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.

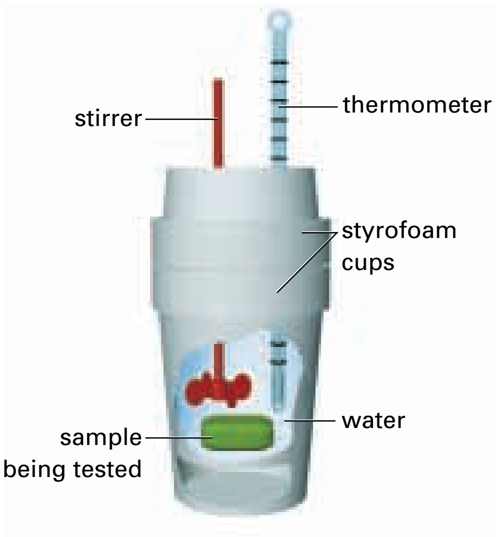

Figure 14.14 **A polystyrene (coffee cup) calorimeter usually consists of two nested polystyrene cups with a polystyrene lid, to provide insulation from the surroundings.**

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.

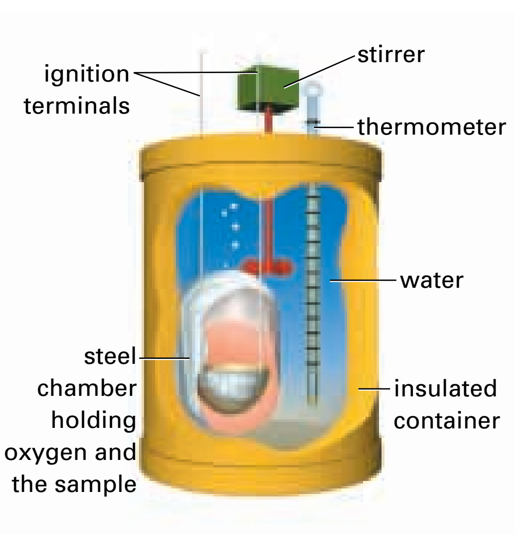

Figure 14.15 **A bomb calorimeter is more sophisticated than a polystyrene calorimeter.**

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.

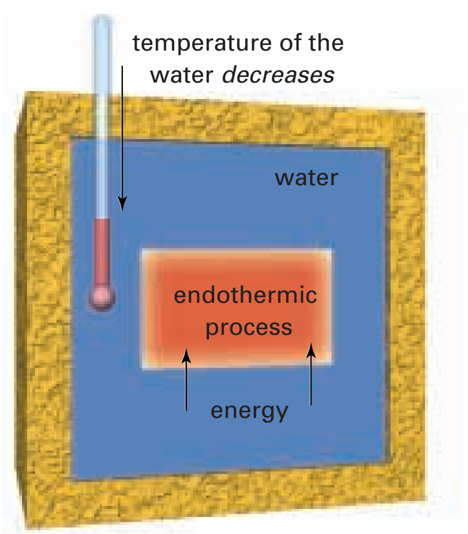

Figure 14.16 An endothermic process, such as ice melting

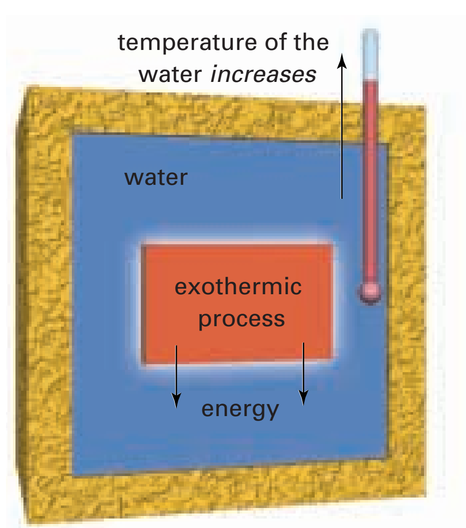

Figure 14.17 An exothermic process, such as the combustion of propane

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.

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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.

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

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.)

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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
= (151 L)(1 kg/L)
= 151 kg

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

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

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) $Q_m = -Q_w$
$= -2.81 \times 10^4$ 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.

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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.

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

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

Act on Your Strategy

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

The calorimeter gained 418 kJ of thermal energy.

$$Q_{sample} = -Q_{cal} = -418 \text{ kJ}$$

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

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

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, $Q_{cal} = C\Delta T$. Recall that $\Delta T = T_f = T_i$. 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.

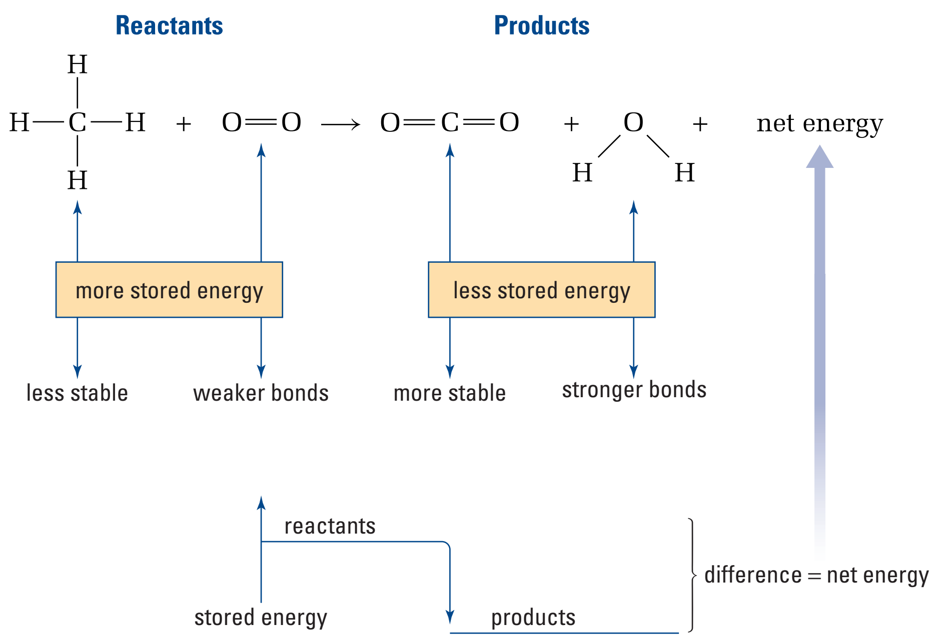


Figure 14.24 **Energy stored in bonds**

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

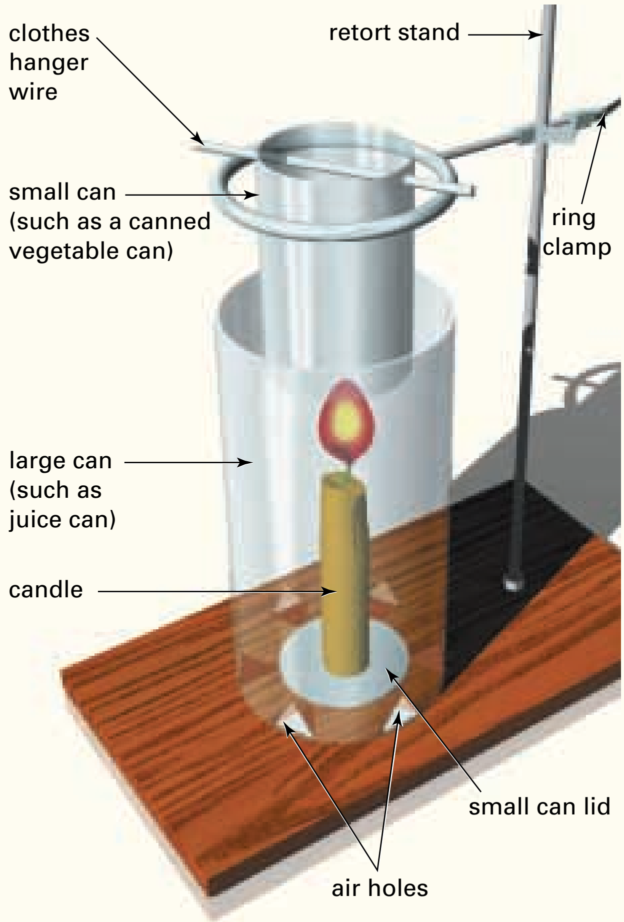
- Tie back long hair and confine any loose clothing. Before you light the candle, check that there are no flammable solvents near by.

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?