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Warmth from Fat

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Heat released by a burning match is absorbed by air.

## DISCOVER IT!

OBSERVING  
HEAT FLOW

You need a clean, medium-sized rubber band.

1. Hook your index fingers through each end of the rubber band. Without stretching the rubber band, place it against your upper lip or forehead. Note the temperature of the rubber band.
2. Move the rubber band away from your skin, quickly stretch and hold it, and then place it back against your skin. Note any temperature change.
3. Fully stretch the rubber band, and then allow it to return to its original shape. Place it against your skin and note any temperature change.
4. Repeat Steps 2 and 3 until you are certain of the temperature change in each step.

Did the rubber band feel cool or warm after it was stretched in Step 2? Did the rubber band feel cool or warm after it returned to its original shape in Step 3? Think about the temperature changes you observed and form some initial answers to the following questions. What is heat? In what direction does heat flow? As you work through this chapter, return to these questions.

# THE FLOW OF ENERGY—HEAT



*L*ava flowing out of an erupting volcano is very hot. Its temperature ranges from 550 °C to 1400 °C. As lava flows down the side of a volcano, it loses heat and begins to cool slowly. In some instances, the lava may flow into the ocean

where it cools rapidly. **Why does lava cool more quickly in water than on land?**

## Energy Transformations

Glowing campfires, the sun's rays, and rubbing your hands together all produce heat. However, other activities, such as melting ice and boiling water, absorb heat. **Thermochemistry** is concerned with the heat changes that occur during chemical reactions. In this chapter, you will examine heat and its effects on a number of chemical and physical processes. First, however, it is important to understand energy transformations.

When you buy gasoline, you are buying the stored potential energy it contains. This energy is used to do work, most often to propel a car. The controlled explosions of the gasoline in the car's engine transform the potential energy into useful work. Work is done when a force is used to move an object. **Energy** is the capacity for doing work or supplying heat. Unlike matter, energy is weightless, odorless, and tasteless. Energy is detected only because of its effects. Energy stored within the structural units of chemical substances is called **chemical potential energy**. Gasoline contains a significant amount of chemical potential energy. Different substances store different amounts of energy. The kinds of atoms and their arrangement in the substance determine the amount of energy stored in the substance.

**Heat**, represented by  $q$ , is energy that transfers from one object to another because of a temperature difference between them. Heat, itself, cannot be detected by the senses or by instruments. Only changes caused by heat can be detected. One of the effects of adding heat is a rise in the temperature of objects. It is the radiant heat of the sun's rays that makes a summer day hot. In this example, air is the object that absorbs heat and

### objectives

- ▶ Explain the relationship between energy and heat
- ▶ Distinguish between heat capacity and specific heat

### key terms

- ▶ thermochemistry
- ▶ energy
- ▶ chemical potential energy
- ▶ heat
- ▶ system
- ▶ surroundings
- ▶ universe
- ▶ law of conservation of energy
- ▶ endothermic process
- ▶ exothermic process
- ▶ calorie
- ▶ joule
- ▶ heat capacity
- ▶ specific heat capacity
- ▶ specific heat

**Figure 11.1**

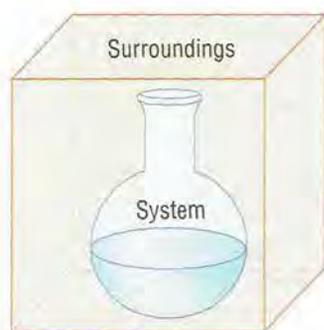
*Chemical potential energy is stored within the bonds of gasoline molecules (a). As the gasoline burns, the energy is released and is used to do work. In this example, the work is to propel race cars around the track (b).*



(a)



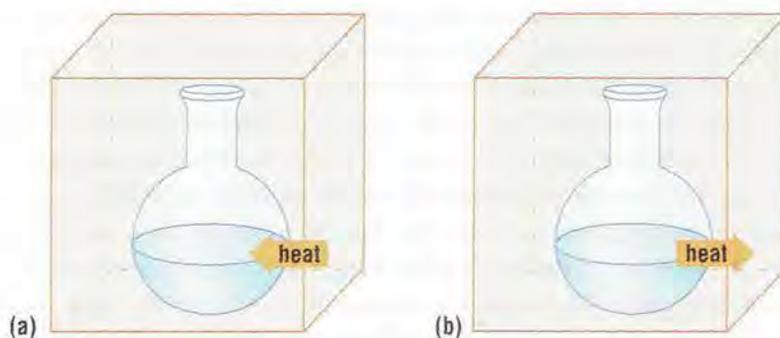
(b)


**Figure 11.2**

The part of the universe being studied is the system. What constitutes the surroundings?

**Figure 11.3**

There are two directions in which the heat of a system can flow. (a) In an endothermic process, heat flows into the system from the surroundings. The system absorbs heat. (b) In an exothermic process, heat flows from the system to the surroundings. The system loses heat.



In thermochemical calculations the direction of the heat flow is given from the point of view of the system. Look at **Figure 11.3a**. Heat flowing into a system from its surroundings is defined as positive;  $q$  has a positive value. A process that absorbs heat from the surroundings is called an **endothermic process**. In an endothermic process, the system gains heat as the surroundings cool down. In **Figure 11.3b**, heat flows out of the system into its surroundings. This type of heat flow is given a negative value;  $q$  is negative because the system is losing heat. A process that releases heat to its surroundings is called an **exothermic process**. In an exothermic process, the system loses heat as the surroundings heat up. **Table 11.1** explains the sign convention for heat changes.

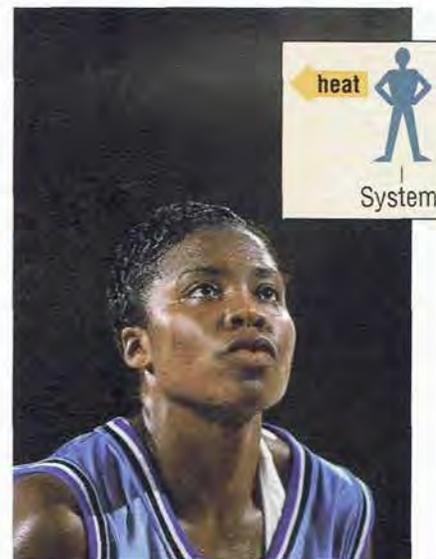
**Table 11.1**

Heat Change Sign Convention		
Direction of heat flow	Sign	Reaction type
Heat flows out of the system	Heat change $< 0$ (negative)	Exothermic
Heat flows into the system	Heat change $> 0$ (positive)	Endothermic

increases in temperature. Heat always flows from a warmer object to a cooler object. If two objects remain in contact, heat will flow from the warmer object to the cooler object until the temperature of both objects is the same.

## Exothermic and Endothermic Processes

Essentially all chemical reactions and changes in physical state involve either the release or the absorption of heat. In studying heat changes, it is useful to define a **system** as the part of the universe on which you focus your attention. The **surroundings** include everything else in the universe. In thermochemical experiments, it is a good approximation to consider the region in the immediate vicinity of the system as the surroundings. In **Figure 11.2**, for example, the mixture of chemicals undergoing a reaction is the system, and everything else is the surroundings. Together, the system and its surroundings constitute the **universe**. A major goal of studying thermochemistry is to examine the flow of heat from the system to its surroundings, or the flow of heat from the surroundings to the system. The **law of conservation of energy** states that in any chemical or physical process, energy is neither created nor destroyed. All of the energy involved in a process can be accounted for as work, stored energy, or heat.



**Figure 11.4**

*A fire helps keep you warm when you are out in the cold. If your body is the system, what is the fire? The human body cools itself by giving off heat when perspiration evaporates. What is the system here? Which of these processes is exothermic and which is endothermic?*

In the photograph on the left in **Figure 11.4**, the system (the people) gains heat from its surroundings (the fire). As shown in the inset illustration, heat flows into the system from its surroundings. What kind of process is this? In the photograph on the right, the system (the body) cools as perspiration evaporates from the skin and heat flows to the surroundings. What kind of process is shown in this illustration?

## Heat Capacity and Specific Heat

You have probably heard of someone exercising to burn off fat and calories. What does it mean to “burn calories”? During exercise your body generates heat, and this heat is measured in units called calories. The heat is generated as your body breaks down sugars and fats into carbon dioxide and water. Although there is not an actual fire burning the sugars and fats within your body, chemical reactions accomplish the same result. In breaking down 10 g of sugar, for example, your body generates a certain amount of heat. The same amount of heat would be produced if 10 g of sugar were completely burned in a fire, producing carbon dioxide and water.

A **calorie** is defined as the quantity of heat needed to raise the temperature of 1 g of pure water 1 °C. There is an important difference, however, between a calorie and a Calorie. The calorie, written with a small c, is defined above and is used except when referring to the energy contained in food. The dietary Calorie, written with a capital C, always refers to the energy in food. One dietary Calorie is actually equal to one kilocalorie, or 1000 calories.

$$1 \text{ Calorie} = 1 \text{ kilocalorie} = 1000 \text{ calories}$$

The statement “10 g of sugar has 41 Calories” means that 10 g of sugar releases 41 kilocalories of heat when completely burned to produce carbon dioxide and water.

**Figure 11.5**

Which food will warm you up more when you are cold: a bowl of hot soup, or two slices of hot buttered toast? Why?



## LINK TO PHYSIOLOGY

### Dietary Calories

Your proper caloric intake depends on your level of physical activity. In an eight-hour day at a desk, you burn about 800 Calories. This is about the number of Calories in two helpings of spaghetti. When exercising, however, you become a relative biochemical blast furnace. In vigorous activities such as running and jumping, you expend 7–10 Calories per minute, or 420–600 Calories per hour. At these rates, a runner who covers a 26-mile marathon course in 3 hours might expend 1800 Calories, or the equivalent of 4.5 helpings of spaghetti.

The calorie is also related to the **joule**, the SI unit of heat and energy named after the English physicist James Prescott Joule (1818–1889). A joule is slightly less than one-fourth of a calorie. One joule of heat raises the temperature of 1 g of pure water 0.2390 °C. You can convert between calories and joules using the following relationships.

$$1 \text{ J} = 0.2390 \text{ cal} \quad 4.184 \text{ J} = 1 \text{ cal}$$

The amount of heat needed to increase the temperature of an object exactly 1 °C is the **heat capacity** of that object. The heat capacity of an object depends on its mass as well as its chemical composition. The greater the mass of the object, the greater its heat capacity. A massive steel girder, for example, requires much more heat to raise its temperature 1 °C than a small steel nail does. Similarly, a cup of water has a much greater heat capacity than a drop of water. Besides varying with mass, the heat capacity of an object also depends on its chemical composition. It follows that different substances with the same mass may have different heat capacities.

On a sunny day, a 20 kg puddle of water may be cool, while a nearby 20 kg iron sewer cover may be too hot to touch. This situation illustrates how different heat capacities affect the temperature of objects. Assuming that both the water and the iron absorb the same amount of radiant energy from the sun, the temperature of the water changes less than the temperature of the iron because the specific heat capacity of water is larger.

**Table 11.2**

Specific Heat Capacities of Some Common Substances		
Specific heat capacity		
Substance	J/(g × °C)	cal/(g × °C)
Water	4.18	1.00
Grain alcohol	2.4	0.58
Ice	2.1	0.50
Steam	1.7	0.40
Chloroform	0.96	0.23
Aluminum	0.90	0.21
Glass	0.50	0.12
Iron	0.46	0.11
Silver	0.24	0.057
Mercury	0.14	0.033

The **specific heat capacity**, or simply the **specific heat**, of a substance is the amount of heat it takes to raise the temperature of 1 g of the substance 1 °C. What is the relationship between specific heat and heat capacity? Table 11.2 gives specific heats for some common substances. Water has a very high specific heat compared with the other substances in the table. You can see from the table that one calorie of heat raises the temperature of 1 g of water 1 °C. Metals, however, have low specific heats. One calorie of heat raises the temperature of 1 g of iron 9 °C. Thus water has a specific heat nine times that of iron. Heat affects the temperature of objects with a high specific heat much less than the temperature of those with a low specific heat. Just as it takes a lot of heat to raise the temperature of water, water also releases a lot of heat as it cools. Water in lakes and oceans absorbs heat from the air on hot days and releases it back into the air on cool days. As illustrated in Figure 11.6, this property of water is responsible for moderate climates in coastal areas. The specific heat of water is often used by farmers to protect their crops. In freezing weather, citrus crops are often sprayed with water to protect the fruit from damage. As the water freezes, it releases heat, which helps to prevent the fruit from freezing. The results of this procedure are shown in Figure 11.7.

To calculate the specific heat of a substance, you divide the heat input by the temperature change times the mass of the substance. The equation for specific heat ( $C$ ) follows, where  $q$  is heat and  $m$  is mass. The symbol  $\Delta T$  (read “delta T”) in the equation represents the change in temperature.  $\Delta T$  is calculated from the equation  $\Delta T = T_f - T_i$ , where  $T_f$  is the final temperature and  $T_i$  is the initial temperature.

$$C = \frac{q}{m \times \Delta T} = \frac{\text{heat (joules or calories)}}{\text{mass (g)} \times \text{change in temperature (}^\circ\text{C)}}$$

As you can see from this equation, specific heat may be expressed in terms of joules or calories. Therefore, the units of specific heat are either  $\text{J}/(\text{g} \times ^\circ\text{C})$  or  $\text{cal}/(\text{g} \times ^\circ\text{C})$ .



**Figure 11.6**

*San Francisco is located on the Pacific coast. The high specific heat of the ocean helps keep the temperature in San Francisco much more moderate than that of towns and cities farther inland.*

**Figure 11.7**

*Water must give off a lot of heat in order to freeze. Sometimes farmers use water's high specific heat capacity to their advantage. In freezing weather, orange groves are often sprayed with water to protect the fruit from frost damage.*

## REVIEW OF SIGNIFICANT FIGURES

This page offers you a chance to refresh your knowledge of significant figures before beginning this chapter—a chapter which involves solving problems containing numerous numerical values. As you learned in Section 3.2, a value in science must be reported using the correct number of *significant figures* (or digits). Answers cannot be more precise than the given data, as they would be misleading.

The rules for determining significant figures are summarized below. (Note that this list is a condensed version of the rules listed in Chapter 3.)

1. Nonzero digits are always significant.
2. A zero is significant only if it is
  - a. at the right end of a number and after a decimal point, or
  - b. between digits that are significant according to rule 1 or 2a.

Zeros to the left of nonzero digits or at the end of a quantity written as a whole number are “placeholders” and are *not* significant.

3. If a quantity is known to be exact, it has an unlimited number of significant figures.

4. If a quantity is written in scientific notation, all digits of the coefficient are significant.

For example, the significant digits in each number below are shown in blue.

30 400	3 significant figures
150.0	4 significant figures
2401	4 significant figures
168.030	6 significant figures
0.0058	2 significant figures
$3.010 \times 10^8$	4 significant figures

When working with significant digits, *round up* the final significant digit *only if* the next digit is 5 or greater. When you do a calculation, round your final answer (not the intermediate steps) according to these rules:

**Addition and Subtraction** Round the result to the same number of decimal places as the measurement with the fewest decimal places.

**Multiplication and Division** Round the result to the same number of significant figures as the measurement with the fewest significant figures.

## Example 1

Evaluate  $2.34 + 1.2$  and express the answer with the correct number of significant figures.

2.34 has 2 decimal places and 1.2 has 1 decimal place. So, the answer should be rounded to 1 decimal place.

Without rounding,  $2.34 + 1.2 = 3.54$ .

After rounding 3.54 to 1 decimal place, the final answer is 3.5.

## Example 2

Use the formula  $C = \frac{q}{m \times \Delta T}$  to calculate  $C$ , the specific heat capacity, if  $q = 516 \text{ J}$ ,  $m = 6.4 \text{ g}$ , and  $\Delta T = 25.80 \text{ }^\circ\text{C}$ .

Because 6.4 g has only 2 significant figures, the answer should be rounded to 2 significant figures.

$$C = \frac{516 \text{ J}}{6.4 \text{ g} \times 25.80 \text{ }^\circ\text{C}} = 3.125 \frac{\text{J}}{\text{g} \times \text{ }^\circ\text{C}} = 3.1 \frac{\text{J}}{\text{g} \times \text{ }^\circ\text{C}}$$

## Practice Problems

Find the number of significant figures in each quantity.

- A. 1340    B. 0.06    C.  $3.400 \times 10^4$     D. 0.00350    E. 16.0

Round each answer to the appropriate number of significant figures.

- F.  $16.382 + 17.5$     I.  $317.04 \div 18.7$     L. Use the formula  $C = \frac{q}{m \times \Delta T}$  to calculate  $C$  if  $q = 14.80 \text{ J}$ ,  $m = 3.056 \text{ g}$ , and  $\Delta T = 10.01 \text{ }^\circ\text{C}$ .
- G.  $1.4 \times 6.03$     J.  $(6.030 \times 10^7) + (1.64 \times 10^5)$
- H.  $128.0 - 64.37$     K.  $(3.0 \times 10^{15}) \div (2.19 \times 10^4)$

## Sample Problem 11-1

The temperature of a piece of copper with a mass of 95.4 g increases from 25.0 °C to 48.0 °C when the metal absorbs 849 J of heat. What is the specific heat of copper?

1. **ANALYZE** List the knowns and the unknown.

Knowns:

- $m_{\text{Cu}} = 95.4 \text{ g}$
- $\Delta T = (48.0 \text{ °C} - 25.0 \text{ °C}) = 23.0 \text{ °C}$
- $q = 849 \text{ J}$

Unknown:

$$\bullet C_{\text{Cu}} = ? \frac{\text{J}}{\text{g} \times \text{°C}}$$

Use the known values and the definition of specific heat,

$$C = \frac{q}{m \times \Delta T}, \text{ to calculate the unknown value } C_{\text{Cu}}.$$

2. **CALCULATE** Solve for the unknown.

Substitute the known values into the equation for the specific heat and solve.

$$C_{\text{Cu}} = \frac{q}{m \times \Delta T}$$

$$C_{\text{Cu}} = \frac{849 \text{ J}}{95.4 \text{ g} \times 23.0 \text{ °C}} = 0.387 \frac{\text{J}}{\text{g} \times \text{°C}}$$

3. **EVALUATE** Does the result make sense?

Remember that water has a very high specific heat (4.18 J/(g × °C)). Metals, however, have low specific heats—values less than 4.18 J/(g × °C). Thus the calculated value of 0.387 J/(g × °C) seems reasonable.

## Practice Problems

1. When 435 J of heat is added to 3.4 g of olive oil at 21 °C, the temperature increases to 85 °C. What is the specific heat of olive oil?
2. A 1.55-g piece of stainless steel absorbs 141 J of heat when its temperature increases by 178 °C. What is the specific heat of the stainless steel?
3. How much heat is required to raise the temperature of 250.0 g of mercury 52 °C?

## Chem ASAP!

## Problem-Solving 3

Solve Problem 3 with the help of an interactive guided tutorial.



## section review 11.1

4. Define energy and explain how energy and heat are related.
5. Explain the difference between heat capacity and specific heat.
6. Will the specific heat of 50 g of a substance be the same as, or greater than, the specific heat of 10 g of the same substance?
7. On a sunny day, why does the concrete deck around an outdoor swimming pool become hot, while the water stays cool?
8. Using calories, calculate how much heat 32.0 g of water absorbs when it is heated from 25.0 °C to 80.0 °C. How many joules is this?
9. A chunk of silver has a heat capacity of 42.8 J/°C. If the silver has a mass of 181 g, calculate the specific heat of silver.
10. How many kilojoules of heat are absorbed when 1.00 L of water is heated from 18 °C to 85 °C?



**Chem ASAP! Assessment 11.1** Check your understanding of the important ideas and concepts in Section 11.1.

# MEASURING AND EXPRESSING HEAT CHANGES



## objectives

- ▶ Construct equations that show the heat changes for chemical and physical processes
- ▶ Calculate heat changes in chemical and physical processes

## key terms

- ▶ calorimetry
- ▶ calorimeter
- ▶ enthalpy ( $H$ )
- ▶ thermochemical equation
- ▶ heat of reaction
- ▶ heat of combustion

*As you know, a burning match gives off heat. When you strike a match, heat is released to the surroundings in all directions. As you have learned, heat cannot be detected by the senses or by instruments. Is there a way to measure exactly how much heat is released from a burning match?*

## Calorimetry

Energy changes occur in many systems, from the inner workings of a clock, to the eruption of volcanoes, to the formation of the solar system. Most chemical and physical changes you will encounter occur at constant atmospheric pressure. For example, a reaction in an open beaker, the formation of ice in a lake, and the reactions in many living organisms all occur at constant atmospheric pressure. By defining a thermodynamic variable called enthalpy—a variable that takes constant pressure into account—you can measure the energy changes that accompany chemical and physical processes.

Heat that is released or absorbed during many chemical reactions can be measured by calorimetry. **Calorimetry** is the accurate and precise measurement of heat change for chemical and physical processes. In calorimetry, the heat released by the system is equal to the heat absorbed by its surroundings. What law describes this relationship? To measure heat changes accurately and precisely, the processes must be carried out in an insulated container. The insulated device used to measure the absorption or release of heat in chemical or physical processes is called a **calorimeter**.

Foam cups, which keep hot drinks hot and cold drinks cold, are excellent heat insulators. Because they do not let much heat in or out, they can be used as simple calorimeters. In fact, the heat change for many chemical reactions can be measured in a constant-pressure calorimeter similar to the one shown in **Figure 11.8**. Because most chemical reactions and physical

**Figure 11.8**

*A simple constant-pressure calorimeter is shown here. In a calorimeter, the thermometer measures the temperature change of the chemicals as they react in water. The stirrer is used to keep the solution at a uniform temperature. The chemical substances that react in solution constitute the system. Is the water, in which the chemicals dissolved, part of the system or part of the surroundings?*

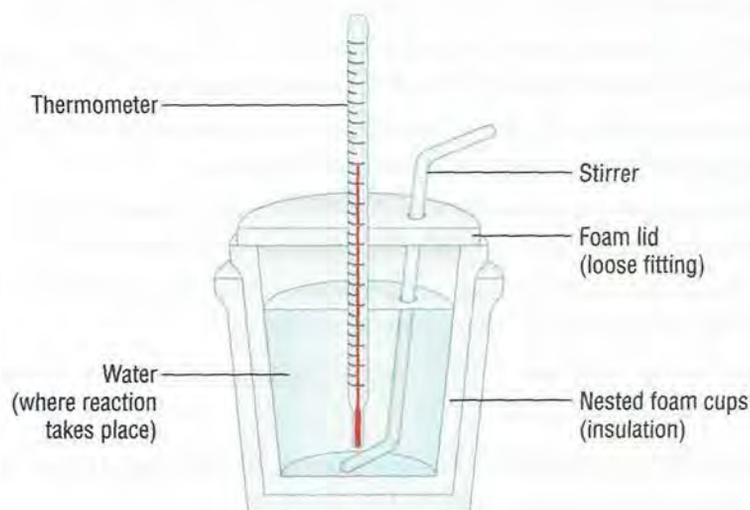


Table 11.3

Enthalpy Sign Convention	
Exothermic reaction	$\Delta H$ is negative ( $\Delta H < 0$ )
Endothermic reaction	$\Delta H$ is positive ( $\Delta H > 0$ )

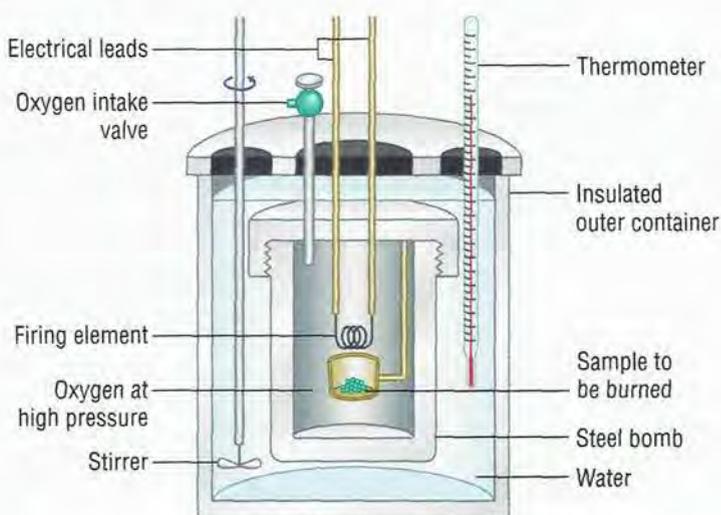
changes carried out in the laboratory are open to the atmosphere, these changes occur at constant pressure. For systems at constant pressure, the heat content is the same as a property called the **enthalpy** ( $H$ ) of the system. Heat changes for reactions carried out at constant pressure are the same as changes in enthalpy, symbolized as  $\Delta H$  (read "delta H"). Because the reactions presented in this textbook occur at constant pressure, the terms heat and enthalpy are used interchangeably. In other words,  $q = \Delta H$ . Recalling the equation for specific heat, we can write the following relationship for the heat change in a chemical reaction carried out in aqueous solution.

$$q = \Delta H = m \times C \times \Delta T$$

$\Delta H$  is the heat change;  $m$  is the mass of the water;  $C$  is the specific heat capacity of water; and  $\Delta T = T_f - T_i$ . The sign of  $\Delta H$  is negative for an exothermic reaction and positive for an endothermic reaction. Table 11.3 summarizes the sign convention for enthalpy.

To measure the heat change for a reaction in aqueous solution in a foam cup calorimeter, you dissolve the reacting chemicals (the system) in known volumes of water (the surroundings). Then measure the initial temperature of each solution and mix the solutions in the foam cup. After the reaction is complete, measure the final temperature of the mixed solutions. Because you know the initial and final temperatures and the heat capacity of water, you can calculate the heat released or absorbed in the reaction using the equation for specific heat.

Calorimetry experiments can also be performed at constant volume using a device called a bomb calorimeter. A bomb calorimeter, similar to the one shown in Figure 11.9, measures the heat released from burning a compound. The calorimeter is a closed system; that is, the mass of the system is constant.



## Chem ASAP!

## Simulation 8



Simulate a combustion reaction and compare the  $\Delta H$  results for several compounds.

Figure 11.9

In a bomb calorimeter, a sample is burned in a constant-volume chamber in the presence of oxygen at high pressure. The heat that is released warms the water surrounding the chamber. By measuring the temperature increase of the water, it is possible to calculate the quantity of heat released during the combustion reaction.

## Sample Problem II-2

To study the amount of heat released during a neutralization reaction (you will learn about neutralization in Chapter 21), 25.0 mL of water containing 0.025 mol HCl is added to 25.0 mL of water containing 0.025 mol NaOH in a foam cup calorimeter. At the start, the solutions and the calorimeter are all at 25.0 °C. During the reaction, the highest temperature observed is 32.0 °C. Calculate the heat (in kJ) released during this reaction. Assume the densities of the solutions are 1.00 g/mL.

 1. **ANALYZE** List the knowns and the unknown.

Knowns:

HCl solution:

- $V_{\text{HCl}} = 25.0 \text{ mL}$
- solution contains 0.025 mol HCl

NaOH solution:

- $V_{\text{NaOH}} = 25.0 \text{ mL}$
- solution contains 0.025 mol NaOH
- $V_{\text{final}} = V_{\text{HCl}} + V_{\text{NaOH}}$   
 $= 25.0 \text{ mL} + 25.0 \text{ mL} = 50.0 \text{ mL}$
- $T_i = 25.0 \text{ }^\circ\text{C}$
- $T_f = 32.0 \text{ }^\circ\text{C}$
- $C_{\text{water}} = 4.18 \text{ J}/(\text{g} \times \text{ }^\circ\text{C})$
- $\text{Density}_{\text{solution}} = 1.00 \text{ g}/\text{mL}$

Unknown:

- $\Delta H = ? \text{ kJ}$

The equation requires the mass of the water used in the experiment, but the mass is not known. Use dimensional analysis to determine the mass of the water.  $\Delta T$  must also be calculated. Once  $m$ ,  $C$ , and  $\Delta T$  are known, use  $\Delta H = m \times C \times \Delta T$  to solve for  $\Delta H$  of the water.

 2. **CALCULATE** Solve for the unknown.

First, calculate the total mass of the water. Only the final volume of the solution ( $V_f$ ) is needed to make the calculation.

$$m = (50.0 \text{ mL}) \times \left( \frac{1.00 \text{ g}}{\text{mL}} \right) = 50.0 \text{ g}$$

Now calculate  $\Delta T$ .

$$\Delta T = T_f - T_i \quad \Delta T = 32.0 \text{ }^\circ\text{C} - 25.0 \text{ }^\circ\text{C} = 7.0 \text{ }^\circ\text{C}$$

Substitute the values for  $m$ ,  $C_{\text{water}}$ , and  $\Delta T$  into the equation and solve for the unknown ( $\Delta H$ ).

$$\begin{aligned} \Delta H &= m \times C \times \Delta T \\ &= (50.0 \text{ g})(4.18 \text{ J}/(\text{g} \times \text{ }^\circ\text{C}))(7.0 \text{ }^\circ\text{C}) \\ &= 1463 \text{ J} = 1.5 \times 10^3 \text{ J} \end{aligned}$$

Convert joules to kilojoules.

$$\Delta H = (1.5 \times 10^3 \text{ J}) \left( \frac{1 \text{ kJ}}{1000 \text{ J}} \right) = 1.5 \text{ kJ}$$

## Practice Problems

- A student mixed 50.0 mL of water containing 0.50 mol HCl at 22.5 °C with 50.0 mL of water containing 0.50 mol NaOH at 22.5 °C in a foam cup calorimeter. The temperature of the resulting solution increased to 26.0 °C. How much heat in kilojoules (kJ) was released by this reaction?
- A small pebble is heated and placed in a foam cup calorimeter containing 25.0 mL of water at 25.0 °C. The water reaches a maximum temperature of 26.4 °C. How many joules of heat were released by the pebble?

**Chem ASAP!**
**Problem-Solving 12**

Solve Problem 12 with the help of an interactive guided tutorial.



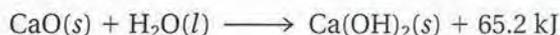
## Sample Problem 11-2 (cont.)

3. **EVALUATE** Does the result make sense?

The sign of  $\Delta H$  for the water is positive; the water absorbs 1.5 kJ of heat. Therefore, this neutralization reaction releases 1.5 kJ of heat into the water in the calorimeter, so the sign of  $\Delta H$  for the reaction is negative. About 4 J of heat is required to raise the temperature of 1 g of water 1 °C. Thus it would take about 200 J to raise the temperature of 50 g of water 1 °C. Further, about 1400 J, or 1.4 kJ, is needed to raise the temperature of 50 g of water 7 °C. This estimated answer is very close to the calculated value of  $\Delta H$  for the neutralization reaction.

## Thermochemical Equations

If you mix calcium oxide with water, an exothermic reaction takes place. The water in the mixture becomes warm. This reaction occurs when cement, which contains calcium oxide, is mixed to make concrete. When 1 mol of calcium oxide reacts with 1 mol of water, 1 mol of calcium hydroxide forms and 65.2 kJ of heat is released. You can show this in the chemical equation by including heat change as a product of the reaction. The diagram in Figure 11.11a shows the heat change that occurs in this exothermic reaction.



You can treat heat change in a chemical reaction like any other reactant or product in a chemical equation. An equation that includes the heat change is called a **thermochemical equation**. A **heat of reaction** is the heat change for the equation exactly as it is written. You will usually see heats of reaction reported as  $\Delta H$ , which is the heat change at constant pressure. The physical state of the reactants and products must also be given. The standard conditions are that the reaction is carried out at 101.3 kPa (1 atmosphere) and that the reactants and products are in their usual physical states at 25 °C.

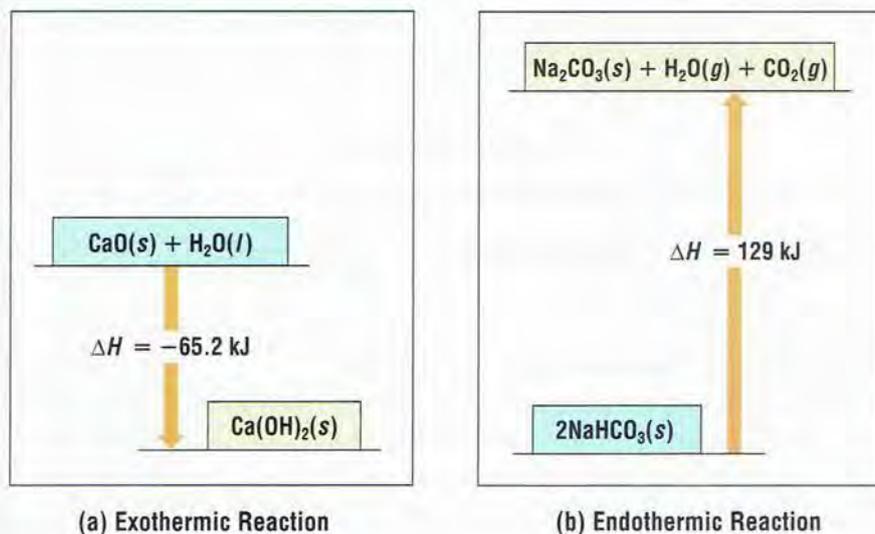


Figure 11.10

The reaction between iron(III) oxide and aluminum, called the thermite reaction, releases so much heat that the iron produced is in the molten state.

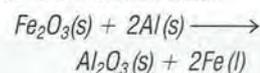
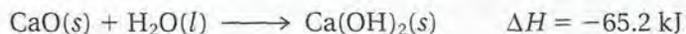


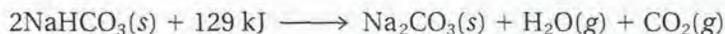
Figure 11.11

These enthalpy diagrams show exothermic and endothermic processes: (a) the reaction of calcium oxide and water and (b) the decomposition of sodium hydrogen carbonate. In which case is the enthalpy of the reactant(s) higher than that of the product(s)?

The heat of reaction, or  $\Delta H$ , in the above example is  $-65.2$  kJ. Each mole of calcium oxide and water that react to form calcium hydroxide produces  $65.2$  kJ of heat.



Other reactions absorb heat from the surroundings. For example, baking soda (sodium hydrogen carbonate) decomposes when it is heated, making it useful in baking. The carbon dioxide released in the reaction causes a cake to rise while baking. This process is endothermic, and the heat of reaction is  $129$  kJ.

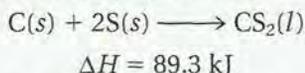


Remember that  $\Delta H$  is positive for endothermic reactions. Therefore, you can write the reaction as follows.

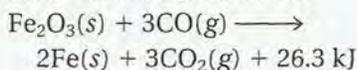


### Practice Problems

13. When carbon disulfide is formed from its elements, heat is absorbed. Calculate the amount of heat (in kJ) absorbed when  $5.66$  g of carbon disulfide is formed.



14. The production of iron and carbon dioxide from iron(III) oxide and carbon monoxide is an exothermic reaction. How many kilojoules of heat are produced when  $3.40$  mol  $\text{Fe}_2\text{O}_3$  reacts with an excess of  $\text{CO}$ ?



#### Chem ASAP!

#### Problem-Solving 14

Solve Problem 14 with the help of an interactive guided tutorial.



### Sample Problem 11-3

Using the equation for the reaction above, calculate the kilojoules of heat required to decompose  $2.24$  mol  $\text{NaHCO}_3(s)$ .

1. **ANALYZE** List the knowns and the unknown.

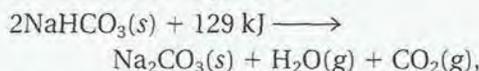
Knowns:

- $2.24$  mol  $\text{NaHCO}_3(s)$  decomposes
- $\Delta H = 129$  kJ

Unknown:

- $\Delta H = ?$  kJ

Use the thermochemical equation,



to write a conversion factor relating kilojoules of heat and moles of  $\text{NaHCO}_3$ . Then use the conversion factor to determine  $\Delta H$  for  $2.24$  mol  $\text{NaHCO}_3$ .

2. **CALCULATE** Solve for the unknown.

The thermochemical equation indicates that  $129$  kJ are needed to decompose  $2$  mol  $\text{NaHCO}_3(s)$ . Using this relationship, the conversion factor is

$$\frac{129 \text{ kJ}}{2 \text{ mol NaHCO}_3(s)}$$

Using dimensional analysis, solve for  $\Delta H$ .

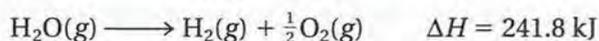
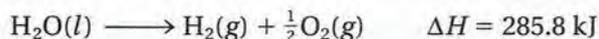
$$\Delta H = 2.24 \text{ mol NaHCO}_3(s) \times \frac{129 \text{ kJ}}{2 \text{ mol NaHCO}_3(s)} \\ = 144 \text{ kJ}$$

3. **EVALUATE** Does the result make sense?

Because the  $\Delta H$  of  $129$  kJ refers to the decomposition of  $2$  mol  $\text{NaHCO}_3(s)$ , the decomposition of  $2.24$  mol should absorb about  $10\%$  more heat than  $129$  kJ, or slightly more than  $142$  kJ. The answer of  $144$  kJ is consistent with this estimate.

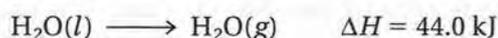
Chemistry problems involving enthalpy changes are similar to stoichiometry problems. The amount of heat released or absorbed during a reaction depends on the number of moles of the reactants involved. The decomposition of 2 mol of sodium hydrogen carbonate, for example, requires 129 kJ of heat. Therefore, the decomposition of 4 mol of the same substance would require twice as much heat, or 258 kJ. **Figure 11.11b** on page 303 shows the heat changes for this reaction. In this and other endothermic processes, the potential energy of the product(s) is higher than the potential energy of the reactant(s).

The physical state of the reactants and products in a thermochemical reaction must also be stated. To see why, compare the following two equations for the decomposition of 1 mol of water.



$$\text{difference} = 44.0 \text{ kJ}$$

Although the two equations are very similar, the different physical states of the  $\text{H}_2\text{O}$  result in different  $\Delta H$  values. In one case, the reactant is a liquid; in the other case, the reactant is a gas. The vaporization of 1 mole of liquid water to water vapor at 25 °C requires an extra 44.0 kJ of heat. Notice also that fractional coefficients are used here for  $\text{O}_2$  because 1 mol  $\text{H}_2\text{O}$  is being decomposed.



**Table 11.4** lists heats of combustion for some common substances. The **heat of combustion** is the heat of reaction for the complete burning of one mole of a substance.

**Table 11.4**

Heats of Combustion at 25 °C		
Substance	Formula	$\Delta H$ (kJ/mol)
Hydrogen	$\text{H}_2(g)$	-286
Carbon	$\text{C}(s)$ , graphite	-394
Carbon monoxide	$\text{CO}(g)$	-283
Methane	$\text{CH}_4(g)$	-890
Methanol	$\text{CH}_3\text{OH}(l)$	-726
Acetylene	$\text{C}_2\text{H}_2(g)$	-1300
Ethanol	$\text{C}_2\text{H}_5\text{OH}(l)$	-1368
Propane	$\text{C}_3\text{H}_8(g)$	-2220
Benzene	$\text{C}_6\text{H}_6(l)$	-3268
Glucose	$\text{C}_6\text{H}_{12}\text{O}_6(s)$	-2808
Octane	$\text{C}_8\text{H}_{18}(l)$	-5471
Sucrose	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(s)$	-5645

**LINK**  
TO  
**BIOLOGY**

### Warmth from Fat

How do animals such as polar bears and seals survive the cold land and water temperatures where they live? A good coat of fur helps, but it is not enough to



keep them warm. These animals also have special fat cells that help generate heat. These special cells are in tissue called brown fat. The cells of brown fat are unlike other fat cells in the animal's body. Most other cells store chemical energy from the breakdown of carbohydrates and fatty acids in adenosine triphosphate (ATP). ATP acts as the central source of energy for the activities and growth of all animals. Heat is mostly a waste product in ATP-producing cells. The heat generated by the brown fat tissue, however, helps the animal keep relatively comfortable even at subzero temperatures.



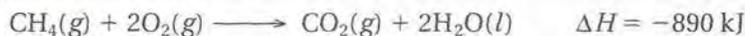
**Figure 11.12**

The combustion of natural gas is an exothermic reaction. As bonds in methane, the main component of natural gas, and oxygen are broken and bonds in carbon dioxide and water are formed, large amounts of energy are released. This energy powers the engine of automobiles. Natural gas is an alternative to gasoline because it substantially reduces air pollution while increasing engine life.

The combustion of natural gas, which is mostly methane, is an exothermic reaction used to heat many homes around the country.



This can also be written as follows.



Burning 1 mol of methane releases 890 kJ of heat. The heat of combustion ( $\Delta H$ ) for this reaction is  $-890 \text{ kJ}$  per mole of carbon burned.

Like other heats of reaction, heats of combustion are reported as the enthalpy changes when the reactions are carried out at 101.3 kPa of pressure and the reactants and products are in their physical states at 25 °C.

### section review 11.2

- When 2 mol of solid magnesium (Mg) combines with 1 mole of oxygen gas ( $\text{O}_2$ ), 2 mol of solid magnesium oxide (MgO) is formed and 1204 kJ of heat is released. Write the thermochemical equation for this combustion reaction.
- Gasohol contains ethanol ( $\text{C}_2\text{H}_5\text{OH}(l)$ ), which when burned reacts with oxygen to produce  $\text{CO}_2(g)$  and  $\text{H}_2\text{O}(g)$ . How much heat is released when 12.5 g of ethanol burns?
- Explain the term heat of reaction.
- Hydrogen gas and fluorine gas react to produce hydrogen fluoride. Calculate the heat change (in kJ) for the conversion of 15.0 g of hydrogen gas to hydrogen fluoride gas at constant pressure.



- Why is it important to give the physical state of a substance in a thermochemical reaction?



**Chem ASAP! Assessment 11.2** Check your understanding of the important ideas and concepts in Section 11.2.

## HEAT IN CHANGES OF STATE



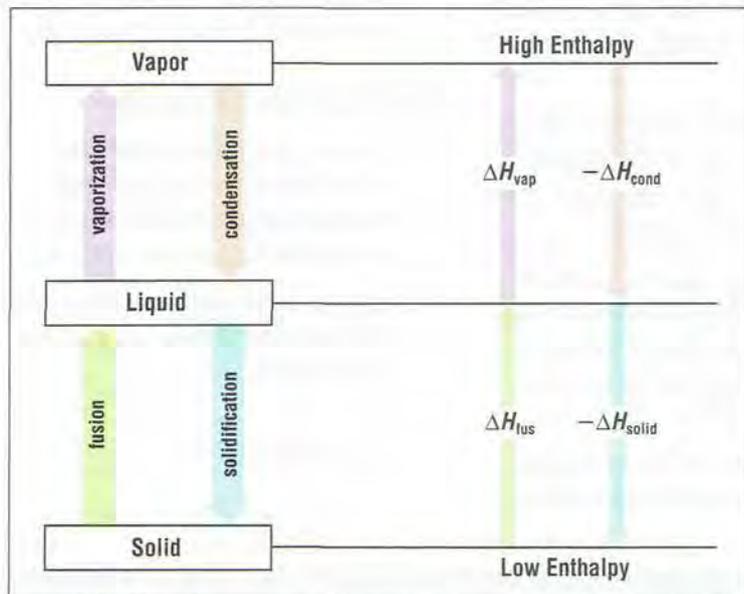
When your body heats up, you start to sweat. The evaporation of sweat is your body's way of cooling itself to a normal temperature. Why does the evaporation of sweat from your skin help to rid your body of excess heat?

## Heats of Fusion and Solidification

What happens if you place an ice cube on a table in a warm room? The ice cube is the system, and the table and air around it are the surroundings. The ice absorbs heat from its surroundings and begins to melt. The temperature of the ice and the water produced remains at 0 °C until all of the ice has melted. The temperature of the water begins to increase only after all of the ice has melted. In this section you will learn about heat changes that occur during changes of state.

Like ice cubes, all solids absorb heat as they melt to become liquids. The heat absorbed by one mole of a substance in melting from a solid to a liquid at a constant temperature is the **molar heat of fusion** ( $\Delta H_{\text{fus}}$ ). The heat lost when one mole of a liquid solidifies at a constant temperature is the **molar heat of solidification** ( $\Delta H_{\text{solid}}$ ). The quantity of heat absorbed by a melting solid is exactly the same as the quantity of heat lost when the liquid solidifies; that is,  $\Delta H_{\text{fus}} = -\Delta H_{\text{solid}}$ , as shown in Figure 11.13. Why is this true? Table 11.5 on the following page gives heats of fusion of some substances.

The melting of 1 mol of ice at 0 °C to 1 mol of water at 0 °C requires the absorption of 6.01 kJ of heat. This quantity of heat is the molar heat of fusion. Likewise, the conversion of 1 mol of water at 0 °C to 1 mol of ice at 0 °C releases 6.01 kJ. This quantity of heat is the molar heat of solidification.



## objectives

- ▶ Classify, by type, the heat changes that occur during melting, freezing, boiling, and condensing
- ▶ Calculate heat changes that occur during melting, freezing, boiling, and condensing

## key terms

- ▶ molar heat of fusion
- ▶ molar heat of solidification
- ▶ molar heat of vaporization
- ▶ molar heat of condensation
- ▶ molar heat of solution

## Chem ASAP!

## Animation 12

Observe the phase changes as ice is converted to steam when heat is added.



Figure 11.13

Enthalpy changes accompany changes in state. Fusion and vaporization are endothermic processes. Solidification and condensation are exothermic processes.

Table 11.5

Heats of Physical Change					
Substance	Formula	Freezing point (K)	$\Delta H_{\text{fus}}$ (kJ/mol)	Boiling point (K)	$\Delta H_{\text{vap}}$ (kJ/mol)
Acetone	CH <sub>3</sub> COCH <sub>3</sub>	177.8	5.72	329.4	29.1
Ammonia	NH <sub>3</sub>	195.3	5.65	239.7	23.4
Argon	Ar	83.8	1.2	87.3	6.5
Benzene	C <sub>6</sub> H <sub>6</sub>	278.7	9.87	353.3	30.8
Ethanol	C <sub>2</sub> H <sub>5</sub> OH	158.7	4.60	351.5	43.5
Helium	He	3.5	0.02	4.22	0.08
Hydrogen	H <sub>2</sub>	14.0	0.12	20.3	0.90
Methane	CH <sub>4</sub>	90.7	0.94	111.7	8.2
Methanol	CH <sub>3</sub> OH	175.5	3.16	337.2	35.3
Neon	Ne	24.5	0.33	27.1	1.76
Nitrogen	N <sub>2</sub>	63.3	0.72	77.4	5.58
Oxygen	O <sub>2</sub>	54.8	0.44	90.2	6.82
Water	H <sub>2</sub> O	273.2	6.01	373.2	40.7

## MINI LAB



### Heat of Fusion of Ice

#### PURPOSE

To estimate the heat of fusion of ice.

#### MATERIALS

- ice
- foam cup
- graduated cylinder
- thermometer
- hot water
- temperature probe (optional)

#### PROCEDURE



Probe version available in the *Probeware Lab Manual*.

1. Fill a 100-mL graduated cylinder with hot tap water. Allow the filled cylinder to stand for 1 minute. Pour the water into the sink.
2. Use the graduated cylinder to measure 70 mL of hot water. Pour the water into the foam cup. Measure the temperature of the water.
3. Add a small ice cube to the cup of water and gently swirl the cup. Measure the temperature of the water immediately after the ice cube has completely melted.
4. Pour the water into the graduated cylinder and measure the volume.

5. Calculate the heat of fusion of ice (kJ/mol) by dividing the heat given up from the water by the moles of ice melted. *Hint:* The mass of ice melted is the same as the increase in the volume of the water: 1 g H<sub>2</sub>O = 1 mL H<sub>2</sub>O.

#### ANALYSIS AND CONCLUSIONS

1. Compare your experimental value for the heat of fusion of ice with the accepted value of 6.01 kJ/mol. Account for any error in your value.
2. Suggest some changes in this procedure that would improve the accuracy of the results.



Figure 11.14

Ice is commonly used to refrigerate perishable foods. What happens to the temperature of the ice as it begins to melt?

### Sample Problem 11-4

How many grams of ice at 0 °C and 101.3 kPa could be melted by the addition of 2.25 kJ of heat?

1. **ANALYZE** List the knowns and the unknown.

Knowns:

- Initial conditions are 0 °C and 101.3 kPa.
- $\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$
- $\Delta H = 2.25 \text{ kJ}$

Unknown:

- $m_{\text{ice}} = ? \text{ g}$

The conditions 0 °C and 101.3 kPa indicate that the standard conditions for the fusion of ice have been met. Use the chemical equation  $\text{H}_2\text{O}(s) + 6.01 \text{ kJ} \rightarrow \text{H}_2\text{O}(l)$  to find the number of moles of ice that can be melted by the addition of 2.25 kJ of heat. Convert moles of ice to grams of ice.

2. **CALCULATE** Solve for the unknown.

The required conversion factors come from  $\Delta H_{\text{fus}}$  and the molar mass of ice. The conversion factors are

$$\frac{1 \text{ mol ice}}{6.01 \text{ kJ}} \quad \text{and} \quad \frac{18.0 \text{ g ice}}{1 \text{ mol ice}}$$

Multiply the known heat change (2.25 kJ) by the conversion factors

$$\begin{aligned} m_{\text{ice}} &= 2.25 \text{ kJ} \times \frac{1 \text{ mol ice}}{6.01 \text{ kJ}} \times \frac{18.0 \text{ g ice}}{1 \text{ mol ice}} \\ &= 6.74 \text{ g ice} \end{aligned}$$

3. **EVALUATE** Does the result make sense?

6.01 kJ is required to melt 1 mol of ice. Because only about one-third of this amount of heat (roughly 2 kJ) is available, only about one-third mol of ice, or  $\frac{18.0 \text{ g}}{3} = 6 \text{ g}$ , should melt.

This estimate and the calculated answer are similar.

### Practice Problems

20. How many grams of ice at 0 °C and 101.3 kPa could be melted by the addition of 0.400 kJ of heat?
21. How many kilojoules of heat are required to melt a 10.0 g popsicle at 0 °C and 101.3 kPa? Assume the popsicle has the same molar mass and heat capacity as water.

#### Chem ASAP!

#### Problem-Solving 21

Solve Problem 21 with the help of an interactive guided tutorial.



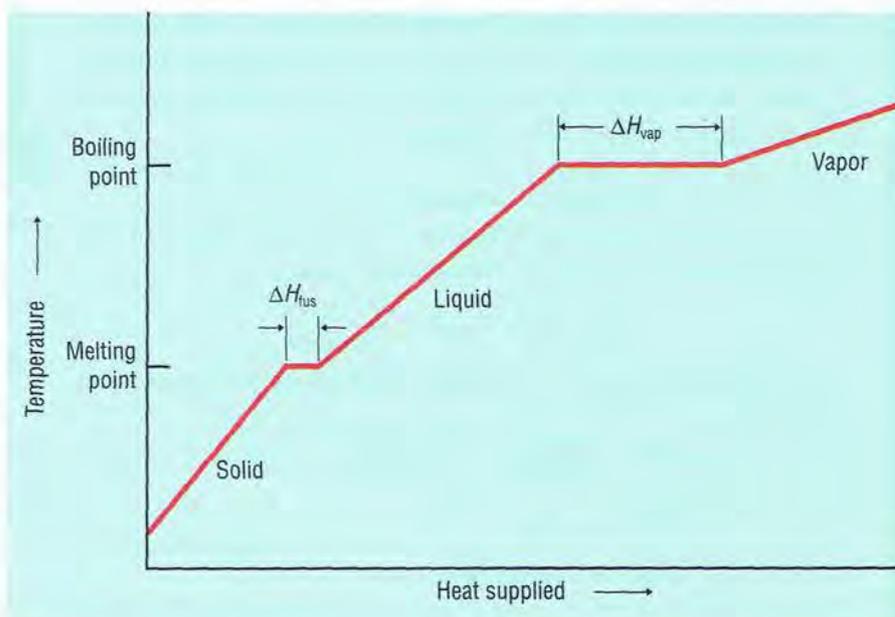
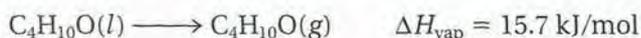
## Heats of Vaporization and Condensation

When liquids absorb heat at their boiling points, they become vapors. Vaporization of a liquid, through boiling or evaporation, cools the environment around the liquid as heat flows from the surroundings to the liquid. The amount of heat necessary to vaporize one mole of a given liquid is called its **molar heat of vaporization**. Table 11.5 on page 308 gives some values of molar heats of vaporization for various compounds. The values are derived at standard conditions: the reactions are carried out at one atmosphere pressure and the reactants and products are in their usual physical states at the same temperature.

The molar heat of vaporization of water is 40.7 kJ/mol. This means that in order to vaporize 1 mol of water, 40.7 kJ of energy must be supplied. This energy converts 1 mol of water molecules in the liquid state to 1 mol of water molecules in the vapor state, given the same temperature and 1 atm pressure. This process is described in the thermochemical equation below.



Diethyl ether ( $\text{C}_4\text{H}_{10}\text{O}$ ) is a low-boiling-point liquid (bp = 34.6 °C) that is a good solvent and was formerly used as an anesthetic. If diethyl ether is poured into a beaker on a warm, humid day, the ether will absorb heat from the beaker walls and evaporate very rapidly. If the beaker loses enough heat, the water vapor in the air may condense and freeze on the beaker walls. If so, a coating of frost will form on the outside of the beaker. Diethyl ether has a molar heat of vaporization ( $\Delta H_{\text{vap}}$ ) of 15.7 kJ/mol. Is this an endothermic or an exothermic process?



**Figure 11.15**

This graph shows the heating curve for water. Notice that the temperature remains constant during melting and vaporization. Notice also that it requires much more energy to vaporize liquid water than it does to melt the same mass of ice.

Condensation is the exact opposite of vaporization. Therefore, the amount of heat released when 1 mol of vapor condenses is called its **molar heat of condensation** ( $\Delta H_{\text{cond}}$ ). This value is numerically the same as the corresponding molar heat of vaporization, however the value has the opposite sign. Because energy is conserved in a physical change,  $\Delta H_{\text{vap}} = -\Delta H_{\text{cond}}$ . Figure 11.15 summarizes the heat changes that occur as a solid is heated to a liquid and then to a gas. You should be able to identify certain trends regarding the temperature during changes of state and the energy requirements that accompany these changes from the graph. The large values for  $\Delta H_{\text{vap}}$  and  $\Delta H_{\text{cond}}$  are the reason hot vapors such as steam can be very dangerous. You can receive a scalding burn from steam when the heat of condensation is released as it touches your skin.



### Sample Problem 11-5

How much heat (in kJ) is absorbed when 24.8 g  $\text{H}_2\text{O}(l)$  at 100 °C is converted to steam at 100 °C?

**1. ANALYZE** List the knowns and the unknown.

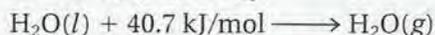
*Knowns:*

- mass of water converted to steam = 24.8 g
- $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$

*Unknown:*

- $\Delta H = ? \text{ kJ}$

The  $\Delta H_{\text{vap}}$  in the following equation is given in kJ/mol, but the quantity of water is given in grams. Thus the first steps in the problem solution are to convert grams of water to moles of water and then multiply by  $\Delta H_{\text{vap}}$ .



**2. CALCULATE** Solve for the unknown.

The required conversion factors come from  $\Delta H_{\text{vap}}$  and the molar mass of water.

$$\frac{1 \text{ mol H}_2\text{O}(l)}{18.0 \text{ g H}_2\text{O}(l)} \quad \text{and} \quad \frac{40.7 \text{ kJ}}{1 \text{ mol H}_2\text{O}(l)}$$

Multiply the known mass of water in grams by the conversion factors.

$$\begin{aligned} \Delta H &= 24.8 \text{ g H}_2\text{O}(l) \times \frac{1 \text{ mol H}_2\text{O}(l)}{18.0 \text{ g H}_2\text{O}(l)} \times \frac{40.7 \text{ kJ}}{1 \text{ mol H}_2\text{O}(l)} \\ &= 56.1 \text{ kJ} \end{aligned}$$

**3. EVALUATE** Does the result make sense?

Knowing the molar mass of water is 18.0 g/mol, 24.8 g  $\text{H}_2\text{O}(l)$  can be estimated to be somewhat less than 1.5 mol  $\text{H}_2\text{O}$ . Thus the calculated heat change should be somewhat less than  $1.5 \text{ mol} \times 40 \text{ kJ/mol} = 60 \text{ kJ}$ , and it is.

### Practice Problems

- 22.** How much heat (in kJ) is absorbed when 63.7 g  $\text{H}_2\text{O}(l)$  at 100 °C is converted to steam at 100 °C?
- 23.** How many kilojoules of heat are absorbed when 0.46 g of chloroethane ( $\text{C}_2\text{H}_5\text{Cl}$ , bp 12.3 °C) vaporizes at its boiling point? The molar heat of vaporization of chloroethane is 26.4 kJ/mol.

**Chem ASAP!**

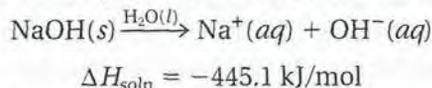
**Problem-Solving 23**

Solve Problem 23 with the help of an interactive guided tutorial.

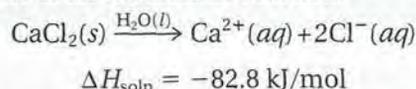


## Heat of Solution

Heat changes can also occur when a solute dissolves in a solvent. The heat change caused by dissolution of one mole of substance is the **molar heat of solution** ( $\Delta H_{\text{soln}}$ ). Sodium hydroxide provides a good example of an exothermic molar heat of solution. When 1 mol of sodium hydroxide ( $\text{NaOH}(s)$ ) is dissolved in water, the solution can become so hot that it steams. The heat from this process is released as the sodium ions and the hydroxide ions separate and interact with the water. The temperature of the solution increases, releasing 445.1 kJ of heat as the molar heat of solution.



A practical application of an exothermic reaction is how a hot pack works. A hot pack mixes calcium chloride ( $\text{CaCl}_2$ ) and water, which produces the heat characteristic of an exothermic reaction.



The dissolution of ammonium nitrate ( $\text{NH}_4\text{NO}_3$ )(s) is an example of an endothermic process. When ammonium nitrate dissolves in water, the solution becomes so cold that frost may form on the outside of the container. Is heat absorbed or released as the ammonium and nitrate ions separate and interact with the water? Heat is released from the water and the temperature of the solution decreases. The cold pack allows water and ammonium nitrate ( $\text{NH}_4\text{NO}_3$ ) to mix, producing an endothermic reaction.

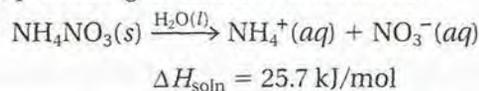
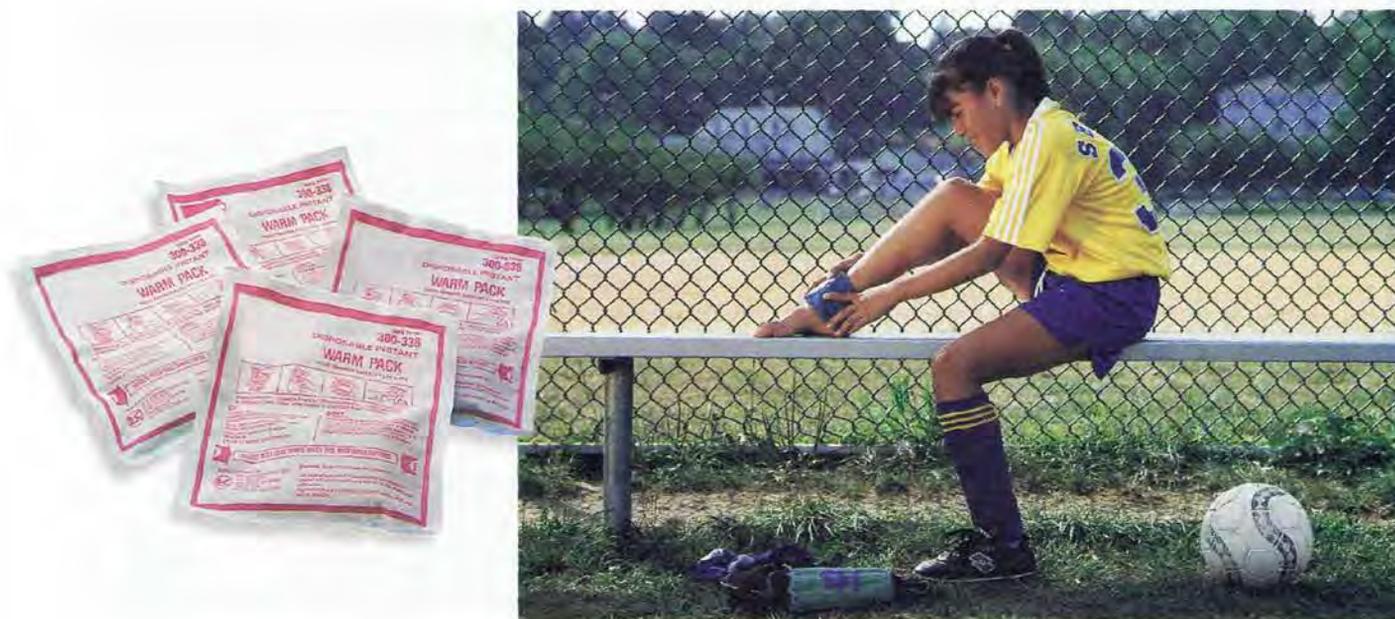


Figure 11.16 illustrates a practical application of heats of solution.



**Figure 11.16**

*Cold packs and hot packs are available for a variety of medical uses.*

## Sample Problem 11-6

How much heat (in kJ) is released when 2.500 mol NaOH(s) is dissolved in water?

1. **ANALYZE** List the knowns and the unknown.

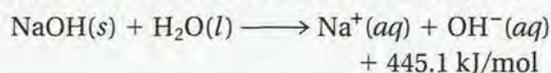
Knowns:

- $\Delta H_{\text{soln}} = -445.1 \text{ kJ/mol}$
- amount of NaOH(s) dissolved: 2.500 mol

Unknown:

- $\Delta H = ? \text{ kJ}$

Use the heat of solution from the following chemical equation to solve for the amount of heat released ( $\Delta H$ ).



2. **CALCULATE** Solve for the unknown.

Multiplying the number of mol NaOH and  $\Delta H_{\text{soln}}$  will yield the value of the unknown ( $\Delta H$ ).

$$\Delta H = 2.500 \text{ mol NaOH}(s) \times \frac{-445.1 \text{ kJ}}{1 \text{ mol NaOH}(aq)} = -1113 \text{ kJ}$$

3. **EVALUATE** Does the result make sense?

By inspection,  $\Delta H$  is 2.5 times greater than  $\Delta H_{\text{soln}}$ , as it should be. Also, the heat of solution in the answer is negative, indicating an exothermic reaction. This is consistent with the heat of solution of sodium hydroxide discussed earlier in the text.

## Practice Problems

24. How much heat (in kJ) is released when 0.677 mol NaOH(s) is dissolved in water?
25. How many moles of  $\text{NH}_4\text{NO}_3(s)$  must be dissolved in water so that 88.0 kJ of heat is released from the water?

**Chem ASAP!**

**Problem-Solving 25**

Solve Problem 25 with the help of an interactive guided tutorial.



## section review 11.3

26. Identify each heat change by name and classify each change as exothermic or endothermic.
- 1 mol  $\text{C}_3\text{H}_8(l) \longrightarrow 1 \text{ mol C}_3\text{H}_8(g)$
  - 1 mol  $\text{NaCl}(s) + 3.88 \text{ kJ/mol} \longrightarrow 1 \text{ mol NaCl}(aq)$
  - 1 mol  $\text{NaCl}(s) \longrightarrow 1 \text{ mol NaCl}(l)$
  - 1 mol  $\text{NH}_3(g) \longrightarrow 1 \text{ mol NH}_3(l)$
  - 1 mol  $\text{Hg}(l) \longrightarrow 1 \text{ mol Hg}(s)$
27. Heavy water, in which the hydrogens are hydrogen-2 instead of the more common hydrogen-1, is called deuterium oxide ( $\text{D}_2\text{O}$ ). Solid  $\text{D}_2\text{O}$  melts at  $3.78^\circ\text{C}$ . The molar heat of fusion of  $\text{D}_2\text{O}(s)$  is  $6.34 \text{ kJ/mol}$ . How much heat is released when  $8.46 \text{ g D}_2\text{O}(l)$  solidifies at its melting point?
28. Why is a burn from steam potentially far more serious than a burn from very hot water?
29. Why does an ice cube melt at room temperature?



**Chem ASAP! Assessment 11.3** Check your understanding of the important ideas and concepts in Section 11.3.

**objectives**

- ▶ Apply Hess's law of heat summation to find heat changes for chemical and physical processes
- ▶ Calculate heat changes using standard heats of formation

**key terms**

- ▶ Hess's law of heat summation
- ▶ standard heat of formation

**E**meralds are beautiful gemstones composed of the elements chromium, aluminum, silicon, oxygen, and beryllium. If you were interested in the heat changes that occur when an emerald is converted to its elements, it is more than likely that you would not want to destroy the emerald by measuring the heat changes directly. **Is there a way to determine the heat of reaction without actually performing the reaction?**

**Hess's Law**

It is possible to talk in general terms about the heat changes that take place in chemical reactions. However, most reactions occur in a series of steps. Suppose, for example, you need to know the heat of reaction for an intermediate step, but it is impossible for you to obtain the value directly. Fortunately, Hess's law makes it possible to measure a heat of reaction indirectly. Thus even when a direct measurement cannot be made, the heat of reaction can still be determined.

Elemental carbon exists as both graphite and diamond at 25 °C. Because graphite is more stable than diamond, you would expect the following process to take place.

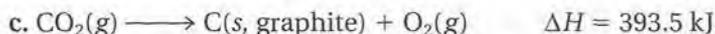


Fortunately for people who own diamonds, the conversion of diamond to graphite takes millions and millions of years. This enthalpy change cannot be measured directly because the reaction is far too slow. Hess's law, however, provides a way to calculate the heat of reaction. Hess's law is expressed by a simple rule: If you add two or more thermochemical equations to give a final equation, then you can also add the heats of reaction to give the final heat of reaction. This rule is **Hess's law of heat summation**.

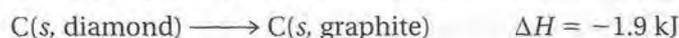
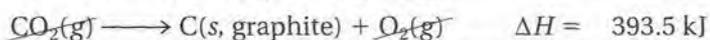
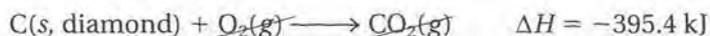
You can use Hess's law to find the enthalpy changes for the conversion of diamond to graphite by using the following combustion reactions and Figure 11.17.



Write equation **a** in reverse to give:



When you write a reverse reaction, you must also change the sign of  $\Delta H$ . If you now add equations **b** and **c**, you get the equation for the conversion of diamond to graphite. The  $\text{CO}_2(\text{g})$  and  $\text{O}_2(\text{g})$  terms on both sides of the summed equations cancel, just as they do in algebra. Now if you also add the values of  $\Delta H$  for equations **b** and **c**, you get the heat of reaction for this conversion.



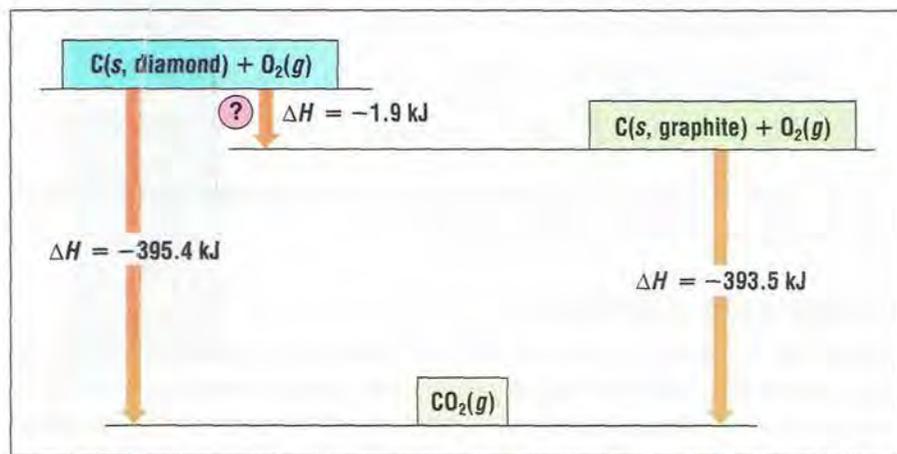
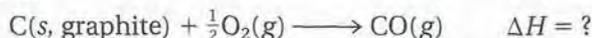


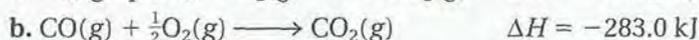
Figure 11.17

Hess's law is used to determine the enthalpy change of a very slow chemical process.  
 $C(s, \text{diamond}) \longrightarrow C(s, \text{graphite})$

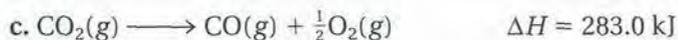
Thus the conversion of diamond to graphite is an exothermic process; its heat of reaction has a negative sign. Conversely, the change of graphite to diamond is an endothermic process. Some reactions give other products in addition to the product of interest. Suppose you want to determine the enthalpy change for the formation of carbon monoxide from its elements. The reaction is:



Although it is easy to write the equation, carrying out the reaction in the laboratory as written is virtually impossible. Carbon dioxide (a "side product") is produced along with carbon monoxide (the "desired product"). Therefore, any measured heat of reaction is related to the formation of both  $CO(g)$  and  $CO_2(g)$ , and not  $CO(g)$  alone. You can solve the problem, however, using Hess's law and two reactions that can be carried out in the laboratory. The reactions are:



Writing the reverse of equation **b** and changing the sign of  $\Delta H$  yields equation **c**.



As shown on the following page, adding equations **a** and **c** gives the expression for the formation of  $CO(g)$  from its elements. Notice that only  $\frac{1}{2}O_2(g)$  cancels in the final equation. See Figure 11.18.

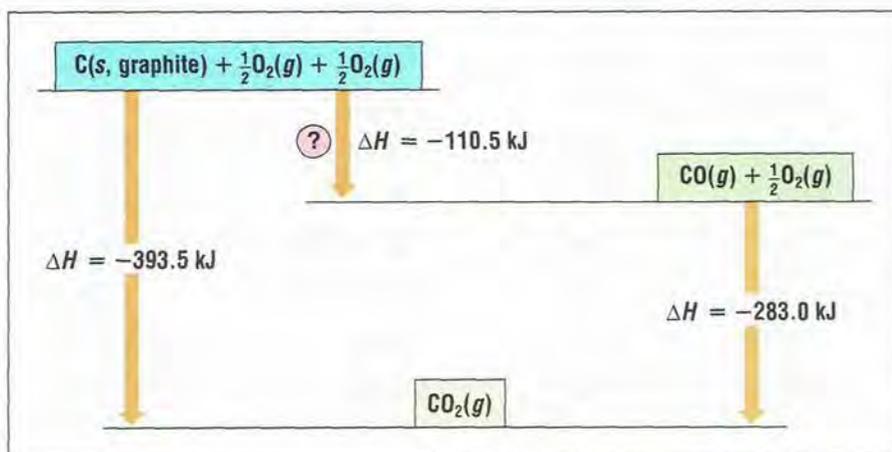
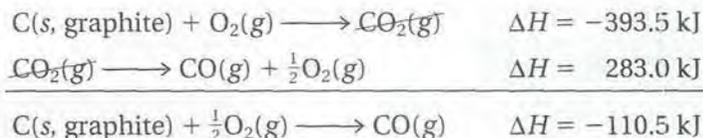


Figure 11.18

Hess's law is used to determine the enthalpy change for the reaction:  
 $C(s, \text{graphite}) + \frac{1}{2}O_2(g) \longrightarrow CO(g)$



The formation of CO(g) is exothermic; 110.5 kJ of heat is given off when 1 mol CO(g) is formed from its elements.

## Standard Heats of Formation

Sometimes it is hard to measure the heat change for a reaction. What are some examples where this might be the case? In such cases you can calculate the heat of reaction from standard heats of formation. The **standard heat of formation** ( $\Delta H_f^\circ$ ) of a compound is the change in enthalpy that accompanies the formation of one mole of a compound from its elements with all substances in their standard states at 25 °C.

The  $\Delta H_f^\circ$  of a free element in its standard state is arbitrarily set at zero. For example,  $\Delta H_f^\circ = 0$  for the diatomic molecules H<sub>2</sub>(g), N<sub>2</sub>(g), O<sub>2</sub>(g), F<sub>2</sub>(g), Cl<sub>2</sub>(g), Br<sub>2</sub>(l), and I<sub>2</sub>(s). Similarly,  $\Delta H_f^\circ = 0$  for the graphite form of carbon, C(s, graphite).

Many values of  $\Delta H_f^\circ$  have been measured. Table 11.6 lists  $\Delta H_f^\circ$  values for some common substances. Standard heats of formation of compounds are handy for calculating heats of reaction at standard conditions. The standard heat of reaction ( $\Delta H^\circ$ ) is the difference between the standard heats of formation of all the reactants and products. This relationship can be expressed by the following equation.

$$\Delta H^\circ = \Delta H_f^\circ(\text{products}) - \Delta H_f^\circ(\text{reactants})$$

Table 11.6

Standard Heats of Formation ( $\Delta H_f^\circ$ ) at 25 °C and 101.3 kPa					
Substance	$\Delta H_f^\circ$ (kJ/mol)	Substance	$\Delta H_f^\circ$ (kJ/mol)	Substance	$\Delta H_f^\circ$ (kJ/mol)
Al <sub>2</sub> O <sub>3</sub> (s)	-1676.0	Fe(s)	0.0	NO(g)	90.37
Br <sub>2</sub> (g)	30.91	Fe <sub>2</sub> O <sub>3</sub> (s)	-822.1	NO <sub>2</sub> (g)	33.85
Br <sub>2</sub> (l)	0.0	H <sub>2</sub> (g)	0.0	Na <sub>2</sub> CO <sub>3</sub> (s)	-1131.1
C(s, diamond)	1.9	H <sub>2</sub> O(g)	-241.8	NaCl(s)	-411.2
C(s, graphite)	0.0	H <sub>2</sub> O(l)	-285.8	O <sub>2</sub> (g)	0.0
CH <sub>4</sub> (g)	-74.86	H <sub>2</sub> O <sub>2</sub> (l)	-187.8	O <sub>3</sub> (g)	142.0
CO(g)	-110.5	HCl(g)	-92.31	P(s, white)	0.0
CO <sub>2</sub> (g)	-393.5	H <sub>2</sub> S(g)	-20.1	P(s, red)	-18.4
CaCO <sub>3</sub> (s)	-1207.0	I <sub>2</sub> (g)	62.4	S(s, rhombic)	0.0
CaO(s)	-635.1	I <sub>2</sub> (s)	0.0	S(s, monoclinic)	0.30
Cl <sub>2</sub> (g)	0.0	N <sub>2</sub> (g)	0.0	SO <sub>2</sub> (g)	-296.8
F <sub>2</sub> (g)	0.0	NH <sub>3</sub> (g)	-46.19	SO <sub>3</sub> (g)	-395.7

Figure 11.19 is a diagram similar to others you have seen, except that it displays the standard heats of formation of the reactants hydrogen and oxygen and the product water. The heat difference between the reactants and products,  $-285.8 \text{ kJ/mol}$ , is the standard heat of formation of liquid water from the gases hydrogen and oxygen. Does water have a lower or higher enthalpy than the elements from which it is formed? On what other basis can you account for your answer?

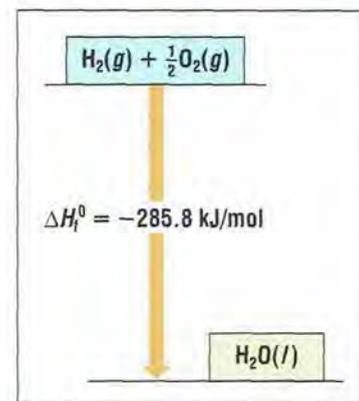


Figure 11.19

This enthalpy diagram shows the standard heat of formation of water.

### Sample Problem 11-7

What is the standard heat of reaction ( $\Delta H^0$ ) for the reaction of gaseous carbon monoxide with oxygen to form gaseous carbon dioxide?

1. **ANALYZE** List the knowns and the unknown.

Knowns (from Table 11.6):

- $\Delta H_f^0 \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$  (free element)
- $\Delta H_f^0 \text{CO}(\text{g}) = -110.5 \text{ kJ/mol}$
- $\Delta H_f^0 \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$

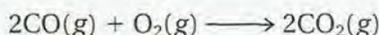
Unknown:

- $\Delta H^0 = ? \text{ kJ}$

Balance the equation of the reaction of  $\text{CO}(\text{g})$  with  $\text{O}_2(\text{g})$  to form  $\text{CO}_2(\text{g})$ . Then determine  $\Delta H^0$  using the standard heats of formation of the reactants and products.

2. **CALCULATE** Solve for the unknown.

First, write the balanced equation.



Next, find and sum the  $\Delta H_f^0$  of all of the reactants, taking into account the number of moles of each.

$$\begin{aligned} \Delta H_f^0(\text{reactants}) &= 2 \frac{\text{mol CO}(\text{g})}{1 \text{ mol CO}(\text{g})} \times \frac{-110.5 \text{ kJ}}{1 \text{ mol CO}(\text{g})} \\ &\quad + 1 \frac{\text{mol O}_2(\text{g})}{1 \text{ mol O}_2(\text{g})} \times \frac{0 \text{ kJ}}{1 \text{ mol O}_2(\text{g})} \\ &= -221.0 \text{ kJ} \end{aligned}$$

Then, find the  $\Delta H_f^0$  of the product in a similar way.

$$\begin{aligned} \Delta H_f^0(\text{product}) &= 2 \frac{\text{mol CO}_2(\text{g})}{1 \text{ mol CO}_2(\text{g})} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol CO}_2(\text{g})} \\ &= -787.0 \text{ kJ} \end{aligned}$$

Finally, find the difference between  $\Delta H_f^0(\text{products})$  and  $\Delta H_f^0(\text{reactants})$ .

$$\begin{aligned} \Delta H^0 &= \Delta H_f^0(\text{products}) - \Delta H_f^0(\text{reactants}) \\ \Delta H^0 &= (-787.0 \text{ kJ}) - (-221.0 \text{ kJ}) \\ \Delta H^0 &= -566.0 \text{ kJ} \end{aligned}$$

### Practice Problems

30. Use the standard heats of formation to calculate the standard heats of reaction ( $\Delta H^0$ ) for these reactions.
- a.  $\text{Br}_2(\text{g}) \longrightarrow \text{Br}_2(\text{l})$
  - b.  $\text{CaCO}_3(\text{s}) \longrightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
  - c.  $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{NO}_2(\text{g})$
31. With one exception, the standard heats of formation of  $\text{Na}(\text{s})$ ,  $\text{O}_2(\text{g})$ ,  $\text{Br}_2(\text{l})$ ,  $\text{CO}(\text{g})$ ,  $\text{Fe}(\text{s})$ , and  $\text{He}(\text{g})$  are identical. What is the exception? Explain.

Chem ASAP!

#### Problem-Solving 30

Solve Problem 30 with the help of an interactive guided tutorial.



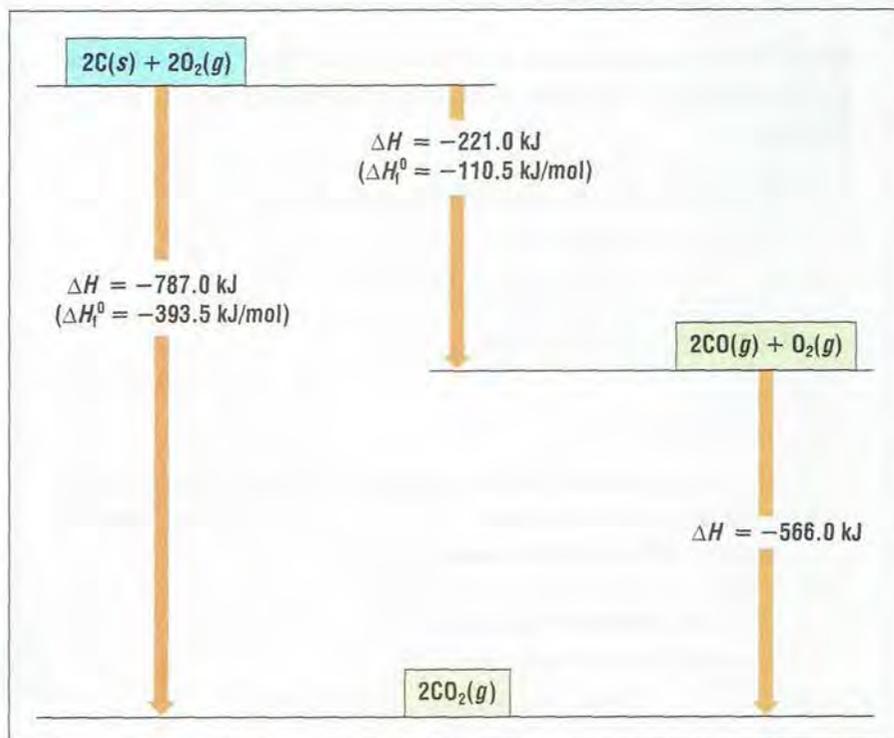
## Sample Problem 11-7 (cont.)

 3. **EVALUATE** Does the result make sense?

The  $\Delta H^0$  is negative. Therefore, the reaction is exothermic, as illustrated in Figure 11.20. This makes sense because the oxidation of carbon monoxide is a combustion reaction. Combustion reactions always release heat.

Figure 11.20

Hess's law is used to determine the enthalpy change for the reaction of carbon monoxide and oxygen.

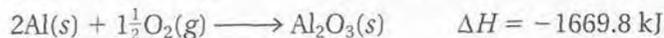
$$2\text{CO}(g) + \text{O}_2(g) \longrightarrow 2\text{CO}_2(g)$$


## section review 11.4

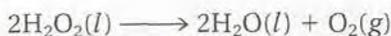
32. Calculate the enthalpy change ( $\Delta H$ ) in kJ for the following reaction.



Use the enthalpy changes for the combustion of aluminum and iron:



33. What is the standard heat of reaction ( $\Delta H^0$ ) for the decomposition of hydrogen peroxide?



34. State Hess's law of heat summation in your own words. Explain its usefulness.
35. What happens to the sign of  $\Delta H$  when the reverse of a chemical reaction is written? Why?



## portfolio project

Some chefs use heavy cast iron skillets; others use lightweight stainless steel. Use what you have learned about heat capacity to list the advantages and drawbacks of these skillets. Write an advertisement for each skillet stressing only its positive features.



**Chem ASAP! Assessment 11.4** Check your understanding of the important ideas and concepts in Section 11.4.

# SMALL-SCALE LAB

## HEAT OF COMBUSTION OF A CANDLE

### SAFETY



Wear safety glasses and follow the standard safety procedures outlined on page 18. Keep the burning candle away from combustible materials.

### PURPOSE

To observe a burning candle and calculate the heat associated with the combustion reaction.

### MATERIALS

- candle
- aluminum foil
- safety matches
- ruler
- balance
- temperature probe (optional)

### PROCEDURE



Probe version available in the *Probeware Lab Manual*.

Measure and record the length of a candle in centimeters. Place the candle on a small piece of aluminum foil and measure the mass of the foil-candle system. Note the time as you light the candle. Let the candle burn for about five minutes. While you wait, begin answering the ANALYSIS questions. After about 5 minutes, extinguish the candle and record the time. Measure the mass of the foil-candle system again. Do not try to measure the mass while the candle is burning.

### ANALYSIS

Using your experimental data, answer the following questions.

1. Observe the candle burn and draw a picture of what you see.
2. Examine the flame closely. Is it the wax or the wick that burns?
3. If you said the wax, how does the wax burn without touching the flame? If you said the wick, what is the function of the wax?

4. If you could measure the temperature near the flame, you would find that the air is much hotter above the flame than it is beside it. Why? Explain.
5. Scientists have often wondered if a candle would burn well in zero gravity. How would zero gravity change the shape of the flame?
6. How much length and mass did the candle lose? Are these data more consistent with the wax or the wick burning?
7. Keeping in mind that 'wick' is also a verb, explain how a candle works.
8. The formula for candle wax can be approximated as  $C_{20}H_{42}$ . Write and balance an equation for the complete combustion of the candle wax.
9. Calculate the moles of candle wax burned in the experiment.
10. Calculate the heat of combustion of candle wax in kJ/mol. The standard heat of formation of candle wax ( $C_{20}H_{42}$ ) is  $-2230$  kJ/mol. The standard heats of formation of carbon dioxide and water are  $-394$  kJ/mol and  $-242$  kJ/mol, respectively. The heat of combustion of candle wax equals the sum of the heats of formation of the products minus the sum of the heats of formation of the reactants.
11. Calculate the kilojoules of heat released in your reaction. *Hint:* Multiply the number of moles of candle wax burned in the experiment by the heat of combustion of candle wax.

### YOU'RE THE CHEMIST!

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Design It!** Design an experiment to show that the candle wax does not burn with complete combustion.
2. **Design It!** Design an experiment to show that water is a product of the combustion of a candle.

## HARNESSING SOLAR ENERGY

When you consider energy sources other than oil or coal, you may think of solar power. Solar power is energy from the sun, and it is free, clean, and readily available.

In theory, solar power is surely the most plentiful energy source available. The enormous power of the sun's rays can be harnessed by solar, or photovoltaic, cells. In reality, this process has been faced with problems that have prevented solar power from becoming a more prevalent part of your life. First, the sun does not always shine on a desired location with maximum power.

The tilt of Earth on its axis, as well as weather conditions, can alter the amount of sunlight falling upon different locations. Second, solar cells are not very efficient at converting a "packet" of light energy, called a photon, into electricity. Finally, storing the electricity is not very practical because small batteries that can hold a lot of charge for extended periods of time have not yet been developed.

Instead of using the sun's energy to generate electricity, what about using it to heat an object directly? The sun's rays can be concentrated onto very small areas. Anyone who has used a magnifying glass to burn a hole in a sheet of paper knows this! The heat collected can be transferred to water, or oil, or even molten NaCl to be stored, or piped, to a remote area. This stored heat energy can then be converted to electricity. One of the drawbacks of this scheme is that some heat will always be lost because the heat-storing system is not at thermal equilibrium with its surroundings.

Scientists have also begun to take advantage of storing solar energy in the form of chemical energy. Think of it this way: It "costs" a certain amount of



energy to break a chemical bond in a molecule; when the process is reversed, the energy "cost" is refunded in the form of heat energy. The first process is an example of an endothermic reaction, which requires heat to be supplied to the system for the reaction to occur. The second (reverse) reaction is exothermic, during which heat is re-

leased as a reaction product.

One system that uses endothermic and exothermic reactions has been successfully tested. In this system,

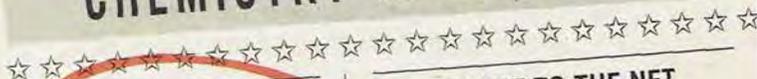
the sun's rays are concentrated to heat a "solar reactor" to about 800 °C, where methane (CH<sub>4</sub>) and carbon dioxide (CO<sub>2</sub>) absorb the heat to form carbon monoxide (CO) and hydrogen gas (H<sub>2</sub>). The product gases (CO and H<sub>2</sub>) are cooled to room temperature and piped to a destination where they are recombined to release the energy that is stored in chemical bonds. This approach is significant because the product gases are not hot, so they do not lose heat to the surroundings. After the CO and H<sub>2</sub> react, heat, CH<sub>4</sub>, and CO<sub>2</sub> are produced. The gases are used to start the process over again, and the heat is used

to generate electricity.

Solar energy is a very dynamic field. Many researchers are actively seeking new and better ways to harness the world's most plentiful energy resource.

In theory, solar power is the most plentiful energy source available. The enormous power of the sun's rays can be harnessed by solar, or photovoltaic, cells.

## CHEMISTRY IN CAREERS



### SOLAR ENGINEER

Must have knowledge of photovoltaics. Use of experimental techniques to test cells and modules required.

See page 873.

CHEMIST NEEDED

### TAKE IT TO THE NET



Find out more about career opportunities:

[www.phschool.com](http://www.phschool.com)

### CHEMICAL SPECIALIST

Local food service distributor seeks responsible self-motivated individ-

# Chapter 11 STUDENT STUDY GUIDE



## Take It to the NET

For interactive study and review, go to [www.phschool.com](http://www.phschool.com)

## KEY TERMS

- ▶ calorie *p.* 295
- ▶ calorimeter *p.* 300
- ▶ calorimetry *p.* 300
- ▶ chemical potential energy *p.* 293
- ▶ endothermic process *p.* 294
- ▶ energy *p.* 293
- ▶ enthalpy (*H*) *p.* 301
- ▶ exothermic process *p.* 294
- ▶ heat *p.* 293
- ▶ heat capacity *p.* 296
- ▶ heat of combustion *p.* 305
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- ▶ joule *p.* 296
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- ▶ specific heat *p.* 297
- ▶ specific heat capacity *p.* 297
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- ▶ surroundings *p.* 294
- ▶ system *p.* 294
- ▶ thermochemical equation *p.* 303
- ▶ thermochemistry *p.* 293
- ▶ universe *p.* 294

## KEY EQUATIONS AND RELATIONSHIPS

- ▶ 1 Calorie = 1 kilocalorie = 1000 calories
- ▶ 1 J = 0.239 cal and 4.184 J = 1 cal
- ▶  $C = \frac{q}{m \times \Delta T} = \frac{\text{heat (joules or calories)}}{\text{mass (g)} \times \text{change in temperature (}^\circ\text{C)}}$
- ▶  $\Delta H = m \times C \times \Delta T$ , where  $\Delta T = T_f - T_i$
- ▶  $\Delta H^0 = \Delta H_f^0(\text{products}) - \Delta H_f^0(\text{reactants})$

## CONCEPT SUMMARY

### 11.1 The Flow of Energy—Heat

- Energy is the capacity to do work or to supply heat. The law of conservation of energy states that energy cannot be created or destroyed.
- A process is exothermic if heat flows from the system to the surroundings and endothermic if heat flows from the surroundings to the system.
- An object's heat capacity is the amount of heat it takes to change the object's temperature by exactly 1 °C. The specific heat capacity of a substance is the amount of heat it takes to raise the temperature of 1 g of the substance 1 °C.

### 11.2 Measuring and Expressing Heat Changes

- Thermochemical equations show the accompanying heat of reaction at constant pressure.
- Calorimetry measures heat changes associated with chemical reactions and phase changes.

### 11.3 Heat in Changes of State

- The molar heats of fusion, solidification, vaporization, and condensation describe the quantity of heat transferred to or from a system when one mole of substance undergoes a change of state at constant temperature.

### 11.4 Calculating Heat Changes

- Hess's law states that in a reaction that is the sum of two or more other reactions,  $\Delta H$  for the overall process is the sum of the  $\Delta H$  values for all of the constituent reactions.
- The enthalpy for the formation of one mole of a compound from its elements in standard states is the enthalpy of formation ( $H_f^0$ ).
- The standard enthalpy change for a reaction ( $\Delta H^0$ ) can be calculated from the  $H_f^0$  of the reactants and products.

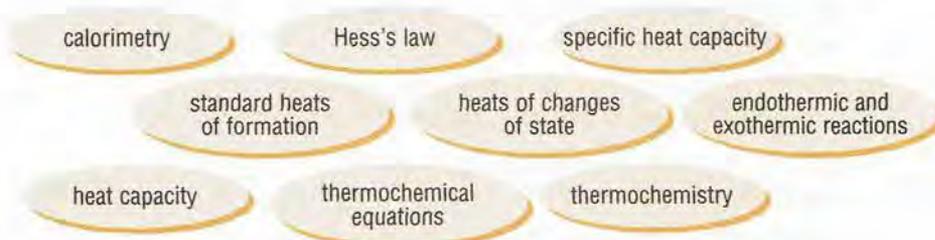
## CHAPTER CONCEPT MAP

Use these terms to construct a concept map that organizes the major ideas of this chapter.



### Chem ASAP! Concept Map 11

Create your Concept Map using the computer.

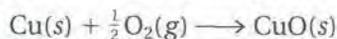


# Chapter 11 REVIEW

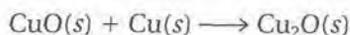
## CONCEPT PRACTICE

36. Explain in your own words the law of conservation of energy. 11.1
37. What always happens when two objects of different temperatures come in contact? Give an example from your own experience. 11.1
38. Define potential energy in terms of chemistry. 11.1
39. What factors determine the heat capacity of an object? 11.1
40. What is the relationship between a calorie and a Calorie? 11.1
41. Make the following conversions. 11.1
- $8.50 \times 10^2$  cal to Calories
  - 444 cal to joules
  - 1.8 kJ to joules
  - $4.5 \times 10^{-1}$  kJ to calories
42. Why do you think it is important to define the system and the surroundings? 11.1
43. Describe the sign convention that is used in thermochemical calculations. 11.1
44. Two substances in a glass beaker chemically react, and the glass beaker becomes too hot to touch. 11.1
- Is this an exothermic or endothermic reaction?
  - If the two substances are defined as the system, what constitutes the surroundings?
45. Classify these processes as exothermic or endothermic. 11.1
- condensing steam
  - evaporating alcohol
  - burning alcohol
  - baking a potato
46. What special name is given to a heat change at constant pressure? 11.2
47. What is the function of a calorimeter? 11.2
48. There are some obvious sources of error in experiments that use foam cups as calorimeters. Name at least three. 11.2
49. What special device would you use to measure the heat released at constant volume? 11.2
50. Give the standard conditions for heat of combustion. 11.2
51. What information is given in a thermochemical equation? 11.2
52. Explain why ice melts at 0 °C without an increase of temperature, even though heat is flowing from the surroundings into the system (the ice). 11.3
53. Calculate the quantity of heat gained or lost in the following changes. 11.3
- 3.50 mol of water freezes at 0 °C.
  - 0.44 mol of steam condenses at 100 °C.
  - 1.25 mol NaOH(s) dissolves in water.
  - 0.15 mol C<sub>2</sub>H<sub>5</sub>OH(l) vaporizes at 78.3 °C.
54. Sodium acetate dissolves readily in water according to the following equation. 11.3
- $$\text{NaC}_2\text{H}_3\text{O}_2(\text{s}) \longrightarrow \text{NaC}_2\text{H}_3\text{O}_2(\text{aq})$$
- $$\Delta H = -17.3 \text{ kJ/mol}$$
- Would this process increase or decrease the temperature of the water?
55. Explain the usefulness of Hess's law of heat summation in thermochemistry. 11.4
56. A considerable amount of heat is required for the decomposition of aluminum oxide. 11.4
- $$2\text{Al}_2\text{O}_3(\text{s}) \longrightarrow 4\text{Al}(\text{s}) + 3\text{O}_2(\text{g})$$
- $$\Delta H = 3352 \text{ kJ}$$
- What is the heat change for the formation of 1 mol of aluminum oxide from its elements?
  - Is the reaction exothermic or endothermic?
57. Calculate the heat change for the formation of lead(IV) chloride by the reaction of lead(II) chloride with chlorine. 11.4
- $$\text{PbCl}_2(\text{s}) + \text{Cl}_2(\text{g}) \longrightarrow \text{PbCl}_4(\text{l})$$
- $$\Delta H = ?$$
- Use the following thermochemical equations.
- $$\text{Pb}(\text{s}) + 2\text{Cl}_2(\text{g}) \longrightarrow \text{PbCl}_4(\text{l})$$
- $$\Delta H = -329.2 \text{ kJ}$$
- $$\text{Pb}(\text{s}) + \text{Cl}_2(\text{g}) \longrightarrow \text{PbCl}_2(\text{s})$$
- $$\Delta H = -359.4 \text{ kJ}$$
58. From the following reactions: 11.4
- $$\frac{1}{2}\text{N}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \longrightarrow \text{NO}(\text{g})$$
- $$\Delta H = 90.4 \text{ kJ/mol}$$
- $$\frac{1}{2}\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow \text{NO}_2(\text{g})$$
- $$\Delta H = 33.6 \text{ kJ/mol}$$
- determine the heat of reaction for:
- $$\text{NO}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \longrightarrow \text{NO}_2(\text{g})$$
- $$\Delta H = ?$$

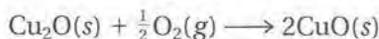
59. Calculate the heat change for the formation of copper(I) oxide from its elements. 11.4



Use the following thermochemical equations to make the calculation.



$$\Delta H = -11.3 \text{ kJ}$$



$$\Delta H = -114.6 \text{ kJ}$$

60. What is the standard heat of formation of a free element in its standard state? 11.4
61. Consider the statement, "the more negative the value of  $\Delta H_f^\circ$ , the more stable the compound." Is this statement true or false? Explain. 11.4
62. Calculate the change in enthalpy (in kJ) for the following reactions. 11.4
- $\text{CH}_4(g) + 2\text{O}_2(g) \longrightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l)$
  - $2\text{CO}(g) + \text{O}_2(g) \longrightarrow 2\text{CO}_2(g)$
63. What is the standard heat of formation of a compound? 11.4

## CONCEPT MASTERY

64. Equal masses of two substances absorb the same amount of heat. The temperature of substance A increases twice as much as the temperature of substance B. Which substance has the higher specific heat? Explain.
65. If 3.20 kcal of heat is added to 1.00 kg of ice at 0 °C, how much water at 0 °C is produced, and how much ice remains?
66. The amounts of heat required to change different quantities of carbon tetrachloride ( $\text{CCl}_4(l)$ ) into vapor are given in the table.

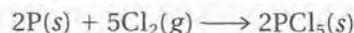


Mass of $\text{CCl}_4$ (g)	Heat	
	(J)	(cal)
2.90	652	156
7.50	1689	404
17.0	3825	915
26.2	5894	1410
39.8	8945	2140
51.0	11453	2740

- Graph the data, using heat as the dependent variable.
- What is the slope of the line?
- The heat of vaporization of  $\text{CCl}_4(l)$  is 53.8 cal/g. How does this value compare with the slope of the line?

67. Calculate the heat change in calories when 45.2 g of steam at 100 °C condenses to water at the same temperature. What is the heat change in joules?

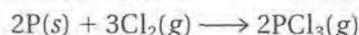
68. Find the enthalpy change for the formation of phosphorus pentachloride from its elements.



Use the following thermochemical equations.



$$\Delta H = 87.9 \text{ kJ}$$



$$\Delta H = -574 \text{ kJ}$$

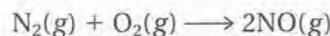
69. Use standard heats of formation ( $\Delta H_f^\circ$ ) to calculate the change in enthalpy for these reactions.
- $2\text{C}(s, \text{graphite}) + \text{O}_2(g) \longrightarrow 2\text{CO}(g)$
  - $2\text{H}_2\text{O}_2(l) \longrightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g)$
  - $4\text{NH}_3(g) + 5\text{O}_2(g) \longrightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)$
70. An ice cube with a mass of 40.0 g melts in water originally at 25.0 °C.
- How much heat does the ice cube absorb from the water when it melts? Report your answer in calories, kilocalories, and joules.
  - Calculate the number of grams of water that can be cooled to 0 °C by the melting ice cube.
71. The molar heat of vaporization of ethanol ( $\text{C}_2\text{H}_5\text{OH}(l)$ ) is 43.5 kJ/mol. Calculate the heat required to vaporize 25.0 g of ethanol at its boiling point.
72. An orange contains 445 kJ of energy. What mass of water could this same amount of energy raise from 25.0 °C to the boiling point?
73. The combustion of ethane ( $\text{C}_2\text{H}_6$ ) is an exothermic reaction.



$$\Delta H = -1.39 \times 10^3 \text{ kJ}$$

Calculate the amount of heat liberated when 4.79 g  $\text{C}_2\text{H}_6$  reacts with excess oxygen.

74. Calculate the heat change ( $\Delta H$ ) for the formation of nitrogen monoxide from its elements.



Use these thermochemical equations.



$$\Delta H = -1.53 \times 10^3 \text{ kJ}$$



$$\Delta H = -1.17 \times 10^3 \text{ kJ}$$

75. How much heat must be removed from a 45.0-g sample of naphthalene ( $C_{10}H_8$ ) at its freezing point to bring about solidification? The heat of fusion of naphthalene is 191.2 kJ/mol.

### CRITICAL THINKING

76. Choose the term that best completes the second relationship.
- a. kilojoules:heat  
 $cm^3$ : \_\_\_\_\_  
 (1) mass (3) energy  
 (2) volume (4) weight
- b. right:left  
 exothermic: \_\_\_\_\_  
 (1) combustion (3) endothermic  
 (2) heat (4) joule
- c. thermometer:temperature  
 calorimeter: \_\_\_\_\_  
 (1) constant pressure (3) endothermic  
 (2) reactants (4) heat
77. Refer to Figure 11.15. Which region of the graph represents the coexistence of solid and liquid? Liquid and vapor?

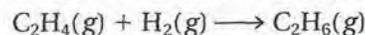
### CUMULATIVE REVIEW

78. What fraction of the average kinetic energy of hydrogen gas at 100 K does hydrogen gas have at 40 K?
79. How many hydrogen molecules are in 44.8 L  $H_2(g)$  at STP?
80. Write the net ionic equation for the reaction of aqueous solutions of sodium chloride and silver acetate.
81. How many grams of oxygen are formed by the decomposition of 25.0 g of hydrogen peroxide?  
 $2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$

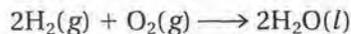
### CONCEPT CHALLENGE

82. The temperature of a person with an extremely high fever can be lowered with a sponge bath of isopropyl alcohol ( $C_3H_7OH$ ). The heat of vaporization of this alcohol is 11.1 kcal/mol.
- a. How many kilocalories of heat are removed from a person's skin when 175 g of isopropyl alcohol evaporates? How many kilojoules?
- b. How many kilograms of water would this energy loss cool in lowering the temperature from 40.0 °C to 36.0 °C?

83. Ethane ( $C_2H_6(g)$ ) can be formed by the reaction of ethene ( $C_2H_4(g)$ ) with hydrogen gas.



Use the heats of combustion for the following reactions to calculate the heat change for the formation of ethane from ethene and hydrogen.



$$\Delta H = -5.72 \times 10^2 \text{ kJ}$$

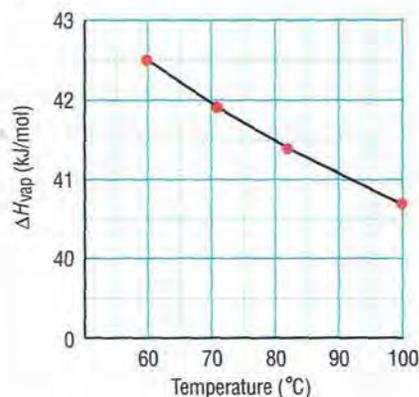


$$\Delta H = -1.401 \times 10^3 \text{ kJ}$$



$$\Delta H = -3.100 \times 10^3 \text{ kJ}$$

84. An ice cube at 0 °C was dropped into 30.0 g of water in a cup at 45.0 °C. At the instant that all of the ice was melted, the temperature of the water in the cup was 19.5 °C. What was the mass of the ice cube?
85. 41.0 g of glass at 95 °C is placed in 175 g of water at 21 °C in an insulated container. They are allowed to come to the same temperature. What is the final temperature of the glass-water mixture? The specific heat of glass is 2.1 cal/(g × °C).
86. The enthalpy of vaporization of water at various temperatures is given in the graph. From this graph, estimate the amount of heat required to convert 1 L of water to steam on the summit of Mount Everest (29 002 ft), where the boiling temperature of water is 70 °C.



# Chapter 11 STANDARDIZED TEST PREP

Select the choice that best answers each question or completes each statement.

- The  $\Delta H_{\text{fus}}$  of ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) is 4.60 kJ/mol. How many kilojoules are required to melt 24.5 g of ethanol at its freezing point?
  - 2.45 kJ
  - 5.33 kJ
  - 245 kJ
  - 18.8 kJ
- How much heat, in kilojoules, must be added to 178 g of water to increase the temperature of the water by 5.0 °C?
  - 890 kJ
  - 36 kJ
  - 3.7 kJ
  - 0.093 kJ
- The standard heat of formation of a free element in its standard state is always
  - zero.
  - positive.
  - negative.
  - higher for solids than gases.
- If  $\Delta H$  for the reaction  $2\text{HgO}(s) \longrightarrow 2\text{Hg}(l) + \text{O}_2(g)$  is +181.66 kJ, then  $\Delta H$  for the reaction  $\text{Hg}(l) + 1/2\text{O}_2(g) \longrightarrow \text{HgO}(s)$  is
  - 90.83 kJ.
  - 90.83 kJ.
  - 181.66 kJ.
  - 181.66 kJ.
- The specific heat capacity of grain alcohol is ten times larger than the specific heat capacity of silver. A hot bar of silver with a mass of 55 g is dropped into an equal mass of cool alcohol. If the temperature of the silver bar drops 45 °C, the temperature of the alcohol
  - increases 45 °C.
  - decreases 4.5 °C.
  - increases 4.5 °C.
  - decreases 45 °C.

The lettered choices below refer to questions 6–9. A lettered choice may be used once, more than once, or not at all.

- (A) kJ/mol  
 (B)  $(\text{J} \times ^\circ\text{C})/\text{g}$   
 (C)  $\text{J}/(\text{g} \times ^\circ\text{C})$   
 (D) kJ  
 (E)  $\text{kJ}/^\circ\text{C}$

Which unit is appropriate for each of the following measurements?

- heat of reaction
- heat capacity
- molar heat of fusion
- specific heat capacity

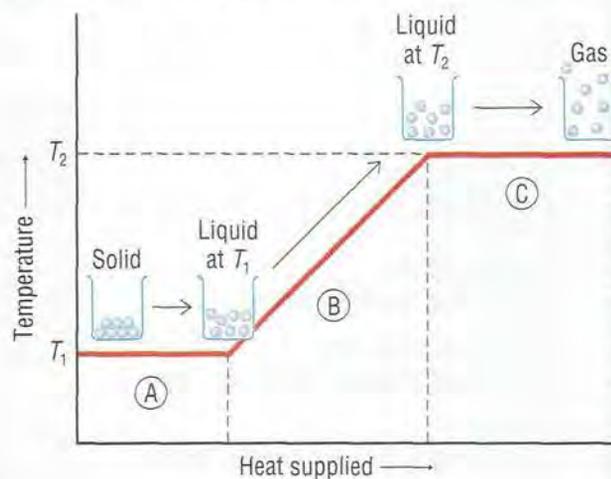
Use the data table to answer questions 10–12.

Alkanes are a class of compounds containing carbon and hydrogen that are often used as fuels.

Alkane	Heat of combustion (kJ/mol)
Methane, $\text{CH}_4$	-890
Ethane, $\text{C}_2\text{H}_6$	-1560
Propane, $\text{C}_3\text{H}_8$	-2220
Butane, $\text{C}_4\text{H}_{10}$	
Pentane, $\text{C}_5\text{H}_{12}$	-3540
Hexane, $\text{C}_6\text{H}_{14}$	-4160
Heptane, $\text{C}_7\text{H}_{16}$	-4810

- Graph the data. Make the number of carbon atoms per molecule the independent variable.
- Estimate the heat of combustion for butane.
- Estimate the heat of combustion for the next alkane in the series, octane,  $\text{C}_8\text{H}_{18}$ .

Use the graph and table to answer questions 13–16. Assume 1.00 mol of substance in each container.



Substance	Freezing point (K)	$\Delta H_{\text{fus}}$ (kJ/mol)	Boiling point (K)	$\Delta H_{\text{vap}}$ (kJ/mol)
Ammonia	195.3	5.65	239.7	23.4
Benzene	278.7	9.87	353.3	30.8
Methanol	175.5	3.16	337.2	35.3
Neon	24.5	0.33	27.1	1.76

- Calculate heat absorbed in region A for neon.
- Calculate heat absorbed in region C for benzene.
- Calculate heat absorbed in regions B and C for methanol. [specific heat =  $81.6 \text{ J}/(\text{g} \times ^\circ\text{C})$ ]
- Calculate heat absorbed in regions A, B, and C for ammonia. [specific heat =  $35.1 \text{ J}/(\text{g} \times ^\circ\text{C})$ ]