Is evaporation endothermic or exothermic? (endothermic)
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 of water. 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 of water.

\[ \text{H}_2\text{O}(s) \rightarrow \text{H}_2\text{O}(l) \quad \Delta H_{\text{fus}} = 6.01 \text{ kJ/mol} \]

\[ \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s) \quad \Delta H_{\text{solid}} = -6.01 \text{ kJ/mol} \]

**SAMPLE PROBLEM 17.4**

**Using the Heat of Fusion in Phase-Change Calculations**

How many grams of ice at 0°C will melt if 2.25 kJ of heat are added?

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

   **Knowns**
   - Initial and final temperatures are 0°C
   - \( \Delta H_{\text{fus}} = 6.01 \text{ kJ/mol} \)
   - \( \Delta H = 2.25 \text{ kJ} \)

   **Unknown**
   - \( m_{\text{ice}} = ? \text{ g} \)

   Use the thermochemical 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.

   Express \( \Delta H_{\text{fus}} \) and the molar mass of ice as conversion factors.

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

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

   \[ m_{\text{ice}} = \frac{2.25 \text{ kJ}}{6.01 \text{ kJ}} \times \frac{1 \text{ mol ice}}{18.0 \text{ g ice}} \times \frac{18.0 \text{ g ice}}{1 \text{ mol ice}} \]

   \[ = 6.74 \text{ g ice} \]

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 18.0 g/3 = 6 g, should melt. This estimate is close to the calculated answer.

**Practice Problems**

21. How many kilojoules of heat are required to melt a 10.0-g popsicle at 0°C? Assume the popsicle has the same molar mass and heat of fusion as water.

22. How many grams of ice at 0°C could be melted by the addition of 0.400 kJ of heat?
Heat of Fusion of Ice

**Objective** Students will melt an ice cube and calculate the heat of fusion of ice.

**Skills Focus** Observing, measuring, calculating, designing experiments

**Quick LAB**

**Heat of Fusion of Ice**

**Purpose** To estimate the heat of fusion of ice.

**Materials**
- ice
- foam cup
- 100-ml graduated cylinder
- thermometer
- hot water
- temperature probe (optional)

**Procedure**
1. Fill the graduated cylinder with hot tap water and let 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 an ice cube to the cup of water. Gently swirl the cup. Measure the temperature of the water as soon as the ice cube has completely melted.
4. Pour the water into the graduated cylinder and measure the volume.

**Analyze and Conclude**
1. Calculate the mass of the ice. (Hint: The mass of ice melted is the same as the increase in the volume of the water.) Convert this into moles.
2. Calculate \( \Delta H_{fus} \) of ice (kJ/mol) by dividing the heat transferred from the water by the moles of ice melted.
3. Compare your experimental value of \( \Delta H_{fus} \) of ice with the accepted value of 6.01 kJ/mol. Account for any error.
4. How might you revise the procedure to achieve more accurate results?

**Heats of Vaporization and Condensation**

When liquids absorb heat at their boiling points, they become vapors. The amount of heat necessary to vaporize one mole of a given liquid is called its molar heat of vaporization \( \Delta H_{vap} \). Table 17.3 lists the molar heats of vaporization for several substances at their normal boiling points.

The molar heat of vaporization of water is 40.7 kJ/mol. This means that it takes 40.7 kJ of energy to convert 1 mol of water molecules in the liquid state to 1 mol of water molecules in the vapor state at the normal boiling point of water (100°C at 101.3 kPa). This process is described in the thermochemical equation below:

\[
H_2O(l) \rightarrow H_2O(g) \quad \Delta H_{vap} = 40.7 \text{ kJ/mol}
\]

**Table 17.3**

<table>
<thead>
<tr>
<th>Substance</th>
<th>( \Delta H_{fus} ) (kJ/mol)</th>
<th>( \Delta H_{vap} ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonia (NH₃)</td>
<td>5.65</td>
<td>23.4</td>
</tr>
<tr>
<td>Ethanol (C₂H₅OH)</td>
<td>4.60</td>
<td>43.5</td>
</tr>
<tr>
<td>Hydrogen (H₂)</td>
<td>0.12</td>
<td>0.90</td>
</tr>
<tr>
<td>Methanol (CH₃OH)</td>
<td>3.16</td>
<td>35.3</td>
</tr>
<tr>
<td>Oxygen (O₂)</td>
<td>0.44</td>
<td>6.82</td>
</tr>
<tr>
<td>Water (H₂O)</td>
<td>6.01</td>
<td>40.7</td>
</tr>
</tbody>
</table>

4. Acceptable answers include the following:
   - To determine the mass of warm water cooled and the mass of ice melted, measure the mass of (1) the empty cup, (2) the cup plus the hot water, and (3) the cup after the ice has melted. Use a covered calorimeter to decrease the influence of room temperature.

**For Enrichment**

Have students design a new experiment that incorporates their suggestions for revising the procedure (their answers to Question 4). If students perform the new experiment, they should compare the results of the two experiments and calculate the respective percent errors.
Diethyl ether (C₄H₁₀O) 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 condensate 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 (ΔHᵥap) of 15.7 kJ/mol.

Condensation is the exact opposite of vaporization. When a vapor condenses, heat is released. The amount of heat released when 1 mol of vapor condenses at the normal boiling point is called its molar heat of condensation (ΔH.cond). This value is numerically the same as the corresponding molar heat of vaporization, however, the value has the opposite sign. The quantity of heat absorbed by a vaporizing liquid is exactly the same as the quantity of heat released when the vapor condenses; that is, ΔHᵥap = −ΔH.cond. Figure 17.10 summarizes the enthalpy changes that occur as a solid is heated to a liquid and then to a vapor. 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 ΔHᵥap and ΔH.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 the steam touches your skin.

H₂O(g) → H₂O(l)  ΔH(cond) = −40.7 kJ/mol

Figure 17.10 A heating curve graphically describes the enthalpy changes that take place during phase changes.

Heats of Vaporization and Condensation

**Interpreting Graphs**

a. at the melting point and the boiling point
b. It takes much less energy to melt a given mass of ice than to vaporize the same mass of water. The ΔH_fus is much smaller than ΔHᵥap.
c. the plateau of the curve between solid and liquid; the plateau between liquid and vapor

**Enrichment Question**

How would the rate of heating affect the heating curve? (The faster the rate of heating is, the steeper the slope of the curve will be.)

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**Exothermic Solidification**

**Purpose** Students will observe the release of heat when a liquid solidifies.

**Materials** 3 wide-mouth test tubes, sodium thiosulfate pentahydrate, hot water bath, 30°C water bath, thermometers, stirring rods

**Procedure** Set three wide-mouth test tubes each containing about 20 grams of sodium thiosulfate pentahydrate (Na₂S₂O₃•5H₂O) in a hot water bath until the chemical is completely melted. Transfer the tubes to a cool water bath at about 30°C. Let the test tubes sit undisturbed until they are close to 30°C and feel lukewarm to touch. Give the tubes to three student volunteers at different locations in the classroom so that many students can see and feel what happens. Give each volunteer a thermometer and a seed crystal. Ask them to quickly take the temperature of the liquid, drop in the seed crystal, and stir gently.

**Expected Outcomes** The temperature of the test tube will rise. Students will observe the change by touching the test tube and reading the thermometer.

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**Facts and Figures**

**Ether**

Before the mid-1840s, patients undergoing surgery received no anesthesia, and they had to endure intense pain. Then the anesthetic properties of diethyl ether (commonly called ether) and chloroform were discovered. By the Civil War, use of these two substances was common; when available, they were used in surgeries on the battlefield. Anesthetics were generally applied by pouring the agent on linen, holding it at some distance from the patient’s nose and mouth for the first inhalation, and then gradually moving the cloth closer until the desired effects were produced. Ether is no longer used as an anesthetic because it is flammable and causes many undesirable side reactions in patients.
Sample Problem 17.5

Answers
23. \(1.44 \times 10^2 \text{ kJ}\)
24. \(1.9 \times 10^{-1} \text{ kJ}\)

Practice Problems Plus

How many kilojoules of heat are required to vaporize 50.0 g of ethanol, \(\text{C}_2\text{H}_5\text{OH}\)? The boiling point of ethanol is 78.5°C. Its molar heat of vaporization is 43.5 kJ/mol. (47.3 kJ)

Math Handbook

For a math refresher and practice, direct students to using a calculator, page R62.

CLASS Activity

Heating Curve for Ethanol

Have students look up the molar heat of fusion, the molar heat of vaporization, and the specific heat of ethanol (grain alcohol) in Tables 17.1 and 17.3. Tell students that ethanol freezes at 158.7 K and boils at 351.5 K. Have students summarize the enthalpy changes that occur as solid ethanol is heated to a liquid and then to a gas by constructing a heating curve like the one shown in Figure 17.10 for water. Ask students to find the amount of heat necessary to convert 1 mole of solid ethanol to ethanol vapor.

SAMPLE PROBLEM 17.5

Using the Heat of Vaporization in Phase-Change Calculations

How much heat (in kJ) is absorbed when 24.8 g \(\text{H}_2\text{O}(l)\) at 100°C and 101.3 kPa is converted to steam at 100°C?

1. Analyze
   List the knowns and the unknown.

   Knowns
   - Initial and final conditions are 100°C and 101.3 kPa
   - mass of water converted to steam = 24.8 g
   - \(\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}\)

   Unknown
   - \(\Delta H = ? \text{ kJ}\)

   Refer to the following thermochemical equation.
   \[
   \text{H}_2\text{O}(l) + 40.7 \text{ kJ/mol} \rightarrow \text{H}_2\text{O}(g)
   \]

   \(\Delta H_{\text{vap}}\) is given in kJ/mol, but the quantity of water is given in grams. You must first convert grams of water to moles of water. 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 mass of water in grams by the conversion factors.

   \[
   \Delta H = 24.8 \text{ g H}_2\text{O}(l) \times \frac{1 \text{ mol H}_2\text{O}(l)}{18.0 \text{ g H}_2\text{O}(l)} \times \frac{40.7 \text{ kJ}}{1 \text{ mol H}_2\text{O}(l)} = 56.1 \text{ kJ}
   \]

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 \(\text{H}_2\text{O}\). The calculated enthalpy change should be a little less than \(1.5 \text{ mol} \times 40 \text{ kJ/mol} = 60 \text{ kJ}\), and it is.

Practice Problems

23. How much heat is absorbed when 63.7 g \(\text{H}_2\text{O}(l)\) at 100°C and 101.3 kPa is converted to steam at 100°C? Express your answer in kJ.

24. 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 normal boiling point? The molar heat of vaporization of chloroethane is 26.4 kJ/mol.
Heat of Solution

If you’ve ever used a hot pack or a cold pack, then you have felt the enthalpy changes that occur when a solute dissolves in a solvent. During the formation of a solution, heat is either released or absorbed. The enthalpy change caused by dissolution of one mole of substance is the molar heat of solution (ΔH_solution). Sodium hydroxide provides a good example of an exothermic molar heat of solution. When 1 mol of sodium hydroxide (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 interact with the water. The temperature of the solution increases, releasing 445.1 kJ of heat as the molar heat of solution.

NaOH(s) \( \rightarrow \) Na\(^{+}\)(aq) + OH\(^{-}\)(aq)

ΔH_solution = −445.1 kJ/mol

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

CaCl\(_2\)(s) \( \rightarrow \) Ca\(^{2+}\)(aq) + 2Cl\(^{-}\)(aq)

ΔH_solution = −82.8 kJ/mol

The dissolution of ammonium nitrate (NH\(_4\)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. Instant cold packs, which are used to treat muscle aches and sore joints, work by endothermic dissolving. The cold pack in Figure 17.11 contains solid ammonium nitrate crystals and water. Once the solute dissolves in the solvent, the pack becomes cold. In this case, the solution process absorbs energy from the surroundings.

NH\(_4\)NO\(_3\)(s) \( \rightarrow \) NH\(_4\)\(^{+}\)(aq) + NO\(_3\)\(^{-}\)(aq)

ΔH_solution = 25.7 kJ/mol

What is the molar heat of solution?

Differentiated Instruction

Less Proficient Readers

Provide at least two brands of cold or hot packs. Have students activate the packs at exactly the same time and then measure the maximum or minimum temperature achieved. Have students determine how long the packs remain effective. Ask students to decide, based on the data and the cost of each pack, which brand provided the best value.

Hot and Cold Packs

• Purchase a number of hot and cold packs used by athletes and show the class how they work. Most instant cold and hot packs operate by utilizing the heat released or absorbed when certain substances dissolve in water. Hot packs usually contain calcium chloride, CaCl\(_2\)(s), while cold packs usually contain ammonium nitrate, NH\(_4\)NO\(_3\)(s). If the labels specify the amount of salt contained in each package, have students use the molar heats of solution provided in the text to calculate the amount of heat absorbed and released in each case.

• Have students design and perform an experiment that measures the temperature change in a hot or cold pack. Also, you could change the solute to something not mentioned in the text (e.g., sodium acetate), and then have students conclude whether the solution process is endothermic or exothermic.

Answers to...

Figure 17.11 The system is the cold pack; the athlete and the air around the cold pack constitute the surroundings.

Checkpoint the enthalpy change caused by the dissolution of one mole of a substance.
Section 17.3 Assessment

27. **Key Concept** How does the molar heat of fusion of a substance compare to its molar heat of solidification?

28. **Key Concept** How does the molar heat of vaporization of a substance compare to its molar heat of condensation?

29. **Key Concept** What enthalpy changes occur when a solute dissolves in a solvent?

30. Identify each enthalpy change by name and classify each change as exothermic or endothermic.
   - a. 1 mol C₃H₈(l) → 1 mol C₃H₈(g)
   - b. 1 mol Hg(l) → 1 mol Hg(s)
   - c. 1 mol NH₃(g) → 1 mol NH₃(l)
   - d. 1 mol NaCl(s) + 3.88 kJ/mol → 1 mol NaCl(aq)
   - e. NaCl(s) → NaCl(l)

31. Why is a burn from steam potentially far more serious than a burn from very hot water?

**Connecting Concepts**

Hydrogen Bonding Reread the description of intermolecular attractions in Section 8.3. Use what you know about hydrogen bonding to explain why water absorbs such a large amount of heat as it vaporizes.

**Interactive Textbook**

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 17.3.