indoors, where there is less oxygen available. This is also why you should make sure that any natural gas or oil-burning furnaces and appliances in your home are working at peak efficiency, to reduce the risk of incomplete combustion. Carbon monoxide detectors are a good safeguard. They warn you if there is dangerous carbon monoxide in your home, due to incomplete combustion.

Figure 14.4 The yellow flame of this candle indicates that incomplete combustion is occurring. Carbon, $C_{(s)}$, emits light energy in the yellow wavelength region of the visible spectrum.

Balancing Combustion Equations

Have you ever seen a construction worker using an oxyacetylene torch? (See Figure 14.6.) A brilliant white light comes from the torch as it cuts through steel. The intense heat that is associated with this flame comes from the combustion of ethyne, a very common alkyne. Ethyne is also known as *acetylene*.

Figure 14.3 Propane burning in a propane torch: A blue flame indicates that complete combustion is occurring.

How do you write the balanced equation for the complete combustion of acetylene (ethyne)? Complete hydrocarbon combustion reactions follow a general format:

$$\text{hydrocarbon} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water vapour}$$

You can use this general format for the complete combustion of any hydrocarbon, no matter how large or how small. For example, both acetylene and propane burn completely to give carbon dioxide and water vapour. Each hydrocarbon, however, produces different amounts, or mole ratios, of carbon dioxide and water.

You have seen, written, and balanced several types of reaction equations so far in this textbook. In the following sample problem, you will learn an easy way to write and balance hydrocarbon combustion equations.

Sample Problem

Complete Combustion of Acetylene

Problem

Write the balanced equation for the complete combustion of acetylene (ethyne).

What Is Required?

You need to write the equation. Then you need to balance the atoms of the reactants and the products.

What Is Given?

You know that acetylene (ethyne) and oxygen are the reactants. Since the reaction is a complete combustion reaction, carbon dioxide and water vapour are the products.

Plan Your Strategy

Step 1 Write the equation.

Step 2 Balance the carbon atoms first.

Step 3 Balance the hydrogen atoms next.

Step 4 Balance the oxygen atoms last.

Act on Your Strategy

Step 1 Write the chemical formulas and states for the reactants and products.

$$\text{ethyne} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water vapour}$$

$$C_2H_{2(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

Step 2 Balance the carbon atoms first.

$$\underset{\text{(2 carbons)}}{C_2H_{2(g)}} + O_{2(g)} \rightarrow \underset{\text{(2 carbons)}}{2CO_{2(g)}} + H_2O_{(g)}$$

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Step 3 Balance the hydrogen atoms next.

$$C_2H_{2(g)} + O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(g)}$$

(2 hydrogens) (2 hydrogens)

Step 4 Balance the oxygen atoms last.

The product coefficients are now set. Therefore, count the total number of oxygen atoms on the product side. Then place an appropriate coefficient in front of the reactant oxygen.

$$C_2H_{2(g)} + ?O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(g)}$$

(4 + 1 oxygens)

You end up with an *odd* number of oxygen atoms on the product side of the equation. When this happens, use a fractional coefficient so that the reactant oxygen balances. Here you have

$$C_2H_{2(g)} + \frac{5}{2}O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(g)}$$

You may prefer to balance the equation with whole numbers. If so, multiply everything by a factor that is equivalent to the denominator of the fraction. Since the fractional coefficient of $O_{2(g)}$ has a 2 in the denominator, multiply *all* the coefficients by 2 to get whole number coefficients.

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$$

Check Your Solution

The same number of carbon atoms appear on both sides of the equation.

The same number of hydrogen atoms appear on both sides of the equation.

The same number of oxygen atoms appear on both sides of the equation.

Sample Problem

Incomplete Combustion of 2,2,4-Trimethylpentane

Problem

2,2,4-trimethylpentane is a major component of gasoline. Write one possible equation for the incomplete combustion of 2,2,4-trimethylpentane.

What Is Required?

You need to write the equation for the incomplete combustion of 2,2,4-trimethylpentane. Then you need to balance the atoms of the reactants and the products. For an incomplete combustion reaction, more than one balanced equation is possible.

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What is given?

You know that 2,2,4-trimethylpentane and oxygen are the reactants. Since the reaction is an incomplete combustion reaction, the products are unburned carbon, carbon monoxide, carbon dioxide, and water vapour.

Plan Your Strategy

Draw the structural diagram for 2,2,4-trimethylpentane to find out how many hydrogen and oxygen atoms it has. Then write the equation and balance the atoms. There are many carbon-containing products but only one hydrogen-containing product, water. Therefore, you need to balance the hydrogen atoms first. Next balance the carbon atoms, and finally the oxygen atoms.

Act on Your Strategy

$$\begin{array}{c}\begin{array}{ccccccccc} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ CH_3 & — & C & — & CH_2 & — & CH & — & CH_3 \\ & & | & & & & & & \\ & & CH_3 & & & & & & \end{array} + O_2 \rightarrow C + CO + CO_2 + H_2O \\ (C_8H_{18})\end{array}$$

Count the carbon and hydrogen atoms, and balance the equation. Different coefficients are possible for the carbon-containing product molecules. Follow the steps you learned in the previous sample problem to obtain the balanced equation shown below.

$$\begin{array}{ccccccccc} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ CH_3 & — & C & — & CH_2 & — & CH & — & CH_3 \\ & & | & & & & & & \\ & & CH_3 & & & & & & \end{array} + \frac{15}{2}O_2 \rightarrow 4C + 2CO + 2CO_2 + 9H_2O$$

or

$$2\left(\begin{array}{ccccccccc} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ CH_3 & — & C & — & CH_2 & — & CH & — & CH_3 \\ & & | & & & & & & \\ & & CH_3 & & & & & & \end{array}\right) + 15O_2 \rightarrow 8C + 4CO + 4CO_2 + 18H_2O$$

Check Your Solution

The same number of carbon atoms appear on both sides of the equation.

The same number of hydrogen atoms appear on both sides of the equation.

The same number of oxygen atoms appear on both sides of the equation.

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Practice Problems

1. The following equation shows the combustion of 3-ethyl-2,5-dimethylheptane:

$$C_{11}H_{24} + 17O_2 \rightarrow 11CO_2 + 12H_2O$$

   (a) Does this equation show *complete* or *incomplete* combustion?

   (b) Draw the structural formula for 3-ethyl-2,5-dimethylheptane.

2. (a) Write a balanced equation for the complete combustion of pentane, C_5H_{12}.

   (b) Write a balanced equation for the complete combustion of octane, C_8H_{18}.

   (c) Write two possible balanced equations for the incomplete combustion of ethane, C_2H_6

3. (a) The flame of a butane lighter is usually yellow, indicating incomplete combustion of the gas. Write a balanced chemical equation to represent the incomplete combustion of butane in a butane lighter. Use the condensed structural formula for butane.

   (b) If you supplied enough oxygen, the butane would burn with a blue flame. Write a balanced chemical equation for the complete combustion of butane.

4. The paraffin wax in a candle burns with a yellow flame. If it had sufficient oxygen to burn with a blue flame, it would burn rapidly and release a lot of energy. It might even be dangerous! Write the balanced chemical equation for the complete combustion of candle wax, $C_{25}H_{52(s)}$.

5. 4-propyldecane burns to give solid carbon, water vapour, carbon monoxide, and carbon dioxide.

   (a) Draw the structural formula for 4-propyldecane.

   (b) Write two different balanced equations for the reaction described in this problem.

   (c) Name the type of combustion. Explain.

Large quantities of acetylene are produced each year by an inexpensive process that combines calcium carbide and water. In the next investigation, you will use this process to produce your own acetylene.

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Communicating results

Investigation 14-A

The Formation and Combustion of Acetylene

In this investigation, you will produce acetylene (ethyne) gas by mixing solid calcium carbide with water.

$$CaC_{2(s)} + 2H_2O_{(\ell)} \rightarrow C_2H_{2(g)} + Ca(OH)_{2(s)}$$

Then you will combine the acetylene with different quantities of air to determine the best reaction ratio for complete combustion. Make sure that you follow all the safety precautions given in this investigation and by your teacher.

Question

What is the ideal ratio of fuel to air for the complete combustion of acetylene (ethyne) gas?

Prediction

Air contains 20% oxygen, $O_{2(g)}$. How much air do you think is needed for the complete combustion of acetylene gas? Predict which proportion will react best:

- $\frac{1}{2}$ acetylene to $\frac{1}{2}$ air
- $\frac{1}{3}$ acetylene to $\frac{2}{3}$ air
- $\frac{1}{5}$ acetylene to $\frac{4}{5}$ air
- $\frac{1}{10}$ acetylene to $\frac{9}{10}$ air.

Safety Precautions

- Be careful of the flames from Bunsen burners. Check that there are no flammable solvents close by. If your hair is long, tie it back. Confine loose clothing.

Materials

4 test tubes (100 mL)
4 rubber stoppers
grease pencil
ruler
400 mL beaker
tweezers
matches (or Bunsen burner and splints)
1 or 2 calcium carbide chips
phenolphthalein indicator
limewater
medicine dropper
distilled water
test tube tongs

Procedure

1. Make a table to record your observations. Give your table a title.
2. Mark each test tube with a grease pencil to indicate one of the following volumes: $\frac{1}{2}$, $\frac{1}{3}$, $\frac{1}{5}$, and $\frac{1}{10}$. To find out where to mark the test tube, measure the total length of the test tube with a ruler. Then multiply the length by the appropriate fraction. Measure the fraction from the *bottom* of the test tube.
3. Fill the four test tubes completely with distilled water.
4. Invert the four test tubes in a 400 mL beaker, half full of distilled water. Make sure that the test tubes stay completely full.
5. Add three to five drops of phenolphthalein to the water in the beaker.
6. Using tweezers, drop a small chip of calcium carbide into the water. CAUTION Do not touch calcium carbide with your hands!
7. Capture the gas that is produced by holding the test tube marked $\frac{1}{2}$ over the calcium carbide chip. Fill the test tube to the $\frac{1}{2}$ mark with the gas. Remove the test tube from the beaker, still inverted. Let the water drain out. Air will replace the water and mix with the gas in the test tube. Insert a rubber stopper. Invert the test tube a few times to mix the acetylene gas with the air in the test tube.

8. Repeat step 7 with the other three test tubes. Fill each test tube to the volume that you marked on it: $\frac{1}{3}$, $\frac{1}{5}$, or $\frac{1}{10}$. After filling each test tube, remove it from the water and insert a rubber stopper.

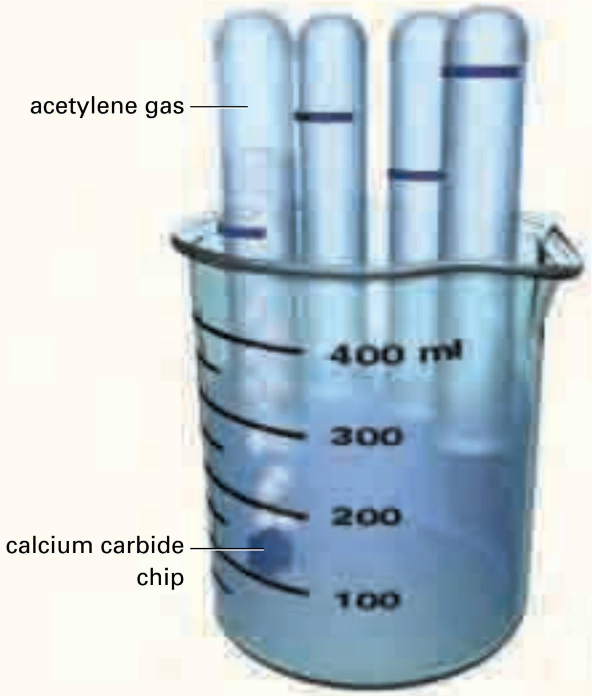


Figure 14.5

9. Invert the first test tube ($\frac{1}{2}$ full). Light a match or a splint. Using test tube tongs, hold the test tube inverted, and take out the stopper. Ignite the gas in the test tube by holding the lighted match or splint near the mouth of the test tube. **CAUTION** If you are using a Bunsen burner and wooden splints, take appropriate safety precautions. Extinguish burning splints by immersing them in water. Be aware of the Bunsen burner flame. Make sure that long hair is tied back and loose clothing is confined.

10. *Immediately* after the reaction in the test tube occurs, use the medicine dropper to add about 1 mL of limewater to the test tube. Stopper the test tube and shake it. **CAUTION** The mouth of the test tube may be hot.

11. Record your observations of the gas when it was ignited. Record your observations of what happened when you added the limewater. Describe any residue left on the test tube.

12. Repeat steps 9 to 11 with the other test tubes in order: $\frac{1}{3}$, $\frac{1}{5}$, and then $\frac{1}{10}$ full.

13. Dispose of all chemical materials as instructed by your teacher.

Analysis

1. What happened to the phenolphthalein indicator during the production of the gas? Explain your observation. **Note**: phenolphthalein is an acid/base indicator.

2. What products may have formed during the combustion of the gas? Support your answers with experimental evidence. **Note**: Limewater reacts with carbon dioxide to produce a milky white solid.

Conclusions

3. (a) Write a balanced chemical equation for the incomplete combustion of acetylene gas.
   (b) Write a balanced chemical equation for the complete combustion of acetylene gas.

4. The air that we breathe is approximately 20% oxygen. Think about the reaction you just wrote for the complete combustion of acetylene. Which ratio in this investigation ($\frac{1}{2}$, $\frac{1}{3}$, $\frac{1}{5}$, or $\frac{1}{10}$) allowed the closest amount of oxygen needed for complete combustion? Support your answer with calculations. Do the observations you made support your answer? Explain.

Applications

5. An automobile engine requires a carburetor or fuel injector to mix the fuel with air. The fuel and air must be mixed in a particular ratio to achieve maximum efficiency in the combustion of the fuel. What might happen if the fuel and air mixture is too rich (if there is too much fuel)?

6. What does the limewater test indicate? Write the balanced chemical equation for the limewater test. **Hint**: Limewater is a dilute solution of calcium hydroxide. A carbonate forms.

In this section, you were introduced to the complete and incomplete combustion of hydrocarbons. You learned that the complete combustion of a hydrocarbon produces water and carbon dioxide. You also learned that the incomplete combustion of a hydrocarbon produces additional products, such as unburned carbon and dangerous carbon monoxide. In the investigation, you had the chance to make and combust a hydrocarbon.

In the next section, you will learn about an important factor of combustion reactions: energy. The combustion of hydrocarbons produces a large amount of energy. This is why they are so useful as fuels. How can you include energy as part of a combustion or other equation? How can you calculate the energy released by fossil fuels? You will learn the answers to these questions in the rest of this chapter.

Section Review

1 K/U What is the difference between incomplete and complete hydrocarbon combustion reactions?

2 K/U Explain why you would usually write (g) for the state of the product water in these combustion reactions. When might you identify the state of water as liquid?

3 (a) K/U Write the balanced equation for the complete combustion of heptane, C_7H_{16}.

(b) K/U Write a balanced equation for the incomplete combustion of 1-pentene, C_5H_{10}.

4 K/U Natural gas is mainly methane gas. If you have a natural gas furnace, stove, or water heater in your home, you must ensure that these appliances are always running at peak efficiency. In other words, the methane gas should undergo complete combustion so that carbon monoxide is not produced. Write a balanced chemical equation to show the complete combustion of methane gas. Write a second balanced equation to show the incomplete combustion of methane gas.

5 C All hydrocarbons have the chemical property of combustion in the presence of oxygen. How do the complete combustion reactions of methane, ethane, and propane differ? How are they similar? **Hint:** Compare the balanced equations for the complete combustion of each.

6 I The complete combustion of ethane is given by the following unbalanced equation:

$$C_2H_{6(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

(a) Balance this equation.

(b) If one mole of ethane is combusted, how many grams of water vapour are produced?

(c) Assume that air contains 20% oxygen gas. What volume of air at STP is needed for the complete combustion of one mole of ethane?

14.2 Thermochemical Equations

Section Preview/ Specific Expectations

In this section, you will

- **write** thermochemical equations
- **relate** bond breaking and bond making to endothermic and exothermic energy changes
- **communicate** your understanding of the following terms: *exothermic*, *thermochemical equation*, *endothermic*, *bond energy*

In the last section, you learned about acetylene, an important fuel in our society (Figure 14.6). You balanced the equation for the complete combustion of acetylene. You then produced acetylene in an investigation. What you did not consider, however, was the most useful product that our society gets from acetylene: heat energy. How can you represent the heat that is released during combustion as part of a chemical equation?

For now, you will use the word "energy" to represent the heat in an equation. In the next section, you will calculate numerical values for this energy and use these energy values in chemical equations.

In previous science courses, you studied reactions that involve a change in energy. Figure 14.7 reviews some important terms.

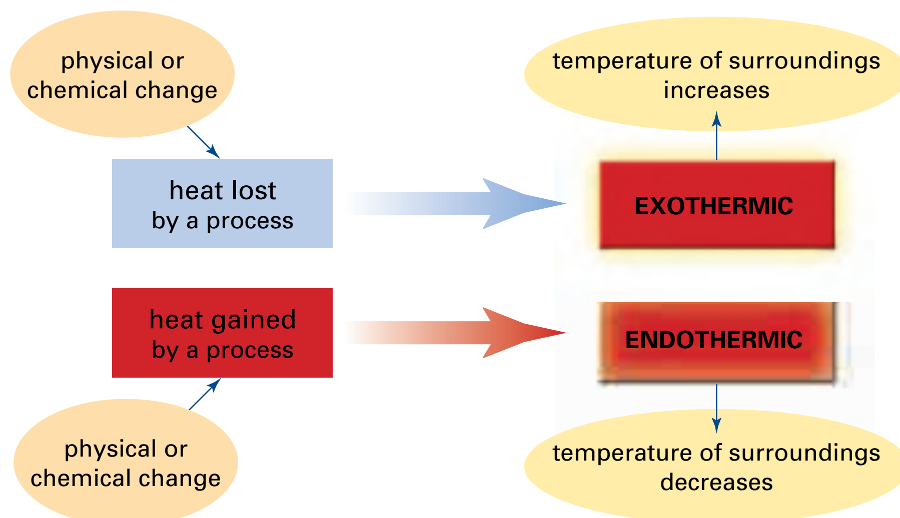


Figure 14.7

An **exothermic** reaction gives off heat. Combustion reactions are exothermic reactions because they produce heat. *Since the energy is a product of the reaction, it is shown on the product side of the equation.*

The combustion of acetylene is a good example of an exothermic reaction:

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)} + \text{energy}$$

When the energy is written as part of the chemical equation, the equation is called a **thermochemical equation**.

An **endothermic** reaction absorbs heat energy. Energy must be added to the reactants for an endothermic reaction to occur. *Since the energy is needed as a reactant, it is shown on the reactant side of the equation.*

Have you ever noticed that perfume or rubbing alcohol feels cool on your skin as it evaporates? The physical process of evaporation is endothermic. Energy is taken away from the surface of your skin, so you feel cool. The energy is added to the liquid alcohol or perfume solvent to make it a gas. The following equation shows the endothermic evaporation of isopropanol (a type of rubbing alcohol) from a liquid to a gas:

$$\text{energy} + CH_3CHOHCH_{3(\ell)} \rightarrow CH_3CHOHCH_{3(g)}$$

How can you tell that this process is not a chemical reaction?

Figure 14.6 **The combustion of acetylene (ethyne) in an oxyacetylene torch produces the highest temperature (about 3300°C) of any known mixture of combustible gases. Metal workers can use the heat from this combustion to cut through most metal alloys.**

Bond Breaking and Bond Making

The formation of acetylene (ethyne) gas from its elements is an endothermic reaction. The combustion of acetylene, however, is exothermic. In fact, it releases enough heat energy to cut steel! How can you explain the formation and combustion of acetylene in terms of bonds being broken and made? The answer to this question is fundamental to your understanding of the energy changes that occur in chemical reactions.

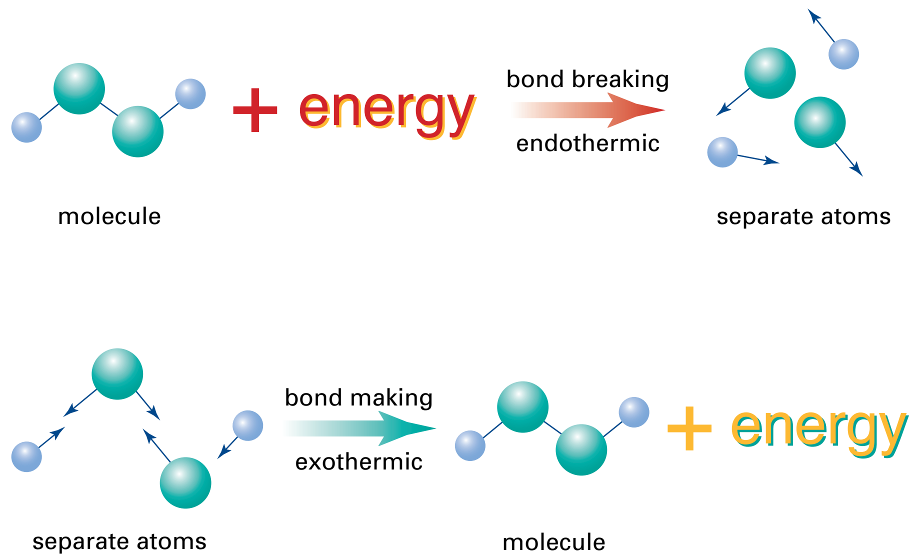


Figure 14.8 **This illustration shows bonds being broken and made during a chemical reaction. If the bonds are strong, there is a large change in energy. If the bonds are weak, there is a small change in energy.**

A chemical bond is caused by the attraction between the electrons and nuclei of two atoms. Energy is needed to break a chemical bond, just like energy is needed to break a link in a chain. On the other hand, making a chemical bond releases energy. The strength of a bond depends on how much energy is needed to break the bond. (see Figure 14.8.)

A specific amount of energy is needed to break each type of bond. When the same type of bond is formed, the same specific amount of energy is released. The energy that is absorbed or released when breaking or making a bond is called the **bond energy**. Bond energy is usually measured in kilojoules (kJ). Table 14.1 shows some average bond energies.

From the table, you can see that 347 kJ of energy is needed to break one mole of C—C bonds in a sample of propane or any other carbon. Similarly, 347 kJ of energy is released if one mole of C—C bonds forms in a sample of butane.

Every chemical reaction involves both bond breaking (reactant bonds are broken) *and* bond making (product bonds are formed). Since there are different types of bonds inside the reactant and product molecules, the bond breaking and bond making energies are different. This results in a net amount of energy for each reaction.

The next sample problem compares bond breaking and bond making in a combustion reaction.

mind STRETCH

(a) Use a molecular model kit to model the formation of butane from its elements: $4C_{(s)} + 5H_{2(g)} \rightarrow C_4H_{10(g)}$ Compare the bonds you break with the bonds you form.

(b) Using Table 14.1, estimate the energy needed to break the bonds of 5 mol of hydrogen gas. Compare this energy with the energy produced by making the bonds of 1 mol of butane gas. Predict whether the formation of butane is exothermic or endothermic.

(c) Look up "heat of formation" in a reference book such as *The CRC Handbook of Chemistry and Physics* to find the actual energy of this reaction. Is it exothermic (negative) or endothermic (positive)?

Note: Since bond energies are only a way of *estimating* the energy produced by the formation of a compound, your answer will not agree with the recognized value. Look at the Thinking Critically question in the Section Review to find out how to estimate the energy of a formation reaction more accurately.

Table 14.1 Average Bond Energies

Type of bond	Average energy (kJ/mol)
H—H	432
C—C	347
C═C	614
C≡C	839
C—H	413
C—O	358
C═O	745
O—H	467
O═O (in O_2)	498

Sample Problem

The Combustion of Acetylene

Problem

Consider bond breaking and bond making to explain why the combustion of acetylene is exothermic. Then write a thermochemical equation for the reaction, using the following balanced equation:

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$$

What Is Required?

You need to describe what happens when the reactant bonds are broken and what happens when the product bonds are formed. You need to compare the energy that is absorbed when the reactant bonds are broken with the energy that is released when the product bonds are formed. Then you need to write a thermochemical equation, using the word "energy."

What Is Given?

You know that energy is absorbed when bonds are broken and energy is released when bonds are formed. You also know that the equation is exothermic. (Overall, energy is released in this reaction.)

Plan Your Strategy

Step 1 Describe what happens when the reactant bonds are broken.

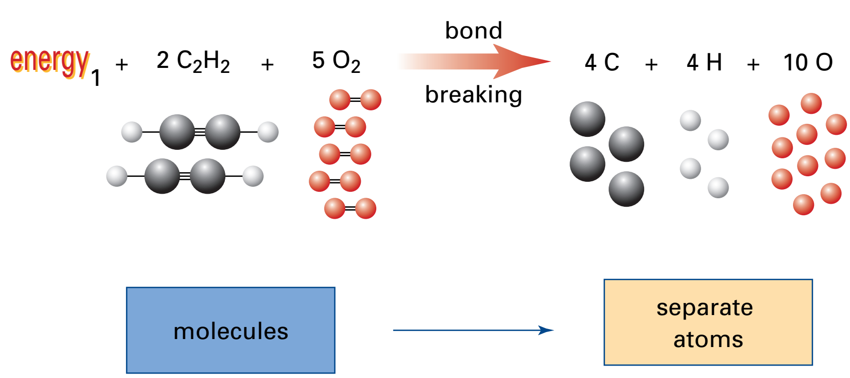
Step 2 Describe what happens when the product bonds are formed.

Step 3 Compare the energy that is absorbed when the reactant bonds are broken with the energy that is released when the product bonds are formed.

Step 4 Write the thermochemical equation, using the word "energy."

Act on Your Strategy

Step 1 The reactant bonds are broken. This process absorbs enough energy to split the reactants into separate atoms.



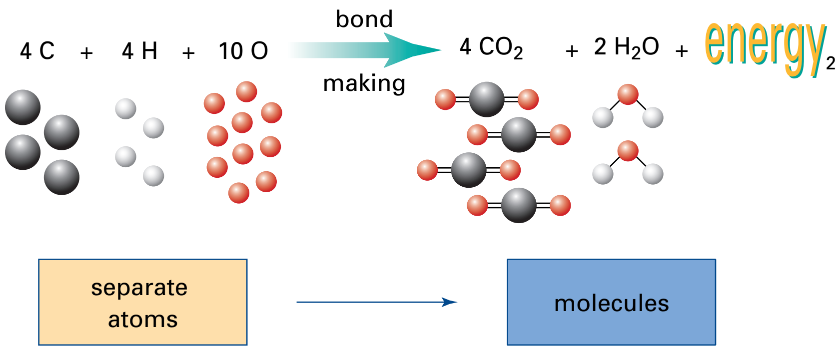
Continued ...

PROBLEM TIP

- If the energy that is needed to break the reactant bonds is greater than the energy that is released when the product bonds form, the reaction is endothermic.
- If the energy that is needed to break the reactant bonds is less than the energy that is released when the product bonds form, the reaction is exothermic.
- If the reaction is endothermic, the energy is on the left side of the equation.
- If the reaction is exothermic, the energy is on the right side of the equation.

Continued ...
FROM PAGE 590

Step 2 The product bonds are made. This process releases energy as the product molecules are formed.



Step 3 The combustion of acetylene is exothermic. Therefore, the energy that is used when the reactant bonds are broken must be less than the energy that is released when the product bonds are formed.

$$\text{energy}_1 < \text{energy}_2$$

Step 4 The thermochemical equation is

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)} + \text{energy}$$

Check Your Solution

The equation is exothermic, so the energy is on the product side of the equation.

Practice Problems

6. The formation of propane from its elements is an exothermic reaction. The combustion of propane is also exothermic.

(a) Write the balanced thermochemical equation for the formation of propane.

(b) Write the balanced thermochemical equation for the combustion of propane. (The balanced equation is on page 580.)

(c) Consider the combustion of propane. Compare the energies of bond breaking and bond making to explain why the reaction is exothermic.

7. (a) Explain why the formation of ethene, $C_2H_{4(g)}$, from its elements is endothermic, while its combustion is exothermic.

(b) Write the balanced thermochemical equations for the formation and the combustion of ethene.

PROBLEM TIP

Remember that different types of bonds have different bond energies. The bonds that are broken in the reactants are different from the bonds that are formed in the products. The net energy for the entire reaction is the difference between energy$_1$ and energy$_2$.

So far, you have been using the word "energy" in thermochemical equations to represent the net energy. It is preferable to have a numerical value for the amount of energy when talking about endothermic and exothermic processes. In the next section, you will see how the net energy in a process can be measured. You will even measure some energy values yourself!

Section Review

1 K/U Choose the correct term in each box to describe the given thermochemical equation. Name the organic compounds.

(a)

$$\text{energy} + 2C_{(s)} + H_{2(g)} \xrightarrow[\boxed{\text{net energy is absorbed/released}}]{\boxed{\text{endothermic/exothermic}}} C_2H_{2(g)}$$

(b)

$$CH_{4(g)} + 2O_{2(g)} \xrightarrow[\boxed{\text{net energy is absorbed/released}}]{\boxed{\text{endothermic/exothermic}}} CO_{2(g)} + 2H_2O_{(g)} + \text{energy}$$

2 K/U If an overall reaction is endothermic, which process involves more energy: breaking the reactant bonds or making the product bonds?

3 **(a)** C When energy is absorbed in a reaction to break the bonds in the reactants, where does the energy go?

(b) C When energy is released to form the bonds in the products, where does the energy go?

4 MC Flour, sugar, eggs, and milk are combined and baked to produce cookies.

(a) Is this reaction exothermic or endothermic? Explain.

(b) Write a word equation that includes the word "energy" to describe this reaction.

5 K/U In the MindStretch on page 589, you used bond energies to estimate the energy of a reaction. Your estimated value was very different from the actual value, however. This is because solid carbon must become gaseous *before* reacting to form a hydrocarbon. The change of state takes additional energy.

$$C_{(s)} + 715 \text{ kJ} \rightarrow C_{(g)}$$

Use the information above, along with the bond energies in Table 14.1, to estimate the energy needed to form acetylene, $C_2H_{2(g)}$, from its elements.

$$2C_{(s)} + H_{2(g)} \rightarrow C_2H_{2(g)}$$

Remember to consider the breaking of the H—H bonds, and the forming of the C—H and C≡C bonds. Compare your answer with the recognized value, 226.7 kJ/mol.

Measuring Energy Changes

14.3

You have probably had plenty of experience with heating and cooling materials around you. You know how to cool yourself by taking a cold shower. You know how to heat hot chocolate on a stove. However, have you ever stopped to think about what energy changes are occuring? Where does the energy come from, and where does it go?

Figure 14.9 Suppose that you are very thirsty, but the tap water is not cold enough. How can you get the water as cold as you want it to be? Your experience tells you to add ice to the water. How does the ice cool the water?

Any time that something is cooled or heated, a change in thermal energy occurs. **Thermal energy** is the kinetic energy of particles of matter. In a glass of tap water with ice (as shown in Figure 14.9), the water started out warmer than the ice. Energy was transferred from the water to the ice. As the ice absorbed the energy, the ice melted. As the water lost energy, the water cooled.

These examples involve heat. **Heat** is the transfer of thermal energy between objects with different temperatures. In this section, you will study how to measure the quantity of thermal energy that is transferred during a process involving an energy change.

Important Factors in Energy Measurement

How can you measure the change in energy when you add ice to water? **Temperature** is a measure of the average kinetic energy of a system. You have already had experience using a thermometer to measure temperature. By placing a thermometer in the water, you can monitor the drop in temperature when the ice is added. Temperature is an important factor in heat energy changes. What other factors are important? Work through the following ThoughtLab to find out.

Section Preview/ Specific Expectations

In this section, you will

- **explain** how mass, specific heat capacity, and change in temperature of an object determine how much heat the object gains or loses
- **solve** problems using the equation $Q = mc\Delta T$
- **communicate** your understanding of the following terms: *thermal energy*, *heat*, *temperature*, ΔT, *specific heat capacity*

CHECKPOINT

What other forms of energy can you think of? Make a list, starting with thermal energy.

mind STRETCH

How does drinking hot chocolate warm you? Describe the energy transfer that takes place.

ThoughtLab Factors in Heat Transfer

Two students performed an experiment to determine what factors need to be considered when determining the quantity of thermal energy lost or gained by a substance undergoing an energy change. They set up their experiment as follows.

Part A

The students placed two different masses of water, at the same initial temperature, in separate beakers. They placed an equal mass of ice (from the same freezer) in each beaker. Then they monitored the temperature of each beaker. Their results are listed in the table below.

Different Masses of Water

Beaker	1	2
Mass of water (g)	60.0	120.0
Initial temperature of water (°C)	26.5	26.5
Mass of ice added (g)	10.0	10.0
Final temperature of mixture (°C)	9.7	17.4
Temperature change (°C)	16.8	9.1

Part B

The students placed equal masses of canola oil and water, at the same initial temperature, in separate beakers. They placed equal masses of ice (from the same freezer) in the two beakers. Then they monitored the temperature of each beaker. Their results are listed in the following table.

Different Liquids

Beaker	1 (canola oil)	2 (water)
Mass of liquid (g)	60.0	60.0
Initial temperature of liquid (°C)	35.0	35.0
Mass of ice added (g)	10.0	10.0
Final temperature of mixture (°C)	5.2	16.9
Temperature change (°C)	29.8	18.1

Procedure

1. For each part of the experiment, identify
   (a) the variable that was changed by the students (the manipulated variable)
   (b) the variable that changed as a result of changing the manipulated variable (the responding variable)
   (c) the variables that were kept constant to ensure a fair test (the controlled variables)
2. Interpret the students' results by answering the following questions.
   (a) If ice is added to two different masses of water, how does the temperature change?
   (b) If ice is added to two different liquids, how does the temperature change?

Analysis

Think about your interpretation of the students' experiment and the discussion prior to this ThoughtLab. What are three important factors to consider when measuring the thermal energy change of a substance?

The ThoughtLab gave you some insight into the factors that are important when measuring energy changes. How can you use these factors to calculate the quantity of heat that is transferred? First you must examine each factor and determine its relationship to heat transfer. You will begin with the most obvious factor: temperature. Then you will look at how the mass of a substance affects the quantity of thermal energy it can store. Finally, you will look at the type of substance and how it affects heat transfer.

Temperature

In the summer, your body temperature is fairly close to the temperature of your surroundings. You do not need to wear extra clothing to keep warm. When winter hits with fierce winds and cold temperatures, however, dressing warmly becomes a necessity. The temperature of your surroundings is now much colder than your body temperature. Heat is transferred from your body to your surroundings, making you feel cold. The extra clothing helps to minimize heat loss from your body. (See Figure 14.10.)

Figure 14.10 How do you control the temperature of your body?

Temperature is directly related to heat transfer. A large change in temperature indicates a large energy change. A small change in temperature indicates a small energy change. Therefore, temperature is an important factor when calculating heat transfer. The temperature variable that is used is the *change in temperature*. This is symbolized by ΔT.

Mass of Substance

Did you know that 70% of Earth's surface is covered with water? This enormous mass of water absorbs and releases tremendous amounts of heat energy. Water makes our climate more moderate by absorbing heat in hot weather and releasing heat in cold weather. The greater the mass of the water, the greater the amount of heat it can absorb and release. (Areas without much water, such as deserts, experience huge variations in temperature.) Therefore, mass is directly related to heat transfer. Mass is a variable in the calculation of heat energy. It is symbolized by a lower-case m.

Type of Substance

In the ThoughtLab on page 594, you probably noticed that the quantity of heat being transferred depends on the type of substance. When you added equal masses of ice to the same mass of oil and water, the temperature change of the oil was almost double the temperature change of the water. "Type of substance" cannot be used as a variable, however, when calculating energy changes. Instead, we use a variable that reflects the individual nature of different substances: specific heat capacity. The **specific heat capacity** of a substance is the quantity of energy, in joules (J), that is required to change one gram (g) of the substance by one degree Celsius (°C). The specific heat capacity of a substance reflects how well the substance can store energy. A substance with a large specific heat capacity can absorb and release more energy than a substance with a smaller specific heat capacity. The symbol that is used for specific heat capacity is a lower-case c. The units are J/g·°C.

The specific heat capacity of water is relatively large: 4.184 J/g·°C. This value helps to explain how water can absorb and release enough energy to moderate Earth's temperature. Examine the values in Table 14.2. Notice that the specific heat capacities of most substances are much lower than the specific heat capacity of water.

Table 14.2 Specific Heat Capacities of Various Substances

Substance	Specific heat capacity (J/g·°C at 25°C)
Elements	
aluminum	0.900
carbon (graphite)	0.711
copper	0.385
gold	0.129
hydrogen	14.267
iron	0.444
Compounds	
ammonia (liquid)	4.70
ethanol	2.46
water (solid)	2.01
water (liquid)	4.184
water (gas)	2.01
Other materials	
air	1.02
concrete	0.88
glass	0.84
granite	0.79
wood	1.76

You have just considered three variables: change in temperature (ΔT), mass (m), and type of substance, which is characterized by specific heat capacity (c). How can you combine these variables into a formula to calculate heat transfer (Figure 14.11)?

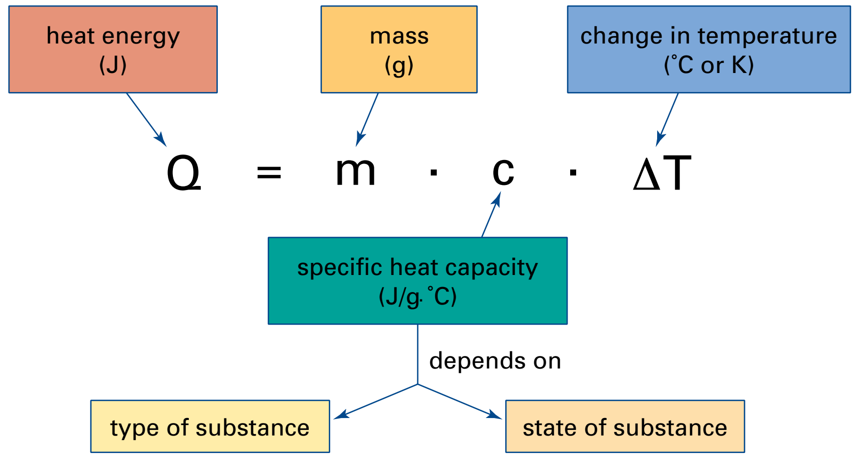


Figure 14.11 **Use this formula to calculate heat (*Q*) transfer.**

How do you solve heat energy problems using $Q = mc\Delta T$? Go back to the ThoughtLab. Some of the data in this ThoughtLab can be used to illustrate the calculation of heat transfer, as shown below.

Sample Problem

Heat Transferred From Water to Ice

Problem

In the ThoughtLab on page 594, 10.0 g of ice was added to 60.0 g of water. The initial temperature of the water was 26.5˚C. The final temperature of the mixture was 9.7˚C. How much heat was lost by the water?

What Is Required?

You need to calculate the quantity of heat (Q) that was lost by the water.

What Is Given?

You know the mass of the water. You also know the initial and final temperatures of the water.

Mass of water (m) = 60.0 g

Initial temperature (T_i) = 26.5˚C

Final temperature (T_f) = 9.7˚C

Plan Your Strategy

You have enough information to solve this problem using $Q = mc\Delta T$. Use the initial and final temperatures to calculate ΔT. You need the specific heat capacity (c) of liquid water. This is given in Table 14.2 (4.184 J/g·˚C). Because you are only concerned with the water, you will not use the mass of the ice.

Continued ...

mind STRETCH

Why does water have such a high specific heat capacity? Do some research to find out. **Hint**: Water's specific heat capacity has something to do with bonding.

Continued ...
FROM PAGE 596

Act on Your Strategy

Substitute the values into the following heat formula, and solve. Remember that $\Delta T = T_f - T_i$

$$\begin{aligned} Q &= mc\Delta T \\ &= (60.0\ \text{g})(4.184 \tfrac{\text{J}}{\text{g}\cdot{}^\circ\text{C}})(9.7^\circ\text{C} - 26.5^\circ\text{C}) \\ &= -4217.472\ (\cancel{\text{g}})(\tfrac{\text{J}}{\cancel{\text{g}}\cdot\cancel{{}^\circ\text{C}}})(\cancel{{}^\circ\text{C}}) \\ &= -4217.472\ \text{J} \\ &= -4.22 \times 10^3\ \text{J (or } -4.22\ \text{kJ)} \end{aligned}$$

The water lost 4.22×10^3 J of heat.

Check Your Solution

The water lost heat, so the heat value should be negative.
Heat is measured in joules or kilojoules. Make sure that the units cancel out to give the appropriate unit for your answer.

Practice Problems

8. 100 g of ethanol at 25°C is heated until it reaches 50°C. How much heat does the ethanol gain? **Hint**: Find the specific heat capacity of ethanol in Table 14.2.

9. In Part A of the ThoughtLab on page 594, the students added ice to 120.0 g of water in beaker 2. Calculate the heat lost by the water. Use the information given for beaker 2, as well as specific heat capacities in Table 14.2.

10. A beaker contains 50 g of liquid at room temperature. The beaker is heated until the liquid gains 10°C. A second beaker contains 100 g of the same liquid at room temperature. This beaker is also heated until the liquid gains 10°C. In which beaker does the liquid gain the most thermal energy? Explain.

11. As the diagram on the next page illustrates, the sign of the heat value tells you whether a substance has lost or gained heat energy. Consider the following descriptions. Write each heat value, and give it the appropriate sign to indicate whether heat was lost or gained.

(a) In Part A of the ThoughtLab on page 594, the ice gained the heat that was lost by the water. When ice was added to 60.0 g of water, it gained 4.22 kJ of energy. When ice was added to 120.0 g of water, it gained 4.6 kJ of energy.

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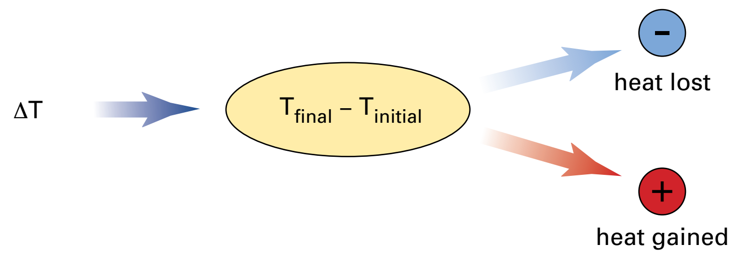

(b) When 2.0 L of water was heated over a campfire, the water gained 487 kJ of energy.

(c) A student baked a cherry pie and put it outside on a cold winter day. There was a change of 290 kJ of heat energy in the pie.

Math LINK

Heat values are often very large. Therefore it is convenient to use kilojoules (kJ) to calculate heat. How does this affect the units of the other variables in the heat equation? Does the specific heat capacity have to change? The following diagram shows how units must be modified in order to end up with kilojoules.

$$Q = m \cdot c \cdot \Delta T$$

Units: $\rightarrow$ $\text{kJ} = \text{kg} \cdot \dfrac{4.184\ \text{kJ}}{\text{kg} \cdot {}^\circ\text{C}}\ {}^\circ\text{C}$

mass-must be kg

Specific heat capacity
- must have kJ (top) and kg (bottom)
- Since "k" is on top andbottom, the number stays the same

In the Sample Problem, heat was *lost* by the water. Therefore the value of Q was *negative*. If the value of Q is *positive*, this indicates that heat is *gained* by a substance.

The heat equation $Q = mc\Delta T$ can be rearranged to solve for any of the variables. For example, in Part B of the ThoughtLab on page 594, ice was added to both canola oil and water. How can you use the information given in Part B to calculate the specific heat capacity of the canola oil?

Figure 14.12 **Canola oil is a vegetable oil that is used in salads and cooking.**

Sample Problem

Calculating Specific Heat Capacity

Problem

Calculate the specific heat capacity of canola oil, using the information given in Part B of the ThoughtLab on page 594. Note that the ice gained 4.0×10^3 J of energy when it came in contact with the canola oil.

What Is Required?

You need to calculate the specific heat capacity (c) of the canola oil.

What Is Given?

From the ThoughtLab, you know the mass (m) and the initial and final temperatures of the canola oil.

Mass of oil (m) = 60.0 g

Initial temperature (T_i) = 35.0°C

Final temperature (T_f) = 5.2°C

Continued ...

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FROM PAGE 598

You also know the quantity of heat gained by the ice. This must be the same as the heat lost by the oil.
Heat gained by the ice = Heat lost by the canola oil = 4.0×10^3 J

Plan Your Strategy

Rearrange the equation $Q = mc\Delta T$ to solve for c. Then substitute the values for Q, m, and ΔT ($T_f - T_i$) into the equation.

Act on Your Strategy

$$\begin{aligned} c &= \frac{Q}{m\Delta T} \\ &= \frac{-4.0 \times 10^3 \text{ J}}{(60.0 \text{ g})(5.2^\circ\text{C} - 35.0^\circ\text{C})} \\ &= 2.2437 \frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}} \\ &= 2.24 \frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}} \end{aligned}$$

Check Your Solution

The specific heat capacity should be positive, and it is. It should have the units $\frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}}$.

Practice Problems

12. Solve the equation $Q = mc\Delta T$ for the following quantities.

(a) m

(b) c

(c) ΔT

13. You know that $\Delta T = T_f - T_i$. Combine this equation with the heat equation, $Q = mc\Delta T$, to solve for the following quantities.

(a) T_i (in terms of Q, m, c, and T_f)

(b) T_f (in terms of Q, m, c, and T_i)

14. How much heat is required to raise the temperature of 789 g of liquid ammonia, from 25.0°C to 82.7°C?

15. A solid substance has a mass of 250.00 g. It is cooled by 25.00°C and loses 4937.50 J of heat. What is its specific heat capacity? Look at Table 14.2 to identify the substance.

16. A piece of metal with a mass of 14.9 g is heated to 98.0°C. When the metal is placed in 75.0 g of water at 20.0°C, the temperature of the water rises by 28.5°C. What is the specific heat capacity of the metal?

17. A piece of gold (c = 0.129 J/g°C) with mass of 45.5 g and a temperature of 80.5°C is dropped into 192 g of water at 15.0°C. What is the final temperature of the system? (**Hint:** Use the equation $Q_w = -Q_g$.)

In this section, you learned that temperature, mass, and type of substance are all important factors to consider when measuring heat change. You saw how these factors can be combined to produce the heat equation:

$$Q = mc\Delta T$$

In the next section, you will learn how to use a calorimeter to measure heat change. You will perform a heat transfer investigation of your own. Using the knowledge you have gained in this section, you will then calculate the specific heat capacity. As well, you will see why the Procedure that was used in the ThoughtLab could be improved.

Section Review

1. K/U Define the term "heat."

2. I What are three important factors to consider when measuring heat energy?

3. I In Part B of the ThoughtLab, 60.0 g of water was in beaker 2. The initial temperature of the water was 35.0˚C, and the final temperature was 16.9˚C.
   - **(a)** Calculate the heat that was lost by the water in beaker 2.
   - **(b)** Where did the heat go?

4. I When iron nails are hammered into wood, friction causes the nails to heat up.
   - **(a)** Calculate the heat that is gained by a 5.2 g iron nail as it changes from 22.0˚C to 38.5˚C. (See Table 14.2.)
   - **(b)** Calculate the heat that is gained by a 10.4 g iron nail as it changes from 22.0˚C to 38.5˚C.
   - **(c)** Calculate the heat that is gained by the 5.2 g nail if its temperature changes from 22.0˚C to 55.0˚C.

5. **(a)** I A 23.9 g silver spoon is put in a cup of hot chocolate. It takes 0.343 kJ of energy to change the temperature of the spoon from 24.5˚C to 85.0˚C. What is the specific heat capacity of solid silver?
   - **(b)** I The same amount of heat energy, 0.343 kJ, is gained by 23.9 g of liquid water. What is the temperature change of the water?

6. C The specific heat capacity of aluminum is 0.902 J/g˚C. The specific heat capacity of copper is 0.389 J/g˚C. The same amount of heat is applied to equal masses of these two metals. Which metal increases more in temperature? Explain.

7. K/U Explain why there is an energy difference between the following reactions.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)} + 802\text{ kJ}$$
$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(\ell)} + 890\text{ kJ}$$

The Technology of Heat Measurement

14.4

Figure 14.13 **How can you measure the energy in a substance?**

Section Preview/ Specific Expectations

In this section, you will

- **describe** some of the physical and chemical properties of hydrocarbons
- **apply** calorimetric techniques to the calculation of energy changes
- **gather** and **interpret** experimental data, and **solve** problems involving calorimetry and the combustion of hydrocarbons
- **communicate** your understanding of the following terms: *isolated system*, *calorimeter*, *calorimetry*, *bomb calorimeter*, *thermal equilibrium*, *heat of combustion*, *heat of solution*, *heat capacity*, *Calorie*, *potential energy*

In this unit, you have learned about the importance of hydrocarbons as fuels. Hydrocarbons are useful because of the energy that is released when they burn. It is often necessary, however, to know the amount of energy that is released. For example, engineers need to know how much energy is released from different fuels when they design an engine and choose an appropriate fuel. Firefighters need to know how much heat can be given off by different materials so they can decide on the best way to fight a specific fire. (See Figure 14.13.)

What about food—the fuel for your body? In order to choose an appropriate and balanced diet, you need to know how much energy each type of food releases when it is digested. Food energy is measured in Calories. (You will learn more about Calories later in this section.)

How do you measure the quantity of energy that is produced? In this section, you will focus on measuring heat changes. You will learn about some technology and techniques to measure heat. You will then apply what you have learned by performing your own heat experiments.

Calorimetry

In the ThoughtLab in section 14.3, two students used beakers with no lids when they measured change in temperature. The students assumed that energy was being exchanged only between the ice and the water. In fact, energy was also being exchanged with the surroundings. As a result, the data that the students obtained had a large experimental error. How could the students have prevented this error?

Much of the technology in our lives is designed to stop the flow of heat. Your home is insulated to prevent heat loss in the winter and heat gain in the summer. If you take hot soup to school for your lunch, you probably use a Thermos™ to prevent heat loss to the environment. Whenever there is a temperature difference between two objects, thermal energy flows from the hotter object to the colder object. When you measure the heat being transferred in a reaction or other process, you must minimize any heat that is exchanged with the surroundings.

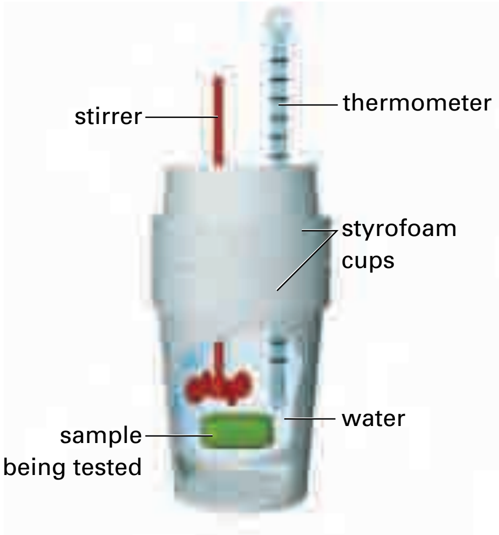

Figure 14.14 **A polystyrene (coffee cup) calorimeter usually consists of two nested polystyrene cups with a polystyrene lid, to provide insulation from the surroundings.**

To measure the heat flow in a process, you need an isolated system, like a Thermos™. An **isolated system** stops matter and energy from flowing in or out of the system. You also need a known amount of a substance, usually water. The water absorbs the heat that is produced by the process, or it releases heat if the process is endothermic. To determine the heat flow, you can measure the temperature change of the water. With its large specific heat capacity (4.184 J/ g·°C) and its broad temperature range (0°C to 100°C), liquid water can absorb and release a lot of energy.

Water, a thermometer, and an isolated system are the basic components of a calorimeter. A **calorimeter** is a device that is used to measure changes in thermal energy. (Figures 14.14 and 14.15 show two types of calorimeters.) The technological process of measuring changes in thermal energy is called **calorimetry**.

How Calorimeters Work

In a polystyrene calorimeter, a known mass of water is inside the polystyrene cup. The water surrounds, and is in direct contact with, the process that produces the energy change. The initial temperature of the water is measured. Then the process takes place and the final temperature of the water is measured. The water is stirred to maintain even energy distribution, and the system is kept at a constant pressure. This type of calorimeter can measure heat changes during processes such as dissolving, neutralization, heating, and cooling.

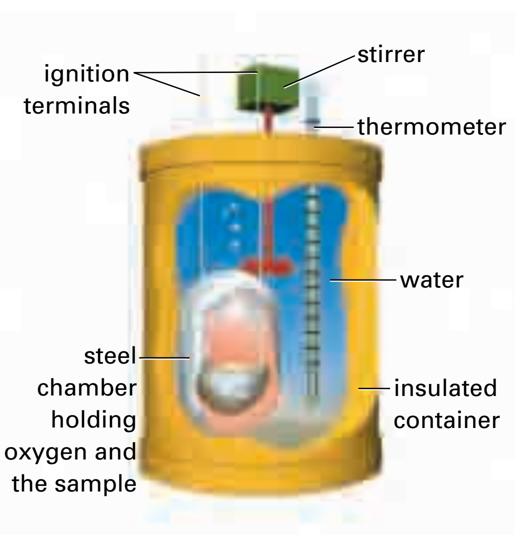

Figure 14.15 **A bomb calorimeter is more sophisticated than a polystyrene calorimeter.**

A **bomb calorimeter** is used for the measurement of heat changes during combustion reactions at a constant volume. It works on the same general principle as the polystyrene calorimeter. The reaction, however, takes place inside an inner metal chamber, called a "bomb." This "bomb" contains pure oxygen. The reactants are ignited using an electric coil. A known quantity of water surrounds the bomb and absorbs the energy that is released by the reaction. You will learn more about bomb calorimeters later in this section.

The law of conservation of energy states that energy can be changed into different forms, but it cannot be created or destroyed. This law allows you to calculate the energy change in a calorimetry experiment. However, you need to make the following assumptions:

- The system is isolated. (No heat is exchanged with the surroundings outside the calorimeter.)
- The amount of heat energy that is exchanged with the calorimeter itself is small enough to be ignored.
- If something dissolves or reacts in the calorimeter water, the solution still retains the properties of water. (For example, density and specific heat capacity remain the same.)

Tools & Techniques

The First Ice Calorimeter

A calorimeter measures the thermal energy that is absorbed or released by a material. Today we measure heat using joules (J) or calories (cal). Early scientists accepted one unit of heat as the amount of heat required to melt 1 kg of ice. Thus two units of heat could melt 2 kg of ice.

The earliest measurements of heat energy were taken around 1760, by a Scottish chemist named Joseph Black. He hollowed out a chamber in a block of ice. Then he wiped the chamber dry and placed a piece of platinum, heated to 38°C, inside. He used another slab of ice as a lid. As the platinum cooled, it gave up its heat to the ice. The ice melted, and water collected in the chamber. When the platinum reached the temperature of the ice, Black removed the water and weighed it to find out how much ice had melted. In this way, he measured the quantity of heat that was released by the platinum.

In 1780, two French scientists, Antoine Lavoisier and Pierre Laplace, developed the first apparatus formally called a calorimeter. Like Black, they used the amount of melted ice to measure the heat released by a material. Their calorimeter consisted of three concentric chambers. The object to be tested was placed in the innermost chamber. Broken chunks of ice were placed in the middle chamber. Ice was also placed in the outer chamber to prevent any heat reaching the apparatus from outside. As the object in the inner chamber released heat, the ice in the middle chamber melted. Water was drawn from the middle chamber by a tube, and then measured.

Lavoisier made many important contributions to the science of chemistry. Unfortunately his interest in political reform led to his arrest during the French Revolution. He was beheaded after a trial that lasted less than a day.

The original calorimeter used by Lavoisier and Laplace

When a process causes an energy change in a calorimeter, the change in temperature is measured by a thermometer in the water. If you know the mass of the water and its specific heat capacity, you can calculate the change in thermal energy caused by the process. See Figures 14.16 and 14.17 for examples.

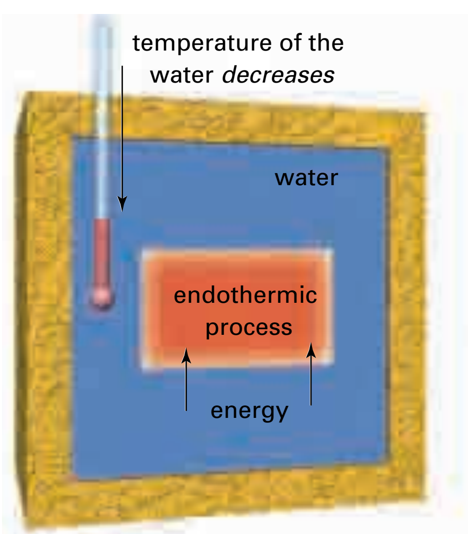

Figure 14.16 **An endothermic process, such as ice melting**

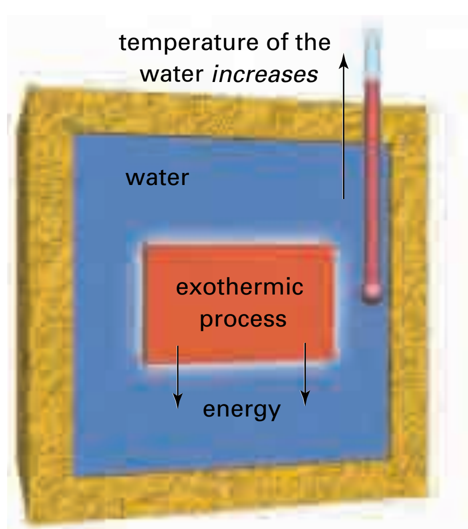

Figure 14.17 **An exothermic process, such as the combustion of propane**

The energy change in a calorimetry experiment can be summarized as follows:

Heat lost by the process = Heat gained by the water

or

Heat gained by the process = Heat lost by the water

In the next Sample Problem, you will use what you have just learned to calculate the specific heat capacity of a metal. This problem is similar to the calculation of the specific heat capacity of canola oil in section 14.3. Here, however, a calorimeter is used to reduce the heat exchange to the environment.

Sample Problem

Determining a Metal's Specific Heat Capacity

Problem

A 70.0 g sample of a metal was heated to 95.0°C in a hot water bath. Then it was quickly transferred to a polystyrene calorimeter. The calorimeter contained 100.0 g of water at an initial temperature of 19.8°C. The final temperature of the contents of the calorimeter was 22.6°C.

(a) How much heat did the metal lose? How much heat did the water gain?

(b) What is the specific heat capacity of the metal?

What Is Required?

(a) You need to calculate the heat lost by the metal (Q_m) and the heat gained by the water (Q_w).

(b) You need to calculate the specific heat capacity of the metal.

What Is Given?

You know the mass of the metal, and its initial and final temperatures.

Mass of metal (m_m) = 70.0 g

Initial temperature of metal (T_i) = 95.0°C

Final temperature of metal (T_f) = 22.6°C

You also know the mass of the water, and its initial and final temperatures.

Mass of water (m_w) = 100.0 g

Initial temperature of water (T_i) = 19.8°C

Final temperature of water (T_f) = 22.6°C

As well, you know the specific heat capacity of water: 4.184 J/g·°C.

Continued ...

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FROM PAGE 604

Plan Your Strategy

(a) You have all the information that you need to find the heat gained by the water. Use the heat equation $Q = mc\Delta T$. To find the heat lost by the metal, assume that $Q_m = -Q_w$.

(b) Calculate the specific heat capacity of the metal by rearranging the heat equation and solving for c.

It is very important that you do not mix up the given information. For example, when solving for the thermal energy change of the water, Q_w, make sure that you only use variables for the water. You must use the initial temperature of the water, 19.8°C, *not* the initial temperature of the metal, 95.0°C. Also, remember that $\Delta T = T_f - T_i$.

Act on Your Strategy

(a) Solve for Q_w.

$$\begin{aligned} Q_w &= m_w c_w \Delta T_w \\ &= (100.0\ \text{g})\left(4.184\ \frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(22.6^\circ\text{C} - 19.8^\circ\text{C}) \\ &= 1171.52\ (\cancel{\text{g}})\left(\frac{\text{J}}{\cancel{\text{g}}\cdot\cancel{{}^\circ\text{C}}}\right)(\cancel{{}^\circ\text{C}}) \\ &= 1.2 \times 10^3\ \text{J} \end{aligned}$$

The water gained 1.2×10^3 J of thermal energy.

Solve for Q_m.

$$\begin{aligned} Q_m &= -Q_w \\ &= -1.2 \times 10^3\ \text{J} \end{aligned}$$

The metal lost 1.2×10^3 J of thermal energy.

(b) Solve for c_m.

$$\begin{aligned} c_m &= \frac{Q_m}{m_m \Delta T_m} \\ &= \frac{-1.2 \times 10^3\ \text{J}}{(70.0\ \text{g})(22.6^\circ\text{C} - 95.0^\circ\text{C})} \\ &= 0.24\ \text{J/g}\cdot{}^\circ\text{C} \end{aligned}$$

The specific heat capacity of the metal is 0.24 J/g·°C.

Check Your Solution

The heat gained by the water is a positive value. The heat lost by the metal is a negative value. The heat is expressed in joules. (Kilojoules are also acceptable.)

The specific heat capacity of the metal is positive, and it has the correct units.

Notice that all the materials in the calorimeter in the Sample Problem had the same final temperature. This is called **thermal equilibrium**.

Figure 14.18

Heat of Combustion

How much energy is needed for a natural gas water heater, like the one shown in Figure 14.18, to heat the hot water in your home? It is probably much more than you think! In the next Sample Problem, you will find out. You will also calculate the **heat of combustion** of a hydrocarbon: the heat that is released when combustion occurs.

Sample Problem

Calculating Thermal Energy

Problem

Many homes in North America use natural gas for general heating and for water heating. Like calorimeters, natural gas water heaters have an insulated container that is filled with water. A gas flame at the bottom heats the water. A typical water heater might hold 151 L of water.

(a) How much thermal energy is needed to raise the temperature of 151 L of water from 20.5°C to 65.0°C? Note: Make the same three assumptions that you made for calorimeters.

(b) If it takes 506 g of methane to heat this water, what is the heat of combustion of methane per gram?

What Is Required?

(a) You need to calculate the quantity of thermal energy (Q) needed to heat 151 L of water.

(b) You need to calculate the heat released per gram of methane burned.

What Is Given?

(a) You know the initial and final temperatures of the water. You also know the volume of the water.
Initial temperature (T_i) = 20.5°C
Final temperature (T_f) = 65.0°C
Volume = 151 L
As well, you know the specific heat capacity of water (c):
4.184 J/g·°C or 4.184 kJ/kg·°C.

(b) You know the mass of the methane. Mass of methane (m) = 506 g

Plan Your Strategy

(a) This problem involves thermal energy and a change in temperature. You can use the heat equation $Q = mc\Delta T$. First calculate the mass of 151 L of water. (Remember that the density of water at room temperature is 1 g/mL, or 1 kg/L.) If you express the mass of the water in kilograms, you must also use the appropriate specific heat capacity of water: 4.184 kJ/kg·°C.

Note: To keep the calculation simple, assume that the density of the water remains the same when it is heated. (This is not strictly true.)

Continued ...

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(b) Use the concept of heat lost = heat gained. Since a loss of heat gives a negative value, use the following equation.

$$Q_m = -Q_w$$

To find the heat per gram, divide the amount of heat by the mass of methane.

Act on Your Strategy

(a) Mass of water = Volume × Density

$$\begin{aligned} &= (151 \text{ L})(1 \text{ kg/L}) \\ &= 151 \text{ kg} \end{aligned}$$

Substitute into $Q = mc\Delta T$, and solve.

$$\begin{aligned} Q &= mc\Delta T \\ &= (151 \text{ kg})(4.184 \text{ kJ/kg·°C})(65.0\text{°C} - 20.5\text{°C}) \\ &= 28\ 114\ (\cancel{\text{kg}})(\text{kJ}/\cancel{\text{kg·°C}})(\cancel{\text{°C}}) \\ &= 2.81 \times 10^4 \text{ kJ} \end{aligned}$$

Therefore, 2.81×10^4 kJ, or 28.1 MJ (megajoules), of energy is needed to heat the water. (This is a great deal of energy!)

(b) $Q_m = -Q_w$

$$= -2.81 \times 10^4 \text{ kJ}$$

Divide the amount of heat by the mass of methane to find the heat per gram.

$$\begin{aligned} Q_m \text{ (per gram)} &= \frac{Q_m}{m_m} \\ &= \frac{-2.81 \times 10^4 \text{ kJ}}{506 \text{ g}} \\ &= -55.5 \text{ kJ/g} \end{aligned}$$

This means that 55.5 kJ of thermal (heat) energy is *released* for each gram of methane that burns.

Check Your Solution

(a) The water gains heat, so the heat value is positive. Heat is expressed in kilojoules. (Joules are also acceptable.)

(b) The methane loses energy, so the heat value is negative. Since this value is the heat per gram, the unit is kJ/g.

Practice Problems

18. A reaction lowers the temperature of 500.0 g of water in a calorimeter by 1.10°C. How much heat is absorbed by the reaction?

19. Aluminum reacts with iron(III) oxide to yield aluminum oxide and iron. The temperature of 1.00 kg of water in a calorimeter increases by 3.00°C during the reaction. Calculate the heat that is released in the reaction.

Continued ...
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20. 5.0 g of an unknown solid was dissolved in 100 g water in a polystyrene calorimeter. The initial temperature of the water was 21.7°C, and the final temperature of the solution was 29.6°C

(a) Calculate the heat change caused by the solid dissolving.

(b) What is the heat of solution per gram of solid dissolved?

21. A 92.0 g sample of a substance, with a temperature of 55.0°C, is placed in a polystyrene calorimeter. The calorimeter contains 1.00 kg of water at 20.0°C. The final temperature of the system is 25.2°C.

(a) How much heat did the substance lose? How much heat did the water gain?

(b) What is the specific heat capacity of the substance?

Heat of Solution

In Practice Problem 20, you calculated the thermal energy change as a solid dissolved in water. This value is called the **heat of solution**: the energy change caused by a substance dissolving. The following ExpressLab deals with the heat of solution of a solid.

ExpressLab The Energy of Dissolving

In this lab you will measure the heat of solution of two solids.

Safety Precautions

- NaOH and KOH can burn skin. If you accidentally spill NaOH or KOH on your skin, wash immediately with copious amounts of cold water.

Materials

balance and beakers or weigh boats
polystyrene calorimeter
thermometer and stirring rod
distilled water
2 pairs of solid compounds:
- ammonium nitrate and potassium hydroxide
- potassium nitrate and sodium hydroxide

Procedure

1. Choose *one* pair of chemicals from the list.
2. For each of the two chemicals, calculate the mass required to make 100.0 mL of a 1.00 mol/L aqueous solution.
3. Measure the required mass of one of the chemicals in a beaker or a weigh boat.
4. Measure exactly 100 g of distilled water directly into your calorimeter.
5. Measure the initial temperature of the water.
6. Pour one of the chemicals into the calorimeter. Put the lid on the calorimeter.
7. Stir the solution. Record the temperature until there is a maximum temperature change.
8. Dispose of the chemical as directed by your teacher. Clean your apparatus.
9. Repeat steps 3 to 8, using the other chemical.

Analysis

1. For each chemical you used, calculate the heat change per gram and the heat change per mole of substance dissolved.
2. Which chemical dissolved endothermically? Which chemical dissolved exothermically?
3. One type of cold pack contains a compartment of powder and a compartment of water. When the barrier between the two compartments is broken, the solid dissolves in the water and causes an energy change. What chemical could be used in this type of cold pack? Why?

A Closer Look at Bomb Calorimetry

Polystyrene calorimeters are reasonably efficient for measuring heat changes during physical processes, such as dissolving and phase changes. They can also be used to measure heat changes during chemical processes, such as neutralization. A stronger and more precise type of calorimeter is needed, however, to measure the heat of combustion of foods, fuels, and other materials. As you learned earlier, bomb calorimeters are used for this purpose. (See Figure 14.19.)

Figure 14.19 **Bomb calorimeters give more accurate measurements than polystyrene calorimeters.**

A bomb calorimeter has many more parts than a polystyrene calorimeter. All of these parts can absorb or release small quantities of energy. Therefore, you cannot assume that the heat lost to the calorimeter is small enough to be negligible. To obtain precise heat measurements, you must know or find out the heat capacity of the bomb calorimeter. **Heat capacity** is the ratio of the heat gained or lost by a system to the change in temperature caused by this heat. It is usually expressed in kJ/°C. Unlike specific heat capacity, which refers to a single substance, heat capacity refers to a system. Thus, the heat capacity of a calorimeter takes into account the heat that *all* parts of the calorimeter can lose or gain. (See Figure 14.20.)

$$C_{\text{total}} = C_{\text{water}} + C_{\text{thermometer}} + C_{\text{stirrer}} + C_{\text{container}}$$

heat capacity of calorimeter

Figure 14.20 **Heat capacity is symbolized by an upper-case *C*. It is usually expressed in the unit J/°C.**

Web LINK

www.school.mcgrawhill.ca/resources

Why do some solids dissolve exothermically, while other solids dissolve endothermically? What factors may be involved? Research two factors: *lattice energies* and *solvation energies*. Go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. Report your results to the class.

CHEM FACT

Have you ever strained a muscle or sprained a joint? You may obtain temporary relief by applying the right heat of solution. A hot pack or cold pack consists of a thick outer pouch that contains water and a thin inner pouch that contains a salt. A squeeze on the outer pouch breaks the inner pouch, and the salt dissolves. Most hot packs use anhydrous $CaCl_2$. Most cold packs use NH_4NO_3. The change in temperature can be quite large. A cold pack, for instance, can bring the solution from room temperature down to 0°C. The usable time, however, is limited to around half an hour.

A bomb calorimeter is calibrated for a constant mass of water. Since the mass of the other parts remain constant, there is no need for mass units in the heat capacity value. The manufacturer usually includes the heat capacity value(s) in the instructions for the calorimeter.

Heat calculations must be done differently when the heat capacity of a calorimeter is included. The next Sample Problem illustrates this.

Sample Problem

Calculating Heat Change in a Bomb Calorimeter

Problem

A laboratory decided to test the energy content of peanut butter. A technician placed a 16.0 g sample of peanut butter in the steel bomb of a calorimeter, along with sufficient oxygen to burn the sample completely. She ignited the mixture and took heat measurements. The heat capacity of the calorimeter was calibrated at 8.28 kJ/°C. During the experiment, the temperature increased by 50.5°C.

(a) What was the thermal energy released by the sample of peanut butter?

(b) What is the heat of combustion of the peanut butter per gram of sample?

What Is Required?

(a) You need to calculate the heat (Q) lost by the peanut butter.

(b) You need to calculate the heat lost per gram of peanut butter.

What Is Given?

You know the mass of the peanut butter, the heat capacity of the calorimeter, and the change in temperature of the system.

Mass of peanut butter (m) = 16.0 g

Heat capacity of calorimeter (C) = 8.28 kJ/°C

Change in temperature (ΔT) = 50.5°C

Plan Your Strategy

(a) The heat capacity of the calorimeter takes into account the specific heat capacities and masses of all the parts of the calorimeter. Calculate the heat change of the calorimeter, Q_{cal}, using the equation

$$Q_{cal} = C\Delta T$$

Note: C is the heat capacity of the calorimeter in J/°C or kJ/°C. It replaces the m and c in other calculations involving specific heat capacity.

First calculate the heat gained by the calorimeter. When the peanut butter burns, the heat lost by the peanut butter sample equals the heat gained by the calorimeter.

$$Q_{sample} = -Q_{cal}$$

Continued ...

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FROM PAGE 610

(b) To find the heat of combustion per gram, divide the heat by the mass of the sample.

Act on Your Strategy

(a)
$$
\begin{aligned}
Q_{\text{cal}} &= C\Delta T \\
&= (8.28\text{kJ}/^\circ\text{C})(50.5^\circ\text{C}) \\
&= 418.14(\text{kJ}/^\circ\cancel{\text{C}})(^\circ\cancel{\text{C}}) \\
&= 418\ \text{kJ}
\end{aligned}
$$

The calorimeter gained 418 kJ of thermal energy.

$$
\begin{aligned}
Q_{\text{sample}} &= -Q_{\text{cal}} \\
&= -418\ \text{kJ}
\end{aligned}
$$

The sample of peanut butter released 418 kJ of thermal energy.

(b)
$$
\begin{aligned}
\text{Heat of combustion per gram} &= \frac{\text{Heat released}}{\text{Mass of sample}} \\
&= \frac{-418\ \text{kJ}}{16.0\ \text{g}} \\
&= -26.2\ \text{kJ/g}
\end{aligned}
$$

The heat of combustion per gram of peanut butter is –26.2 kJ/g.

Check Your Solution

Heat was lost by the peanut butter, so the heat value is negative.

Practice Problems

22. Use the heat equation for a calibrated calorimeter, $Q_{\text{cal}} = C\Delta T$. Recall that $\Delta T = T_{\text{f}} = T_{\text{i}}$. Solve for the following quantities.

(a) C

(b) ΔT

(c) T_{f} (in terms of C, ΔT, and T_{i})

(d) T_{i} (in terms of C, ΔT, and T_{f})

23. A lab technician places a 5.00 g food sample into a bomb calorimeter that is calibrated at 9.23 kJ/°C. The initial temperature of the calorimeter system is 21.0°C. After burning the food, the final temperature of the system is 32.0°C. What is the heat of combustion of the food in kJ/g?

24. A scientist places a small block of ice in an uncalibrated bomb calorimeter. The ice melts, gains 10.5 kJ (10.5×10^3 J) of heat and undergoes a temperature change of 25.0°C. The calorimeter undergoes a temperature change of 1.2°C.

(a) What mass of ice was added to the calorimeter? (Use the heat capacity of liquid water.)

(b) What is the calibration of the bomb calorimeter in kJ/°C?

CHECKPOINT

It takes one calorie (small c) to heat 1 g of water by 1˚C. What mass of water can one Calorie (large C) heat by 1˚C?

Food as a Fuel

Food is the fuel for your body. It provides you with the energy you need to function every day. Unlike the peanut butter in the previous problem, the food you digest is not burned. The process of digestion, however, is very similar to burning. In fact, people often talk about "burning off Calories." When food is digested, it undergoes slow combustion (without flames!) as it reacts with the oxygen you breathe. Eventually this combustion produces the materials that your body needs. It also releases carbon dioxide and water vapour as waste.

Figure 14.21 **Food contains energy, which is usually measured in Calories.**

You have probably noticed that Calories are used more often than kilojoules when discussing the energy in food (Figure 14.21). How are these terms related?

For years, chemists used the calorie as a unit of energy. One calorie is equal to 4.184 J. This is the amount of energy that is required to heat 1 g of liquid water by 1˚C. The food **Calorie** (notice the upper-case C) is equal to one thousand calories, or one kilocalorie. *Therefore one food Calorie is equal to 4.184 kJ.*

You will recall that in the last sample problem, 16.0 g of peanut butter released 418 kJ of energy. To translate a value in kilojoules into Calories, multiply it by the fraction $\frac{1 \text{ Cal}}{4.184 \text{ kJ}}$. For the Sample Problem,

$$418 \text{ kJ} \times \frac{1 \text{ Cal}}{4.184 \text{ kJ}} = 100 \text{ Cal}$$

Therefore, 16.0 g of peanut butter released 100 Cal of energy. Table 14.3 gives energy values in Calories for some foods.

Table 14.3 Energy Values for Some Common Foods

Food	Quantity	Energy (kJ)	Energy (Cal)
almonds (shelled, whole)	75 g	1880	449
apple	100 g	283	68
beef (broiled)	90 g	1330	318
chicken (breast, broiled)	84 g	502	120
tuna (canned)	90 g	740	177
carrots (raw)	50 g	80	19
bread (white, enriched)	30 g	340	81
spaghetti (cooked)	148 g	690	165
olive oil	232 g (1 cup)	8580	2051
caramels (plain)	30 g (3 caramels)	480	115

When you take in food energy, your body stores excess food energy in the form of fat. If your body needs energy later, it will use up some of this fat. This is the secret behind hibernation. Why is excess energy in your body stored as fat? Why is it not stored as protein or carbohydrates? The next ThoughtLab will examine these questions. You will do calculations to find out which substance releases the most heat when burned.

Figure 14.22 **Many mammals in the animal kingdom rely on fat that is stored in their bodies. By surviving on fat reserves that are stored during the autumn, bears can hibernate throughout the winter without eating.**

ThoughtLab Energy Content in Fat and Carbohydrates

By comparing the thermal energy that is released when fats, proteins, and carbohydrates are burned in a bomb calorimeter, you can compare the energy that is stored in these compounds. Natural fats are made up of various types of *fatty acids*. Fatty acids are long chain organic acids. The most common fatty acid in nature is oleic acid.

Glucose is a common sugar in the body. Most sugars that you ingest are broken down into glucose before they are digested further. Collagen is one of the most common proteins in your body.

Which compound releases more energy per gram: a fat (assume oleic acid, $C_{17}H_{33}COOH_{(\ell)}$), a carbohydrate such as sugar (assume glucose, $C_6H_{12}O_{6(s)}$), or a protein (assume collagen, molar mass 300 000 g/mol)?

A sample of glucose is placed in a bomb calorimeter, along with oxygen. The glucose is completely burned. The process is repeated with a sample of oleic acid and a sample of collagen. The following results are recorded:

Compound	sugar (glucose)	fat (oleic acid)	protein (collagen)
mass of compound (g)	1.35	1.23	1.31
initial temperature (°C)	25.20	25.00	25.10
final temperature (°C)	27.65	30.56	28.74
heat capacity of calorimeter (kJ/°C)	8.28	8.28	8.28

Procedure

1. For each substance, calculate the heat energy released per gram of substance burned.
2. For each substance, calculate the heat of combustion per mole of substance burned.

Analysis

1. Based on your calculations, which substance stores more energy? How do you know?

Application

2. Investigate why fat produces more energy than sugar when combusted. You may want to research and compare the bond structures of glucose and oleic acid.

Comparing Fats and Hydrocarbons

One of the substances that you considered in the last ThoughtLab was oleic acid. Oleic acid has a long hydrocarbon chain in each molecule. This hydrocarbon chain is similar to the hydrocarbon chains of fossil fuels, such as octane. (See Figure 14.23.)

oleic acid

$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH=CH-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-COOH$

hydrocarbon chain

$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$

octane

Figure 14.23 **By comparing octane and oleic acid, you can see that the heat of combustion of fossil fuels must be similar to the heat of combustion of fats.**

mind STRETCH

Which stores more energy: *saturated* fat or *unsaturated* fat? Do research to find out. Then compare the structures of these two types of fat to explain why.

What is the relationship between the chemical bonds in a compound and the amount of energy that the compound can store? You already know that the net energy of a reaction equals the difference between *the energy absorbed when the reactant bonds are broken* and *the energy released when the product bonds are formed*. The size of the difference reflects the strength of the bonds in the reactant molecules compared with the strength of the bonds in the product molecules.

Figure 14.24 shows an exothermic reaction: the combustion of methane. The stored energy of the products is less than the stored energy of the reactants. Therefore, a net amount of energy is released by the reaction. In an endothermic reaction, the stored energy of the products is greater than the stored energy of the reactants.

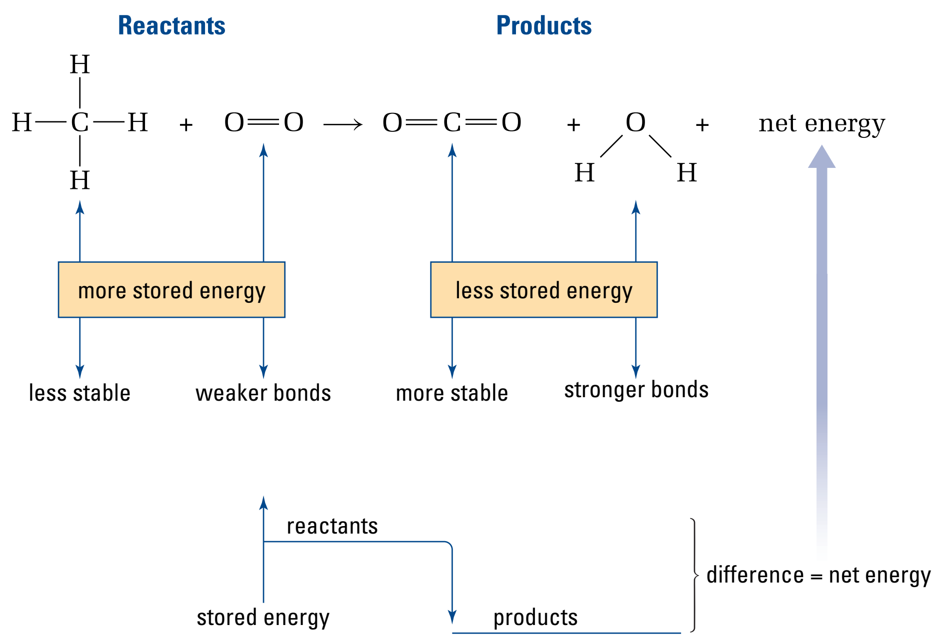


Figure 14.24 **Energy stored in bonds**

Most of the energy change in a reaction is due to a change in the **potential energy** (stored energy) of the bonds. All hydrocarbon fuels are made of the same elements: carbon and hydrogen. They react to give the same combustion products. As a result, the difference in the quantity of heat energy that is released by any two hydrocarbons is directly related to the stored energy in the bonds of these compounds. The next ThoughtLab compares the heat energy that is released by the combustion of two fuels: propane and butane.

ThoughtLab Heat Combustion of Propane and Butane

Propane and butane are very common hydrocarbon fuels. As successive members of the alkane family, they are closely related, but have a different number of bonds. Does this difference in the number of bonds affect the quantity of heat released during combustion?

Samples of propane and butane were completely burned in a bomb calorimeter. The calorimeter was calibrated with a heat capacity of 8.28 kJ/°C. The observations are given in the table below.

Substance	Mass of sample (g)	Initial temperature (°C)	Final temperature (°C)
propane	1.50	25.00	34.03
butane	1.50	25.00	33.87

Procedure

1. Predict which compound will release the greater quantity of energy per mole of gas.
2. Calculate the heat of combustion per mole for each substance.
3. (a) Which substance has the higher heat of combustion per mole?
   (b) Draw complete structural diagrams of propane and butane. Use these diagrams to explain your answer to part (a).

Analysis

1. Was your prediction correct?
2. Find and compare the boiling points of propane and butane. Which fuel would be better for winter camping? Why?

The Combustion of Candles

So far in this chapter, you have focussed on gaseous hydrocarbons, such as methane and acetylene. You have examined their heats of combustion and various processes in which they are used. In your everyday life, you may have encountered another type of hydrocarbon: paraffins. Paraffins are long chain hydrocarbons. They are semisolid or solid at room temperature. One type of paraffin has been a household item for centuries—paraffin wax, $C_{25}H_{52(s)}$, better known as candle wax. (See Figure 14.25.)

Like other hydrocarbons, the paraffin wax in candles undergoes combustion when burned. It releases thermal energy in the process. In the following investigation, you will measure this thermal energy.

Figure 14.25 Paraffin wax candles have been an important light source for hundreds of years.

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Investigation 14-B

The Heat of Combustion of a Candle

You have probably gazed into the flame of a candle without thinking about chemistry! Now, however, you will use the combustion of candle wax to gain insight into the measurement of heat changes. You will also evaluate the design of this investigation and make suggestions for improvement.

Question

What is the heat of combustion of candle wax?

Prediction

Will the heat of combustion of candle wax be greater or less than the heat of combustion of other fuels, such as propane and butane? Record your prediction, and give reasons.

Safety Precautions

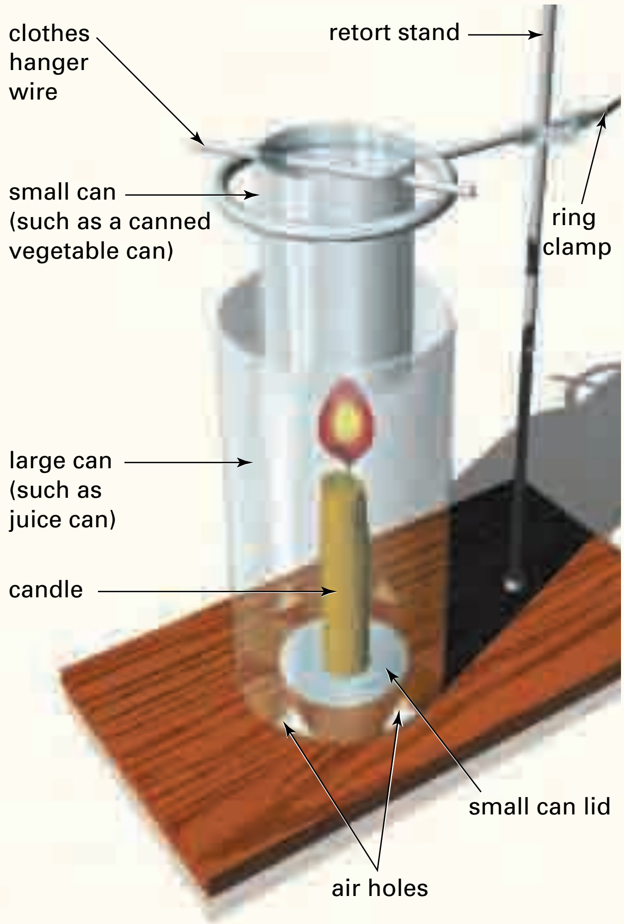
- Tie back long hair and confine any loose clothing. Before you light the candle, check that there are no flammable solvents near by.

Materials

balance
calorimeter apparatus (see the diagram to the right)
thermometer
stirring rod
matches
water
candle

Procedure

1. Burn the candle to melt some wax. Use the wax to attach the candle to the smaller can lid. Blow out the candle.
2. Set up the apparatus as shown in the diagram, but do not include the large can yet. Adjust the ring stand so that the small can is about 5 cm above the wick of the candle. The tip of the flame should just touch the bottom of the small can.
3. Measure the mass of the candle and the lid.
4. Measure the mass of the small can and the hanger.
5. Place the candle inside the large can on the retort stand.
6. Fill the small can about two-thirds full of cold water (10°C to 15°C). You will measure the mass of the water later.



7. Stir the water in the can. Measure the temperature of the water.

8. Light the candle. Quickly place the small can in position over the candle. **CAUTION** Be careful of the open flame.

9. Continue stirring. Monitor the temperature of the water until it has reached 10°C to 15°C above room temperature.

10. Blow out the candle. Continue to stir. Monitor the temperature until you observe no further change.

11. Record the final temperature of the water. Examine the bottom of the small can, and record your observations.

12. Measure the mass of the small can and the water.

13. Measure the mass of the candle, lid, and any drops of candle wax.

Analysis

1. (a) Calculate the mass of the water.
   (b) Calculate the mass of candle wax that burned.

2. Calculate the thermal energy that was absorbed by the water.

3. Calculate the heat of combustion of the candle wax per gram.

4. (a) Assume that the candle wax is pure paraffin wax, $C_{25}H_{52(s)}$. Calculate the heat of combustion per mole of paraffin wax.
   (b) Write a balanced thermochemical equation for the complete combustion of paraffin wax. **Hint**: You calculated the heat of combustion per mole of paraffin wax in part (a). The chemical equation is balanced for one mole of $C_{25}H_{52(s)}$. Therefore, you can use an actual value for the energy in the equation.

Conclusions

5. (a) List some possible sources of error that may have affected the results you obtained.
   (b) Evaluate the design and the procedure of this investigation. Consider the apparatus, the combustion, and anything else you can think of. Make suggestions for possible improvements.

6. What if soot (unburned carbon) accumulated on the bottom of the small can? Would this produce a greater or a lower heat value than the value you expected? Explain.

mind STRETCH

Go to a library, or access the Internet. Complete a brief research report on how "no-drip" candles are made. How does the composition of these candles affect the combustion of the paraffin?

In this section, you considered the importance of fossil fuels in our society, based on the energy that we can obtain from fuel combustion. You have seen how thermal energy can be measured. You now know how much energy can be extracted from molecules that stored solar energy millions of years ago. Can our society continue to depend on this energy source without harming the environment? What risks result from our use of fossil fuels? You will explore these questions in the next section.

Section Review

1. K/U What is the difference between *specific heat capacity* and *heat capacity*?

2. I List two characteristics of a calorimeter that are necessary for successful heat measurement.

3. I A calorimeter is calibrated at 7.61 kJ/°C. When a sample of coal is burned in the calorimeter, the temperature increases by 5.23°C. How much heat was lost by the coal?

4. I A reaction in a calorimeter causes 150 g of water to decrease in temperature by 5.0°C. What is the thermal energy change of the water?

5. I A company claims that its new Calorie-reduced dessert has less than 10 Cal per serving. To test this claim, a technician at the department of Consumer and Corporate Affairs completely burns a serving of the dessert in a bomb calorimeter. The temperature change is 4.86°C. The heat capacity of the calorimeter is 8.28 kJ/°C. Is the company's claim correct?

6. **(a)** C In Chapter 13, you compared the boiling points, a physical property, of some hydrocarbons. In this section, you compared the heats of combustion, a chemical property, of propane, butane, and paraffin. Use a reference book, such as *The CRC Handbook of Chemistry and Physics*. Look up the following hydrocarbons: methane, ethane, propane, butane, pentane, and hexane. Compare their densities, melting points, boiling points, and heats of combustion. Record your findings in charts and/or graphs. What patterns do you observe? Which properties are physical, and which are chemical? Explain.

   (b) C Use your findings to estimate the heat of combustion of heptane.

7. MC At a comfortable indoor temperature of 22°C, your body is at a higher temperature than its surroundings. Therefore, it is constantly radiating heat to the environment. It needs, however, to maintain an internal temperature of 37°C. How does it replace the heat that it is constantly losing to the environment?

8. K/U A heat pack can be used to supply heat to injuries. One type of heat pack is re-usable. 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 that is released by this kind of pack?

The Impact of Petroleum Products

14.5

Products made from petrochemicals (which are obtained from petroleum and natural gas) have profoundly influenced your life. Because of petrochemicals, you can use numerous products that did not exist 50 years ago (See Figures 14.26, 14.27, and 14.28). For example, one of the most common materials around you—plastic—is made from petrochemicals. The large numbers of petrochemicals that are used in manufacturing, however, are insignificant when compared with the vast quantities of hydrocarbons that are consumed each year as fossil fuels. As you learned in Chapter 13, almost 95% of all petroleum that is extracted from the ground is used as fuels and lubricants for vehicles.

Is there a cost for all that we obtain from hydrocarbons? Do these amazing products carry any drawbacks? In this section, you will examine some pros and cons of our society's reliance on hydrocarbons. You will begin to assess the benefits and the risks of materials that you take for granted and use every day.

Section Preview/ Specific Expectations

In this section, you will

- **assess** some risks and benefits of hydrocarbon use in our society
- **communicate** your understanding of the following terms: *risk*, *benefit*, *risk-benefit analysis*, *greenhouse gases*, *global warming*, *sustainable development*

Figure 14.27 Polystyrene foam insulation, such as Styrofoam™, prevents heat flow and does not absorb water.

Figure 14.26 Materials such as Gore-Tex® fabric (made from a polymer of ethene) are important in the outdoor activities industry.

Figure 14.28 Artificial hip joints allow people who are disabled by arthritis to walk again.

CHECKPOINT

Hydrocarbons, such as fossil fuels, carry both risks and benefits. In a group, brainstorm to identify some risks and benefits.

Risks and Benefits

A **risk** is a chance of possible negative or dangerous results. Riding a bicycle carries the risk of falling off. Driving a car carries the risk of an accident. Almost everything you do has some kind of risk attached. Fortunately most risks are relatively small, and they may never happen. Many of the activities that carry risks also carry benefits. A **benefit** is an advantage, or positive result. For example, riding a bicycle provides the benefits of exercise, transportation, and enjoyment. When deciding to do an activity, it may be a good idea to compare the risks and benefits involved. (See Figure 14.29.)

Figure 14.29 Would you like to have a coal-burning power plant near your home? Some people might be upset by this idea because there is a health risk caused by pollution from the plant. Other people might think that a coal-burning power plant poses no threat at all. Who is right? How do you decide?

Risk-Benefit Analysis

Knowing more about an issue helps you assess its risks and benefits more accurately. How can you make the most informed decision possible? Follow these steps to do your own assessment of risks and benefits, called a **risk-benefit analysis**.

Step 1 Identify possible risks and benefits of the activity. Decide how to research these risks and benefits.

Step 2 Research the risks and benefits. You need information from reliable sources to make an accurate analysis.

Step 3 Weigh the effects of the risks and benefits. You may find that the risks are too great and decide not to do the activity. On the other hand, you may find that the benefits are greater than the risks.

Step 4 Compare your method for doing the activity with other possible methods. Do you use the safest method to do the activity? One method may be much safer than another.

In the next Sample Problem, you will see how a risk-benefit analysis can help you make informed decisions.

Sample Problem

Smoking

Problem

Many of your friends smoke, including some people you respect. Lately you have been thinking about taking up smoking too. Should you smoke? Perform a risk-benefit analysis to help you decide.

What Is Required?

You need to perform a risk-benefit analysis of smoking. This includes identifying, researching, and weighing the risks and benefits of smoking. It also includes looking at different methods of smoking, which might reduce the risk.

What Is Given?

The problem mentions that many of your friends smoke. This indicates a possible benefit of smoking—you will be imitating people

Continued ...

Continued ...
FROM PAGE 620

you like and respect. You need to do some research to identify more benefits and risks.

Plan Your Strategy

Use these four steps to identify and assess the risks and benefits of smoking.

Step 1 Identify possible risks and benefits.

Step 2 Research the risks and benefits. Use the Internet, reference books such as encyclopedias, and other sources to help you. It is important to choose reliable sources from which to obtain information.

Step 3 Weigh the risks and benefits.

Step 4 Compare different methods of smoking. Is one method less risky than another?

Act on Your Strategy

Step 1 Identify possible risks and benefits.

Possible risks	Possible benefits
Smoking is hazardous to your health.	Your friends smoke.
	Smoking may help you relax.

At this point, the benefits appear to outweigh the risks. You need to do more research, however, to make an informed decision.

Step 2 Research the risks and benefits.

Further research on the Internet provides more information on smoking risks:

Risks	Benefits
The tobacco smoke inhaled when smoking contains many toxic substances, such as carbon monoxide and ammonia gas.	Your friends smoke.
Smoking is the number one cause of lung cancer, heart disease, and emphysema.	Smoking may help you relax.
Smoking is addictive.	
About 50% of smokers end up dying from a tobacco-related disease.	
Second-hand smoke harms the people around you.	

Step 3 Weigh the risks and benefits of smoking.

The risks of smoking heavily outweigh the benefits.

Step 4 Compare different methods of smoking. Is one method less risky?

Smoking only one cigarette a day is less risky than smoking a pack a day. You cannot count on this method, however. Any kind of smoking is risky.

Continued ...

Continued ...
FROM PAGE 621

Practice Problems

25. Your town is considering dumping its plastic waste in a nearby lake. Identify possible risks and benefits for this plan. Explain where you could find more information to help your town make a decision.

26. Earth's reservoir of fossil fuels, including natural gas, will not last forever. As well, burning fossil fuels releases carbon dioxide (a greenhouse gas) and other pollutants that can cause acid rain. On the other hand, alternate energy sources, such as solar panels, are more expensive. They can also be less reliable than using fossil fuels. Perform a risk-benefit analysis to decide if you should heat your home using natural gas or solar panels. You will need to do more research to make an informed decision. (See the Internet Link on this page.)

Web LINK

www.school.mcgrawhill.ca/resources
To access more information on solar panels, natural gas heating, and the environmental effects of burning fossil fuels, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. Use this information to help you answer Practice Problem 26.

Hydrocarbons: A Risky Business?

How do hydrocarbons benefit our society? How do they affect the environment? What are the benefits and risks of using hydrocarbon fuels and petrochemicals? These are important questions for our global community. See Figure 14.30 for some ideas.

Figure 14.30

Hydrocarbon fuels have changed the way we live. Our dependence on them, however, has affected the world around us. The greenhouse effect, global warming, acid rain, and pollution are familiar topics on the news today. Our use of petroleum products, such as oil and gasoline, is linked directly to these problems.

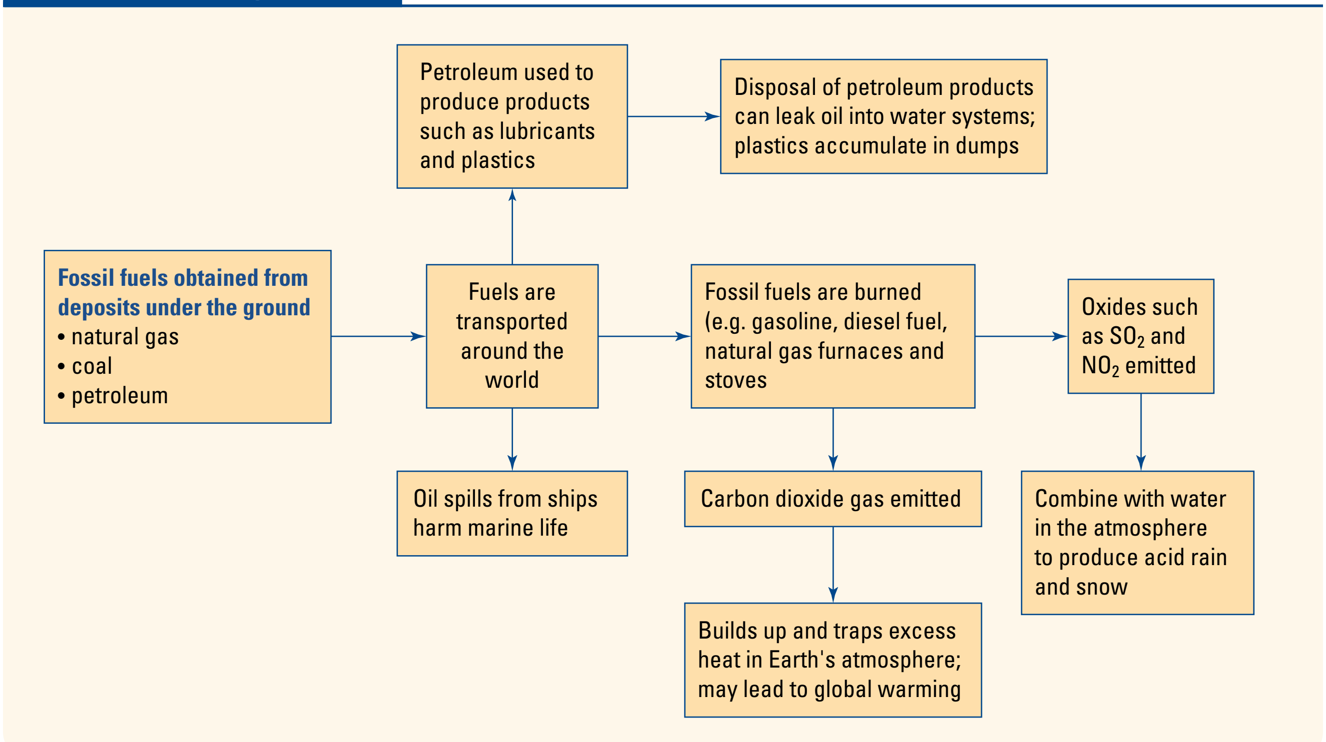
Web LINK

www.school.mcgrawhill.ca/resources/
To learn more about greenhouse gases, global warming, and acid rain, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

The Greenhouse Effect and Global Warming

Roads, expressways, service stations, and parking lots occupy almost 40% of Toronto. They are the result of our demand for fast and efficient transportation. Every day, Toronto's vehicles produce nearly 16 000 t of carbon dioxide by the combustion of fossil fuels. Carbon dioxide is an important greenhouse gas. **Greenhouse gases** trap heat in Earth's atmosphere and prevent the heat from escaping into outer space. Scientists think that a build-up of carbon dioxide in the atmosphere may lead to an increase in global temperature, known as **global warming**. The diagram below shows how these concepts are connected to fossil fuels.

Concept Organizer Hydrocarbons and the Enviroment



Acid Rain

The combustion of fossil fuels releases sulfur and nitrogen oxides. These oxides react with water vapour in the atmosphere to produce acid rain. Some lakes in northern Canada are "dead" because acid rain has killed the plants, algae, and fish that used to live in them. Forests in Québec and other parts of Canada have also suffered from acid rain.

Oil Spill Pollution

Our society demands a regular supply of fossil fuels. Petroleum is transported from oil-rich countries to the rest of the world. If an oil tanker carrying petroleum has an accident, the resulting oil spill can be disastrous to the environment.

Careers in Chemistry

Oil Spill Advisor

Developed nations, such as Canada, depend heavily on petroleum. Our dependence affects the environment in many ways. Oil spills are a dramatic example of environmental harm caused by petrochemicals. In the news, you may have seen oceans on fire and wildlife choked with tar. What can we do?

Obviously the best thing to do is to prevent oil spills from taking place. Stricter regulations and periodic inspections of oil storage companies help to prevent oil leakage. Once an oil spill has occurred, however, *biological*, *mechanical*, and *chemical* technologies can help to minimize harm to the environment.

Biological methods involve helpful micro-organisms that break down, or *biodegrade*, the excess oil. Mechanical methods depend on machines that physically separate spilled oil from the environment. For example, barriers and booms are used to contain an oil spill and prevent it from spreading. Materials such as sawdust are sprinkled on a spill to soak up the oil.

Two main chemical strategies are also used to clean up oil spills. In the first strategy, *gelling agents* are added to react with the oil. The reaction results in a bulky product that is easier to collect using mechanical methods. In the second strategy, *dispersing agents* break up oil into small droplets that mix with the water. This prevents the oil from reaching nearby shorelines. Dispersing agents work in much the same way as a bar of soap!

The scientific advisor for an oil spill response unit assesses a spill and determines the appropriate clean-up methods. She or he acts as part of a team of advisors. Most advisors have an M.Sc. or Ph.D. in an area of expertise such as organic chemistry, physical chemistry, environmental chemistry, biology, oceanography, computer modelling, or chemical engineering.

Oil spill response is handled by private and public organizations. All these organizations look for people with a background in chemistry. In fact, much of what you are learning about hydrocarbons can be related to oil spill response. Hydrocarbon chemistry can lead you directly to an important career, helping to protect the environment.

Make Career Connections

Create a technology scrapbook. Go through the business and employment sections in a newspaper. Cut out articles about clean-up technologies. What kinds of companies are doing this work? What can you learn about jobs in this field? What qualifications does a candidate need to apply for this type of job?

Web LINK

www.school.mcgrawhill.ca/resources/
Go to the web site above to learn about the famous oil spill caused by the *Exxon Valdez*. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. When and where did this oil spill occur? How did it affect the environment? What was done to clean it up?

Everyday Oil Pollution

The biggest source of oil pollution comes from the everyday use of oil by ordinary people. Oil that is dumped into water in urban areas adds to oil pollution from ships and tankers. In total, *three million tonnes* of oil reach the ocean each year. This is equivalent to having an oil spill disaster every day!

A student from Thornhill, Ontario, did a home experiment to discover how much oil remains in "empty" motor oil containers that are thrown out. He collected 100 empty oil containers from a local gas station. Then he measured the amount of oil that was left in each container. He found an average of 36 mL per container. Over 130 million oil containers are sold and thrown out in Canada each year. Using these figures, he calculated that nearly five million litres of oil are dumped into landfill sites every year, just in "empty" oil containers!

Once oil reaches the environment, it is almost impossible to clean up. Oil leaking from a landfill site can contaminate drinking water in the area. Because oil can dissolve similar substances, pollutants such as chlorine and pesticides, and other organic toxins, mix with the oil. They are carried with it into the water system, increasing the problem.

Solutions to Environmental Problems

All of the problems described above hinge on our use of fossil fuels. Thus, cutting back on our use of fossil fuels will help to reduce environmental damage. Cutting back on fossil fuels, however, depends on the consumers who buy petrochemicals and use fossil fuels. In other words, it depends on you and the people you know.

Corporations that are looking for profit have little incentive to change their use of fossil fuels. For example, the technology is available to build cars that can drive about 32 km on a single litre of fuel. Because this technology is not financially profitable, cars are still being produced that drive about 8 km per litre of fuel. If consumers demand and purchase more fuel-efficient cars, however, car manufacturers will have an incentive to produce such cars. Tougher government standards may also help to push the vehicle industry towards greater fuel efficiency.

Governments can also bring about change by endorsing the principle of sustainable development. This principle was introduced at the 1992 Earth Summit Conference. **Sustainable development** takes into account *the environment*, *the economy*, *and the health and needs of society*. (See Figure 14.31.)

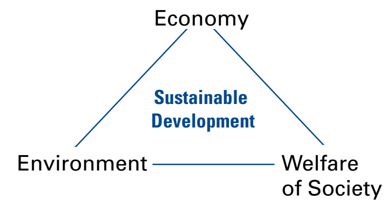


Figure 14.31 **Canada and other members of the United Nations endorse the principle of sustainable development. This principle states that the world must find ways to meet our current needs, without compromising the needs of future generations.**

Hydrocarbon fuels and products can benefit our society if they are managed well. They can cause great environmental damage, however, if they are managed irresponsibly. With enough knowledge, you can learn to make informed decisions on these important issues. Here are some suggestions of ways you can reduce your consumption of petroleum products. Why not choose one or more methods to practice? Or, brainstorm with your classmates to think of other ways to reduce consumption.

- Contact your local government and local power companies. Suggest using alternative fuels, such as solar energy and wind power.
- Ride a bicycle or walk more.
- Express your concerns by writing letters to the government or to newspapers.
- Become more informed by researching issues that concern you.
- Fix oil leaks in vehicles, and avoid dumping oil down the sink.
- If you are cold at home, put on an extra sweater instead of turning up the heat.
- Recycle and re-use petrochemical products, such as plastic shopping bags.
- Repair a broken item rather than buying a new one.

Section Wrap-up

In this section, you learned about some of the risks and benefits resulting from our use of fossil fuels. We obtain gasoline, heating oil, jet fuel, diesel fuel, fertilizers, and plastics from the oil and petroleum industry. Burning fossil fuels, however, produces carbon dioxide (a greenhouse gas) and other pollutants that lead to acid rain. Transporting oil also carries the risk of oil spills. Do the benefits of fossil fuels outweigh the risks? Complete these Section Review questions to help you decide.

Section Review

1. K/U You are about to try water-skiing for the first time. Should you try it? Will your activity affect the environment? How will you make your decision?
   - **(a)** Describe the steps you would follow to do a risk-benefit analysis.
   - **(b)** What are some possible risks and benefits?
   - **(c)** Where might you find the information you need to make a decision?
2. C Identify three benefits and three risks associated with the use of petroleum products and petrochemicals.
3. C Identify some steps that you can take to reduce your dependence on petroleum.
4. MC A construction company is planning to level a forest near your home to build a strip mall with a large parking lot. Many people enjoy walking in the forest, and many children play there. You have also observed wildlife, such as rabbits, snakes, frogs, and many kinds of birds living in the forest. In a group, brainstorm ways that the company could consider the environment and human welfare, as well as the economy. (Think of the economy as the owner of the property and the stores that will be built.)
5. MC Perform a risk-benefit analysis of the petroleum industry. Use the information in this section to help you identify possible risks and benefits. Use the Internet, or reference books to do more research on these risks and benefits.
6. MC Impure coal and gasoline contain nitrogen and sulfur compounds ($NO_{(g)}$ and $S_{(s)}$). The combustion of nitrogen and sulfur produces oxides that lead to acid rain.
   - **(a)** Balance the following six equations.

     $S_{8(s)} + O_{2(g)} \rightarrow SO_{2(g)}$

     $S_{8(s)} + O_{2(g)} \rightarrow SO_{3(g)}$

     $NO_{(g)} + O_{2(g)} \rightarrow NO_{2(g)}$

     $SO_{2(g)} + H_2O_{(\ell)} \rightarrow H_2SO_{3(aq)}$

     $SO_{3(g)} + H_2O_{(\ell)} \rightarrow H_2SO_{4(aq)}$

     $NO_{2(g)} + H_2O_{(\ell)} \rightarrow HNO_{3(aq)} + HNO_{2(aq)}$
   - **(b)** Suggest possible sources for the reactants $O_{2(g)}$ and $H_2O_{(\ell)}$.
   - **(c)** Explain how these equations show the production of acid rain.

CHAPTER 14 Review

Reflecting on Chapter 14

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Describe how fossil fuels are an important source of energy in our society.
- Describe how to write combustion equations, and how to measure the heat changes caused by physical processes such as dissolving.
- Describe the importance of isolating a system to reduce heat flow when measuring heat.
- Describe society's use of fossil fuels, and the resulting effects on the environment.
- Explain how to weigh the risks and benefits of an activity, and perform a risk-benefit analysis. This skill will help you make more informed decisions on issues that affect society, the economy, and the environment.

Reviewing Key Terms

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

benefit	bomb calorimeter
bond energy	Calorie
calorimeter	calorimetry
ΔT	combustion
complete combustion	endothermic
exothermic	fossil fuels
global warming	greenhouse gases
heat	heat capacity
heat of combustion	incomplete combustion
isolated system	potential energy
risk	risk-benefit analysis
specific heat capacity	sustainable development
temperature	thermal energy
thermal equilibrium	thermochemical equation

Knowledge/Understanding

1. (a) What are the products of the complete combustion of a hydrocarbon?
   (b) What products form if the combustion is incomplete?
2. (a) Why can incomplete combustion be dangerous if it occurs in your home?
   (b) How can you tell, by looking at a flame, that incomplete combustion is taking place?
3. Indicate whether each process is endothermic or exothermic.
   (a) water evaporating
   (b) a piece of paper burning
   (c) rubbing your hands together
   (d) clouds forming
4. How can a balanced thermochemical equation tell you whether a chemical reaction is exothermic or endothermic?
5. Describe the relationship between the amount of thermal energy that is released by water and
   (a) the mass of water
   (b) the temperature change of the water
6. Why is energy needed to sustain an endothermic reaction?
7. (a) The combustion of paraffin, $C_{25}H_{52(s)}$, is exothermic. Explain why by comparing the energy changes observed when chemical bonds are broken and formed.
   (b) The formation of 1-pentyne, $C_5H_{8(\ell)}$, is endothermic. Explain why by comparing the energy changes observed when chemical bonds are broken and formed.
8. 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.
   $$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)} + 484\text{ kJ}$$
   (a) Is this reaction exothermic or endothermic?
   (b) How much energy does the complete combustion of 1 g of hydrogen provide?
9. Write a balanced chemical equation for each reaction.
   (a) complete combustion of 4-methyl-1-pentene, $C_6H_{12(\ell)}$
   (b) incomplete combustion of benzene, $C_6H_{6(\ell)}$
   (c) incomplete combustion of propene, $C_3H_{6(g)}$
   (d) complete combustion of 3-ethyhexane, $C_8H_{18(\ell)}$
10. List three assumptions that you make when using a polystyrene calorimeter.

11. The same amount of heat is added to aluminum (c_{Al} = 0.900 J/g·°C) and nickel (c_{Ni} = 0.444 J/g·°C). Which metal will have a greater temperature increase? Explain.

12. Does propane burning in an outdoor barbecue have a negative or positive heat of combustion? Explain.

Inquiry

13. To make four cups of tea, 1.00 kg of water is heated from 22.0°C to 99.0°C. How much energy is needed?

14. Two different foods are burned in a calorimeter. Sample 1 has a mass of 6.0 g and releases 25 Cal of heat. Sample 2 has a mass of 2.1 g and releases 9.0 Cal of heat. Which food releases more heat per gram?

15. 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?

16. A substance is burned completely in a bomb calorimeter. The temperature of the 2000 g of water in the calorimeter rises from 25.0°C to 43.9°C. How much energy is released?

17. A horseshoe can be shaped from an iron bar when the iron is heated to temperatures near 1500°C. The hot iron is then dropped into a bucket of water and cooled. An iron bar is heated from 1500°C and then cooled in 1000 g of water that was initially at 20.0°C. How much heat energy does the water absorb if its final temperature is 65.0°C?

18. A group of students decide 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.
   (a) Calculate the heat energy that is released by the popcorn. Express your answer in both kilojoules and Calories per gram of popcorn.
   (b) Another student tells the group that she has read the label on the popcorn bag. The label states that 30 g of popcorn yields 110 Cal. What is this value in Calories per gram? How can you account for the difference between the two values?

19. In Chapters 13 and 14, you have examined many properties of hydrocarbons. Describe one physical property and one chemical property of hydrocarbons. Explain how these two properties vary from one hydrocarbon to another. Describe how you might measure each property in a lab.

20. A reaction in a calorimeter causes 250.0 g of water to decrease in temperature by 2.40°C. How much heat did the reaction absorb?

21. A chemist wants to calibrate a new bomb calorimeter. He completely burns a mass of 0.930 g of carbon in a calorimeter. The temperature of the calorimeter changes from 25.00°C to 28.15°C. If the thermal energy change is 32.8 kJ/g of carbon burned, what is the heat capacity of the new calorimeter? What evidence shows that the reaction was exothermic?

22. 200 g of iron at 350°C is added to 225 g of water at 10.0°C. What is the final temperature of the iron-water mixture?

23. In this chapter, you learned that fats have long hydrocarbon sections in their molecular structure. Therefore, they have many C—C and C—H bonds. Sugars have fewer C—C and C—H bonds but more C—O bonds. Use Table 14.1 in this chapters. Explain why you can obtain more energy from burning a fat than from burning a sugar.

24. 2,2,4-trimethylpentane, an isomer of C_8H_{18}, is a major component of gasoline.
   (a) Write the balanced thermochemical equation, using the word "energy," for the complete combustion of this compound. Use $C_8H_{18(\ell)}$ as the formula for the compound.
   (b) What is the ideal ratio of fuel to air for this fuel?
   (c) In the previous unit, you learned how to solve problems involving gases. Calculate the volume of carbon dioxide, at 20.0°C and 105 kPa, that is produced from the combustion of 1.00 L of $C_8H_{18(\ell)}$. **Note:** The density of $C_8H_{18(\ell)}$ is 0.69 g/mL.

25. 100 g of calcium carbide is used to produce acetylene in a laboratory, as you did in Investigation 14-A.

(a) What volume of water (density 1.00 g/mL) is needed to react completely with the calcium carbide?

(b) What volume of acetylene gas will be produced at STP when the calcium carbide and water are mixed?

Communication

26. Earlier in the chapter, you learned that poisonous carbon monoxide can form during incomplete hydrocarbon combustion. The use of carbon monoxide detectors in homes and businesses has reduced the number of deaths due to carbon monoxide poisoning. Are all carbon monoxide detectors the same? Telephone your local fire department, go to a library, or search the Internet to find out about carbon monoxide detectors.

27. Design a poster or a brochure to explain the concept of sustainable development to a student in a much younger grade.

28. Prepare a concept map to illustrate the effects of hydrocarbons on our society and the environment.

Making Connections

29. When energy is wasted during an industrial process, what actually happens to this energy?

30. Look at Table 14.3. Compare caramels with raw carrots. Which food gives more Calories per gram?

31. Petrochemical products, such as plastics, have affected your life. Identify one benefit and one possible risk associated with the use of petrochemical products.

32. Define sustainable development. Suggest a condition that you feel society must agree on to achieve sustainable development.

33. The Hibernia Oil Field is located off the Grand Banks of Newfoundland, on Canada's east coast. It started oil production in the fall of 1997. Research and write a report on some of the risks and benefits of this massive oil operation. Consider ecological, economic, and social issues.

34. On an episode of *The Nature of Things*, Dr. David Suzuki made the following comment: "As a society and as individuals, we're hooked on it [oil]." Discuss his comment. Explain how our society has benefitted from hydrocarbons. Describe some of the problems that are associated with the use of hydrocarbons. Also describe some possible alternatives for the future.

Answers to Practice Problems and Short Section Review Questions:

Practice Problems **1.(a)** complete **2.(a)** $C_5H_{12} + 8O_2 \rightarrow 5Co_2 + 6H_2O$ **(b)** $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$ **(c)** $6C_2H_6 + 15O_2 \rightarrow 4CO_2 + 4CO + 4C + 18H_2O$, $3C_2H_6 + 6O_2 \rightarrow CO_2 + CO + 4C + 9H_2O$, **3.(a)** $C_4H_{10} + 4O_2 \rightarrow CO_2 + CO + 2C + 5H_2O$ **(b)** $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$ **4.** $C_{25}H52 + 38O_2 \rightarrow 25CO_2 + 26H_2O$ **5.(a)** $C_{13}H_{28} + 9O_2 \rightarrow CO_2 + 2CO + 10C + 14H_2O$, $C_{13}H_{28} + 10O_2 \rightarrow 2CO_2 + 2CO + 8C + 14H_2O$ **(c)** incomplete **6.(a)** $3C_{(s)} + 4H_{2(g)} \rightarrow C_3H_{8(g)} +$ energy **(b)** $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)} +$ energy **7.(b)** energy $+ 2C_{(s)} + 2H_{2(g)} \rightarrow C_2H_{4(g)}$, $C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(g)} +$ energy **8.** 6.2×10^3 J **9.** -4.6×10^3 J **11.(a)** +4.22 kJ, +4.6 kJ **12.(a)** $m = Q/c\Delta T$ **(b)** $c = Q/m\Delta T$ **(c)** $\Delta T = Q/mc$ **13.(a)** $T_I = T_f - Q/mc$ **(b)** $T_f = T_I + Q/mc$ **14.** 2.14×10^5 J **15.** 0.7900 J/g•°C, granite **16.** 12.1 J/g•°C **17.** 15.5•°C **18.** -2.30×10^3 J **19.** 1.26×10^4 J **20.** 661 J/g **21.(a)** -2.2×10^4 J, 2.2×10^4 J **(b)** 8.0 J/g•°C **22.(a)** $C = Q/\Delta T$ **(b)** $\Delta T = Q/C$ **(c)** $T_f = T_I + Q/C$ **(d)** $T_i = T_f - Q/C$ **23.** −20.3 kJ/g **24.(a)** 100 g **(b)** 8.75 kJ/°C

Section Review: 14.1: **3.(a)** $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$ **(b)** $C_5H_{10} + 6O_2 \rightarrow 3CO_2 + CO + C + 5H_2O$ **4.** $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, $4CH_4 + 6O_2 \rightarrow CO_2 + 2CO + C + 8H_2O$ **6.(a)** $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$ **(b)** 54 g **(c)** 392 L **14.2:** **1.(a)** endothermic, net energy is absorbed, formation of ethyne (acetylene) **(b)** exothermic, net energy is released, combustion of methane **4.(a)** endothermic **5.** 197 kJ/mol **14.3:** **3.(a)** -4.54×10^3 J **4.(a)** 38.1 J **(b)** 76.2 J **(c)** 76.2 J **5.(a)** 0.237 J/g•°C **(b)** 3.43°C **6.** copper **14.4:** **3.** −39.8 kJ **4.** -3.1×10^3 J **5.** 9.63 Cal, yes **14.5:** **6.(a)** $S_8 + 8O_2 \rightarrow 8SO_2$, $S_8 + 12O_2 \rightarrow 8SO_3$, $2NO + O_2 \rightarrow 2NO_2$, $SO_2 + H2O \rightarrow H_2SO_3$, $SO_3 + H_2O \rightarrow H_2SO_4$, $2NO_2 + H_2O \rightarrow HNO_3 + HNO_2$