Chapter 17: Thermochemistry

17.1 The Flow of Energy – Heat and Work

The Flow of Energy—Heat and Work

- The temperature of lava from a volcano ranges from 550°C to 1400°C. As lava flows, it loses heat and begins to cool. You will learn about heat flow and why some substances cool down or heat up more quickly than others.

Energy Transformations

- **Heat**, *q*, - energy that transfers from one object to another because of a temperature difference
  - Flows from warmer to cooler

Energy Transformations

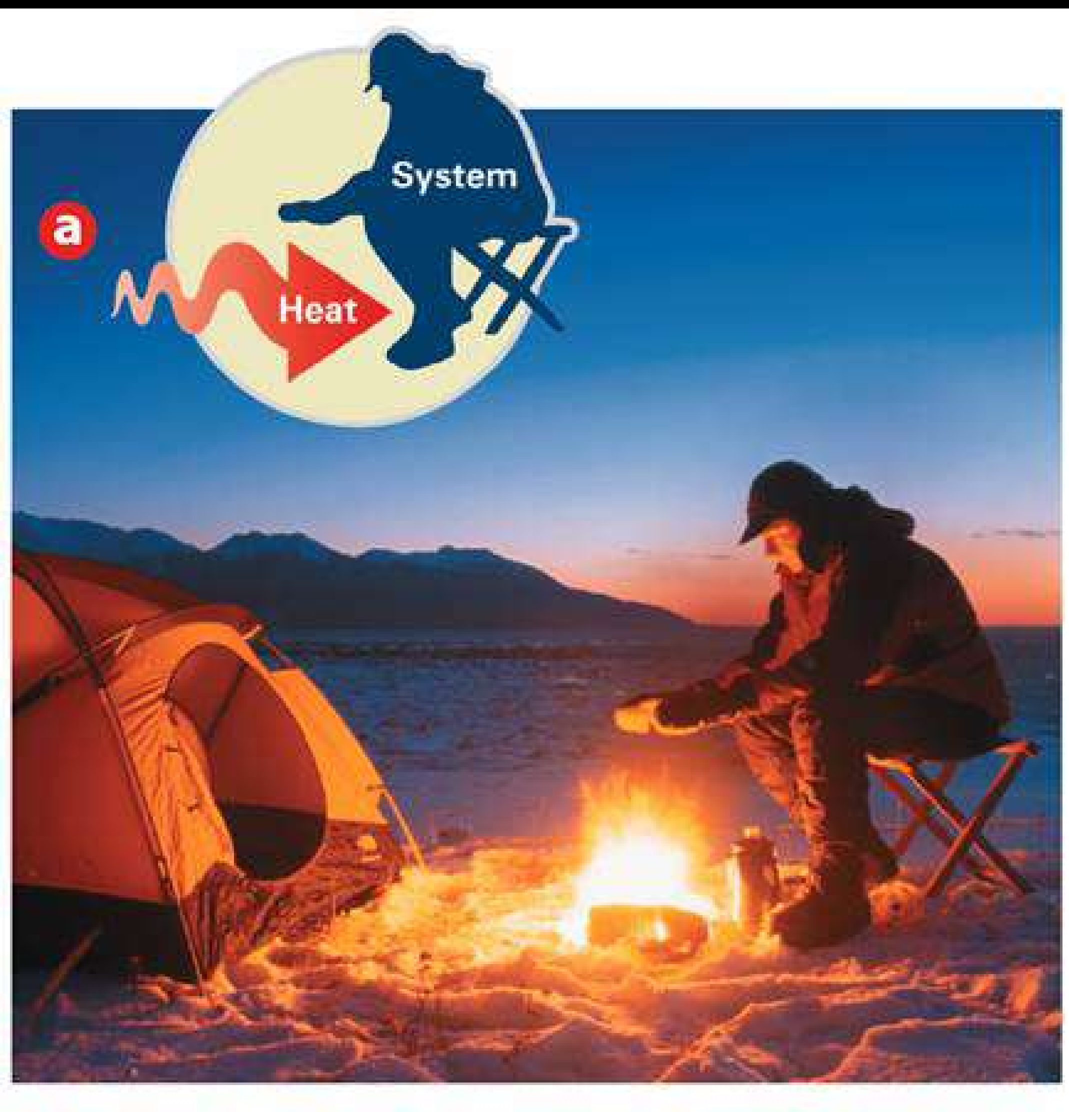
- **Thermochemistry** - study of energy changes that occur during chemical reactions and changes in state.
- Energy stored in the chemical bonds of a substance is called **chemical potential energy.**

Energy Transformations

- When fuel is burned in a car engine, chemical potential energy is released and is used to do work.

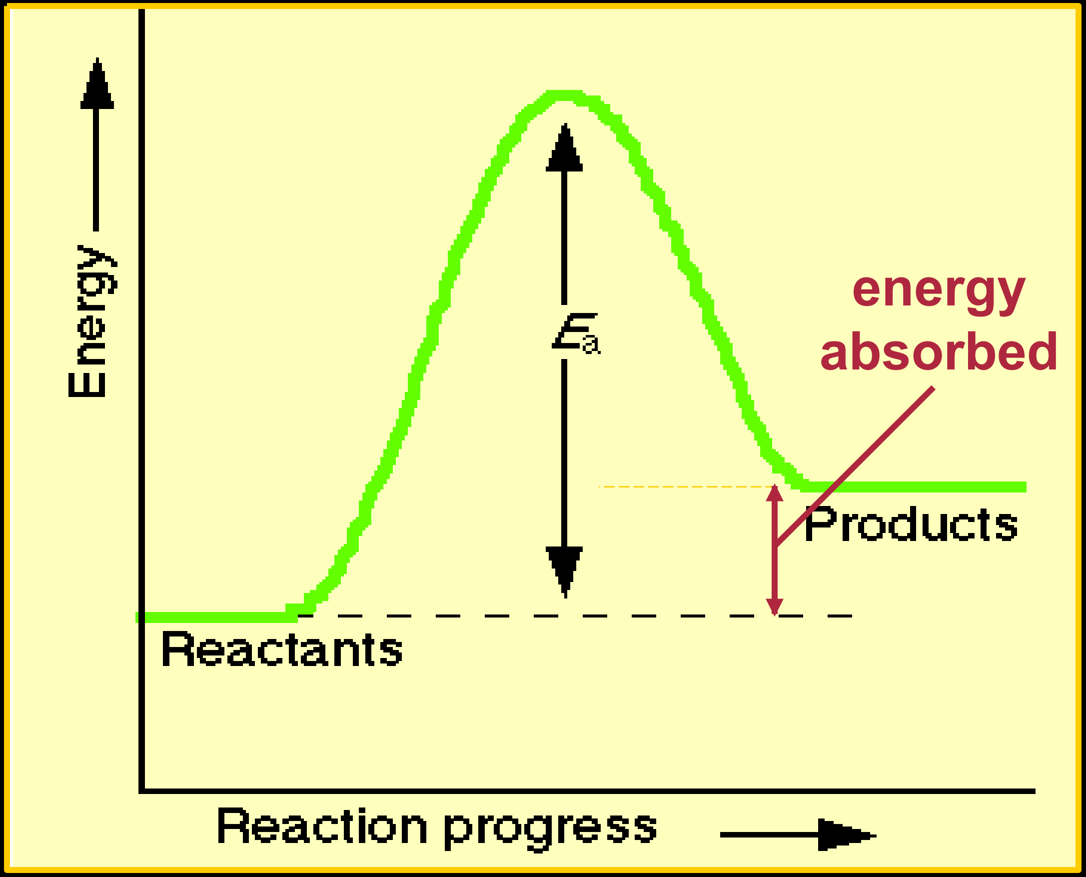
Exothermic and Endothermic Processes

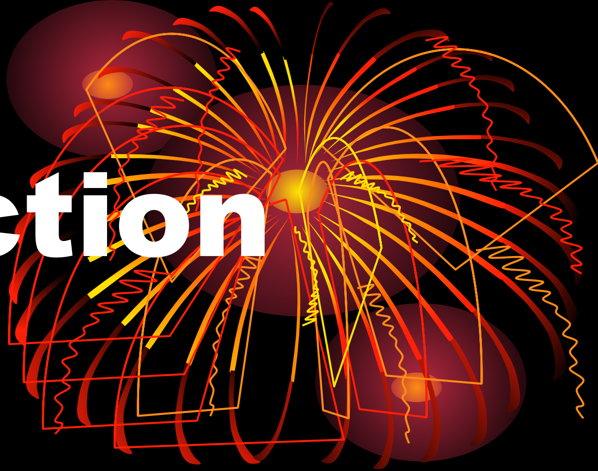
- **System** - part of the universe on which you focus your attention
- **Surroundings** - everything else in the universe.
- **Law of conservation of energy** - in any chemical or physical process, energy is neither created nor destroyed.



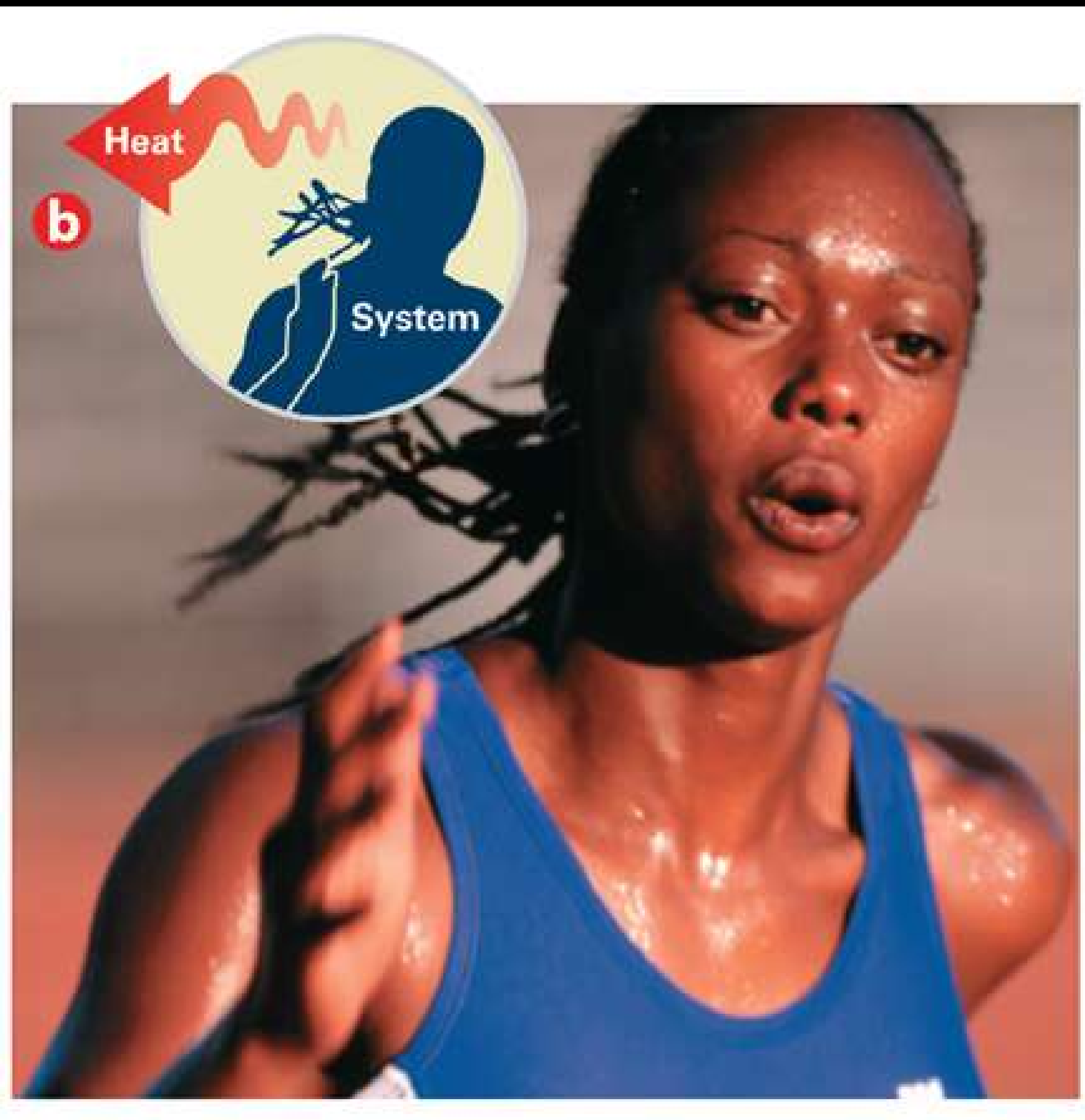
- An **endothermic process** is one that absorbs heat from the surroundings.

Endothermic Reaction



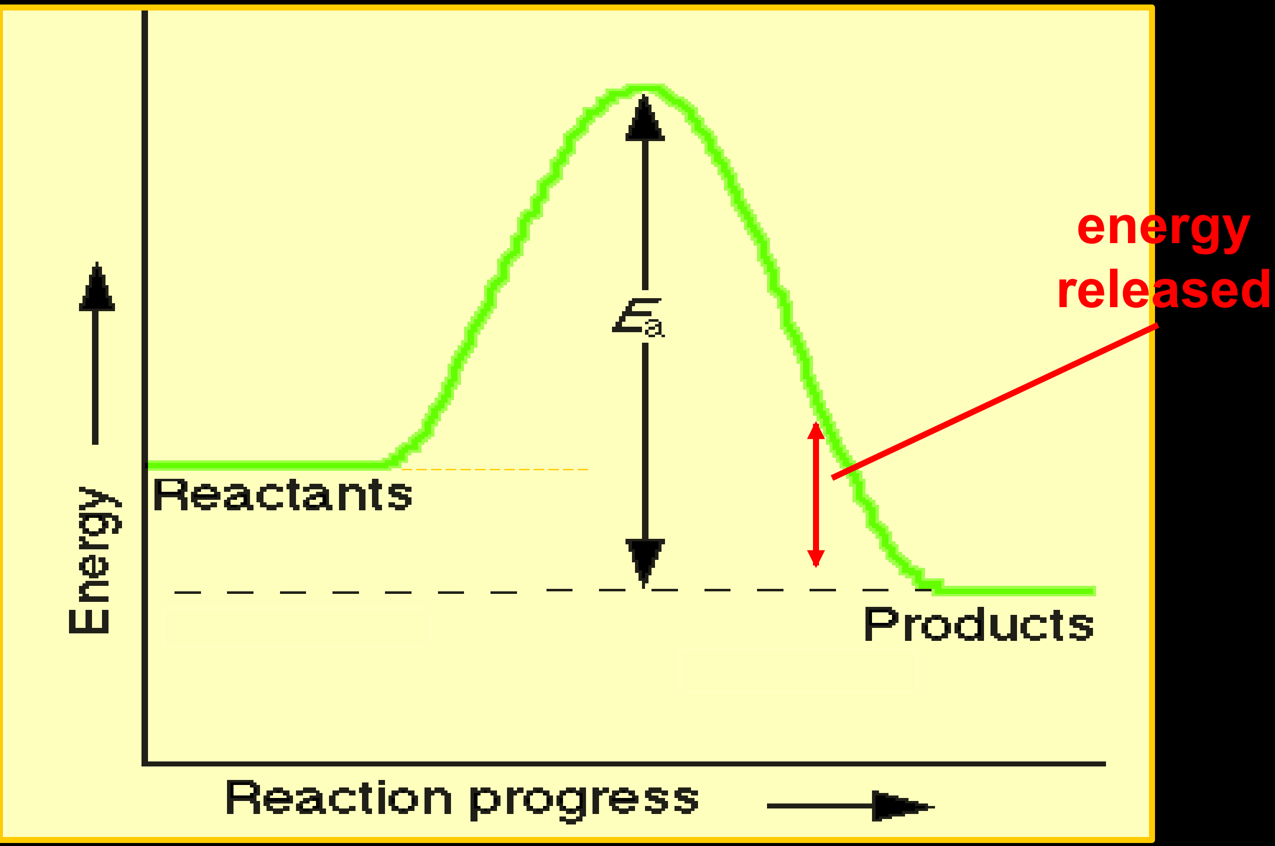


$$2Al_2O_3 + \text{energy} \rightarrow 4Al + 3O_2$$



- An **exothermic process** is one that releases heat to its surroundings.

Exothermic Reaction



$$2H_2(l) + O_2(l) \rightarrow 2H_2O(g) + \text{energy}$$

Conceptual Problem 17.1

Recognizing Exothermic and Endothermic Processes

On a sunny winter day, the snow on a rooftop begins to melt. As the melt-water drips from the roof, it refreezes into icicles. Describe the direction of heat flow as the water freezes. Is this process endothermic or exothermic?

Conceptual Problem 17.1

1 Analyze ***Identify the relevant concepts.***

Heat always flows from a warmer object to a cooler object. An endothermic process absorbs heat from the surroundings. An exothermic process releases heat to the surroundings.

Conceptual Problem 17.1

2 Solve ***Apply concepts to this situation.***

In order for water to freeze, its temperature must decrease. So heat must flow out of the water (the system). Because heat is released from the system to the surroundings (the air), the process is exothermic.

for Conceptual Problem 17.1

1. A container of melted paraffin wax is allowed to stand at room temperature until the wax solidifies. What is the direction of heat flow as the liquid wax solidifies? Is the process exothermic or endothermic?

Units for Measuring Heat Flow

- Heat flow is measured in two common units, the calorie and the joule (1 calorie = 4.184 J)
  - The energy in food is usually expressed in Calories.

1 Calorie = 1 kilocalorie = 1000 calories

calorie-joule conversions

- 1) 60.1 cal to joules
- 2) 34.8 cal to joules
- 3) 47.3 J to cal
- 4) 28.4 J to cal

Heat Capacity and Specific Heat

- The amount of heat needed to increase the temperature of an object exactly 1°C is the **heat capacity** of that object.
  - Depends on mass and chemical composition

Heat Capacity and Specific Heat

- **Specific heat** - amount of heat it takes to raise the temperature of 1 g of the substance 1°C.

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

Heat Capacity and Specific Heat

- Water releases a lot of heat as it cools. During freezing weather, farmers protect citrus crops by spraying them with water.

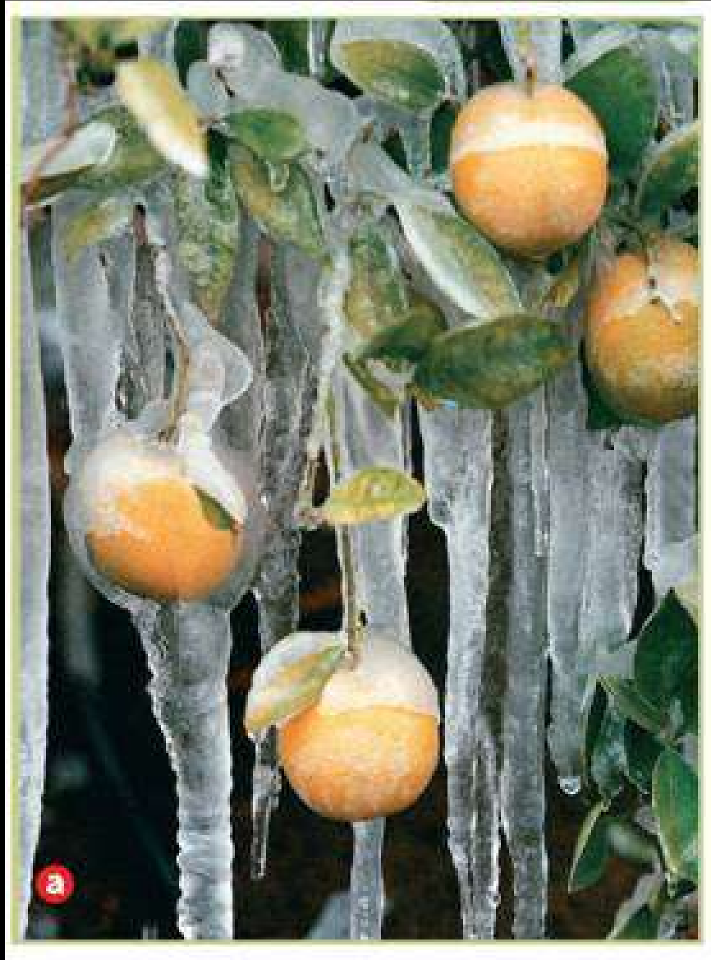


Table 17.1

Specific Heats of Some Common Substances

Substance	Specific Heat	
	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
Iron	0.46	0.11
Silver	0.24	0.057
Mercury	0.14	0.033

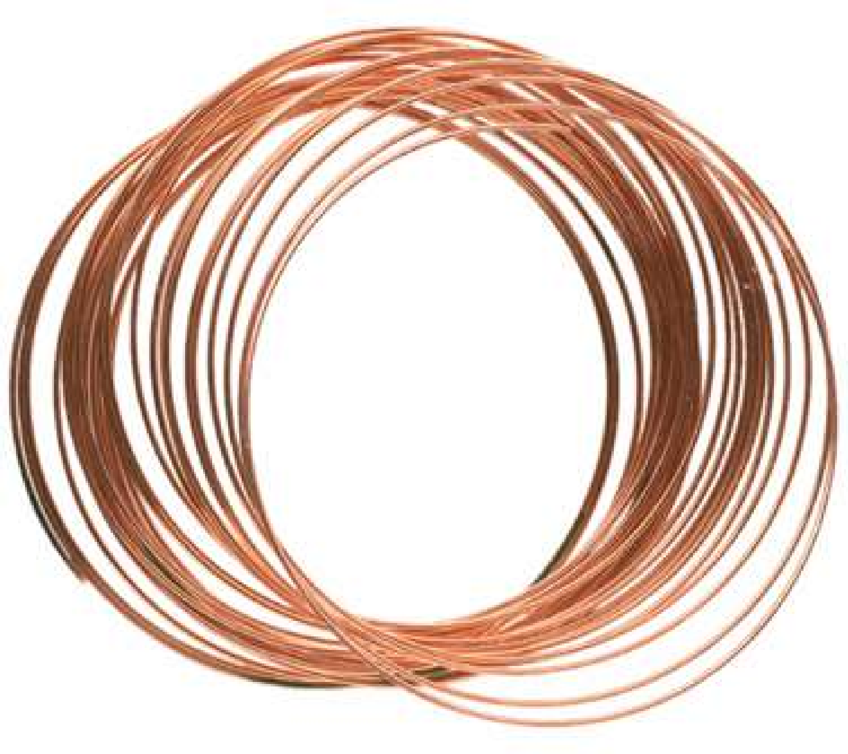
Heat Capacity and Specific Heat

- Because it is mostly water, the filling of a hot apple pie is much more likely to burn your tongue than the crust.

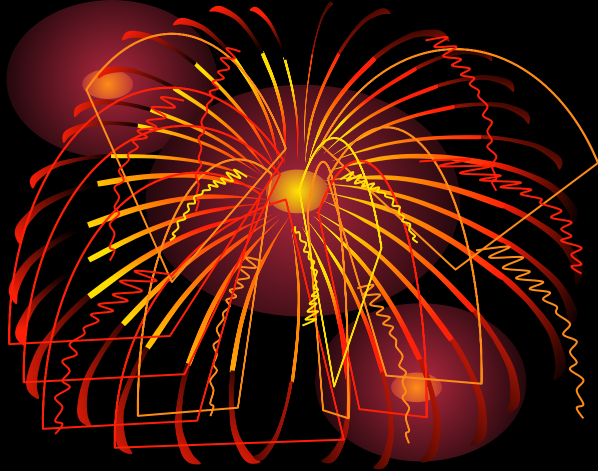
Sample Problem 17.1

Calculating the Specific Heat of a Metal

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



17.1



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

Knowns

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

Unknown

- $C_{Cu} = ?\ J/(g{\cdot}°C)$

17.1

Calculate ***Solve for the unknown.***

Use the known values and the definition of specific heat, $C = \frac{q}{m \times \Delta T}$, to calculate the unknown value C_{Cu}.

$$C_{Cu} = \frac{q}{m \times \Delta T} = \frac{849\ \text{J}}{95.4\ \text{g} \times 23.0°\text{C}} = 0.387\ \text{J/(g·°C)}$$

for Sample Problem 17.1

4. How much heat is required to raise the temperature of 250.0 g of mercury 52°C?

Practice

- 1) Calculate the joules of energy to heat 454 g of water from 5.4°C to 98.6°C.
- 2) What quantity of energy in joules is required to heat 1.3 g of iron from 25°C to 46°C?

Practice

- 3) A 2.6 g sample requires 15.6 J of energy to change its temp from 21°C to 34°C. What metal is it?
- 4) A sample of aluminum requires 3.1 J of energy to change its temp from 19°C to 37°C. What is the mass?

17.1 Section Quiz.

1. The energy released when a piece of wood is burned has been stored in the wood as

   a) sunlight.

   b) heat.

   c) calories.

   d) chemical potential energy.

17.1 Section Quiz.

2. Which of the following statements about heat is false?

   a) Heat is the same as temperature.

   b) Heat always flows from warmer objects to cooler objects.

   c) Adding heat can cause an increase in the temperature of an object.

   d) Heat cannot be specifically detected by senses or instruments.

17.1 Section Quiz.

3. Choose the correct words for the spaces: In an endothermic process, the system __________ heat when heat is __________ its surroundings, so the surroundings ________________.

a) gains, absorbed from, cool down.

b) loses, released to, heat up.

c) gains, absorbed from, heat up.

d) loses, released to, cool down.

17.1 Section Quiz.

4. Which of the relationships listed below can be used to convert between the two units used to measure heat transfer?

a) 1 g = 1°C

b) 1 cal = 4.184 J

c) 1°C = 1 cal

d) 1 g = 4.184 J

17.1 Section Quiz.

5. Assuming that two samples of different materials have equal mass, the one that becomes hotter from a given amount of heat is the one that

   a) has the higher specific heat capacity.

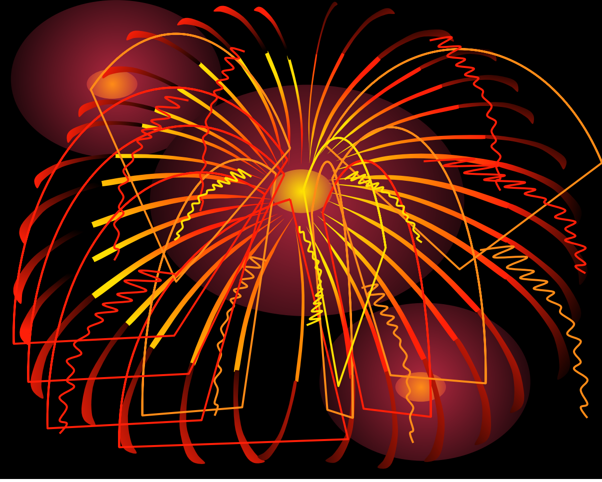
   b) has the higher molecular mass.

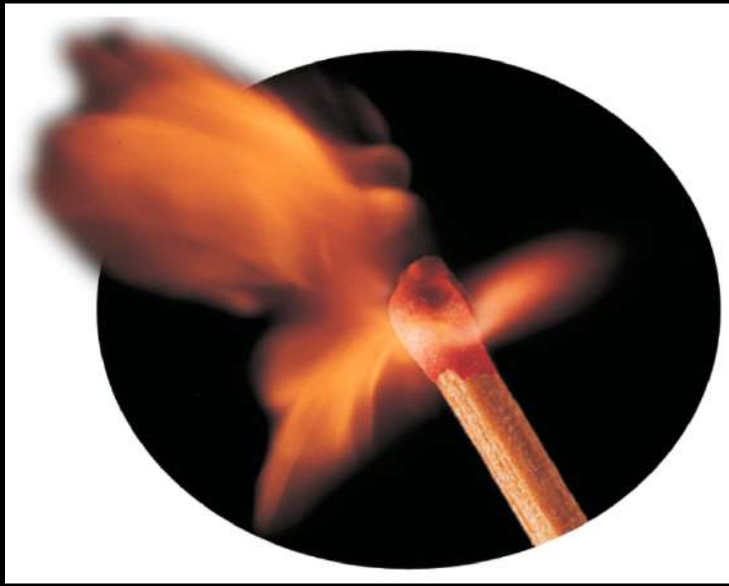
   c) has the lower specific heat capacity.

   d) has the higher density.

17.2 Measuring and Expressing Enthalpy Changes

- A burning match releases heat to its surroundings in all directions. How much heat does this exothermic reaction release? You will learn to measure heat flow in chemical and physical processes by applying the concept of specific heat.





Calorimetry

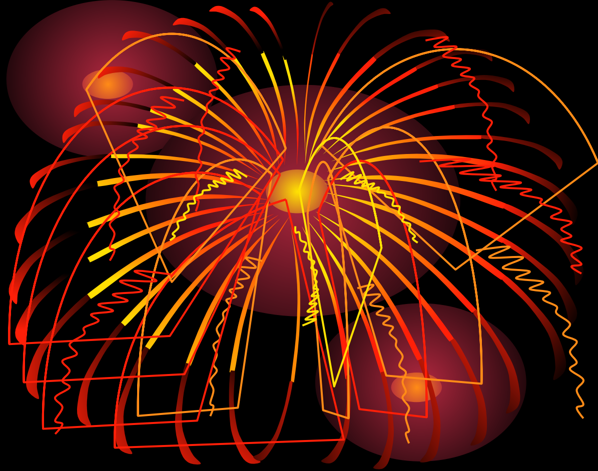
- **Calorimetry** - precise measurement of the heat flow into or out of a system for chemical and physical processes.

Calorimetry

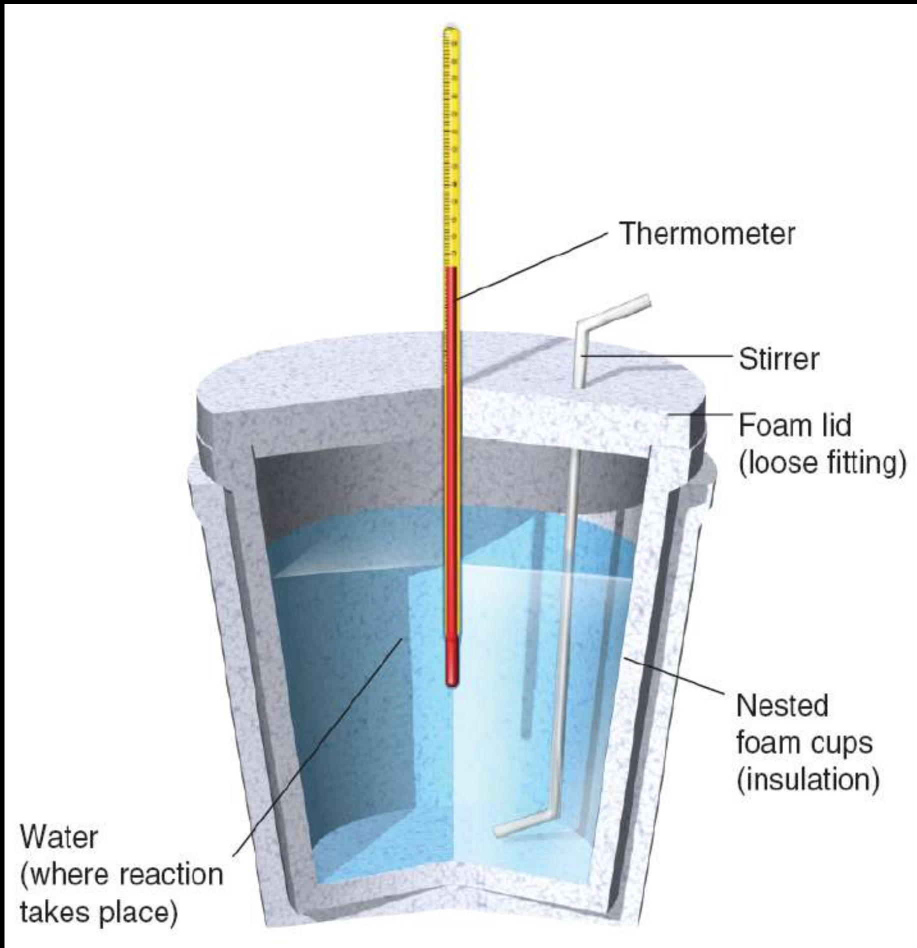
- Heat released by the system is equal to the heat absorbed by its surroundings

$$q_{sys} = \Delta H = -q_{surr} = -m \times C \times \Delta T$$

Calorimetry



- The insulated device used to measure the absorption or release of heat in chemical or physical processes is called a **calorimeter.**

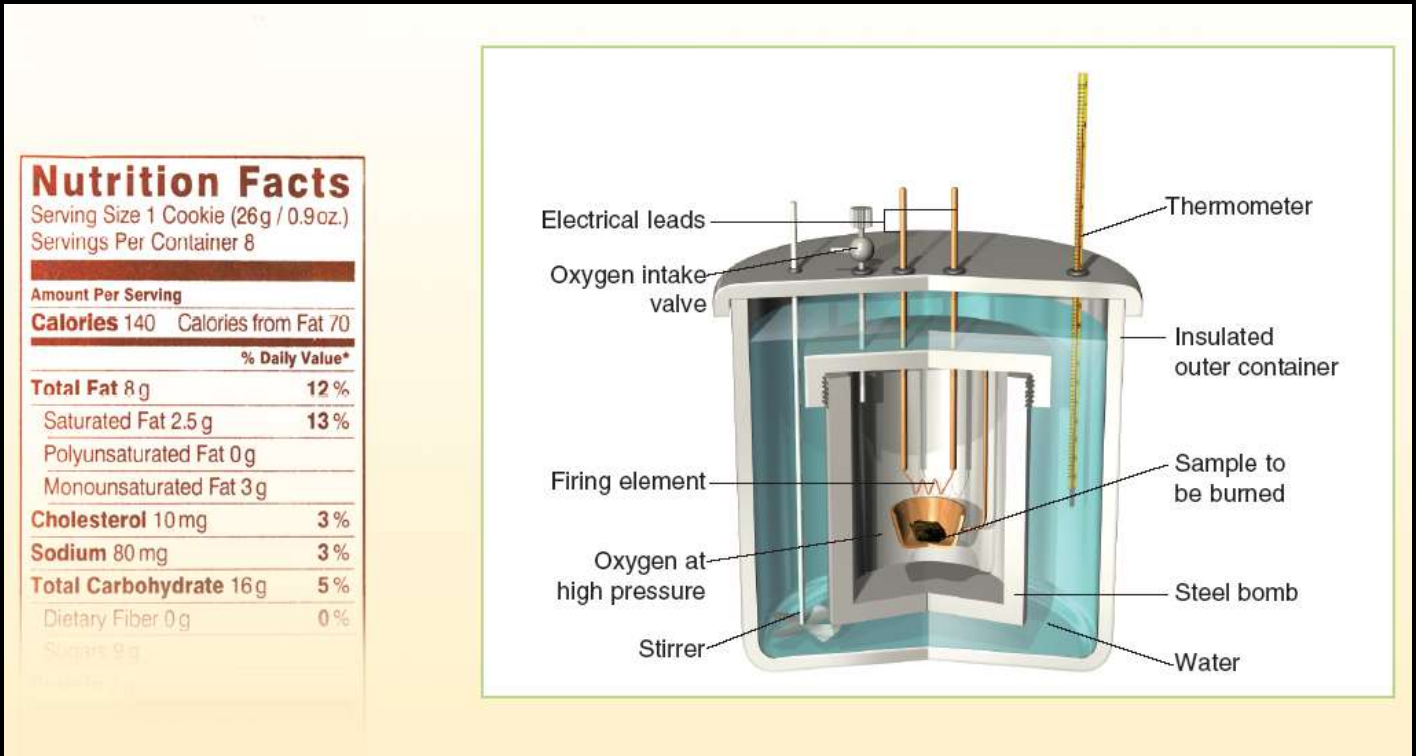


Calorimetry

- The heat content of a system at constant pressure is called the **enthalpy** (*H*) of the system.

Constant-Volume Calorimeters

- Calorimetry experiments can be performed at a constant volume using a bomb calorimeter.



Sample Problem 17.2

Enthalpy Change in a Calorimetry Experiment

When 25.0 mL of water containing 0.025 mol HCl at 25.0°C is added to 25.0 mL of water containing 0.025 mol NaOH at 25.0°C in a foam cup calorimeter, a reaction occurs. Calculate the enthalpy change in kJ) during this reaction if the highest temperature observed is 32.0°C. Assume the densities of the solutions are 1.00 g/mL.

17.2

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

Knowns

- $C_{water} = 4.18\ J/(g{\cdot}°C)$
- $V_{final} = V_{HCl} + V_{NaOH}$
  $= 25.0\ mL + 25.0\ mL = 50.0\ mL$
- $T_i = 25.0°C$
- $T_f = 32.0°C$
- $Density_{solution} = 1.00\ g/mL$

Unknown

- $\Delta H = ?\ kJ$

Use dimensional analysis to determine the mass of the water. You must also calculate ΔT. Use $\Delta H = -q_{surr} = -m \times C \times \Delta T$ to solve for ΔH.

Calculate ***Solve for the unknown.***

First, calculate the total mass of the water.

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

Now calculate ΔT.

$$\Delta T = T_\text{f} - T_\text{i} = 32.0°\text{C} - 25.0°\text{C} = 7.0°\text{C}$$

Use the values for m, C_water, and ΔT to calculate ΔH.

$$\Delta H = -m \times C \times \Delta T = -(50.0\ \cancel{\text{g}})(4.18\ \text{J}/(\cancel{\text{g}} \times \cancel{°\text{C}}))(7.0\cancel{°\text{C}})$$

$$= -1463\ \text{J} = -1.5 \times 10^3\ \text{J} = -1.5\ \text{kJ}$$

for Sample Problem 17.2

13. 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?

Thermochemical Equations

- In a chemical equation, the enthalpy change for the reaction can be written as either a reactant or a product.

Thermochemical Equations

- A chemical equation that includes the enthalpy change is called a **thermochemical equation.**

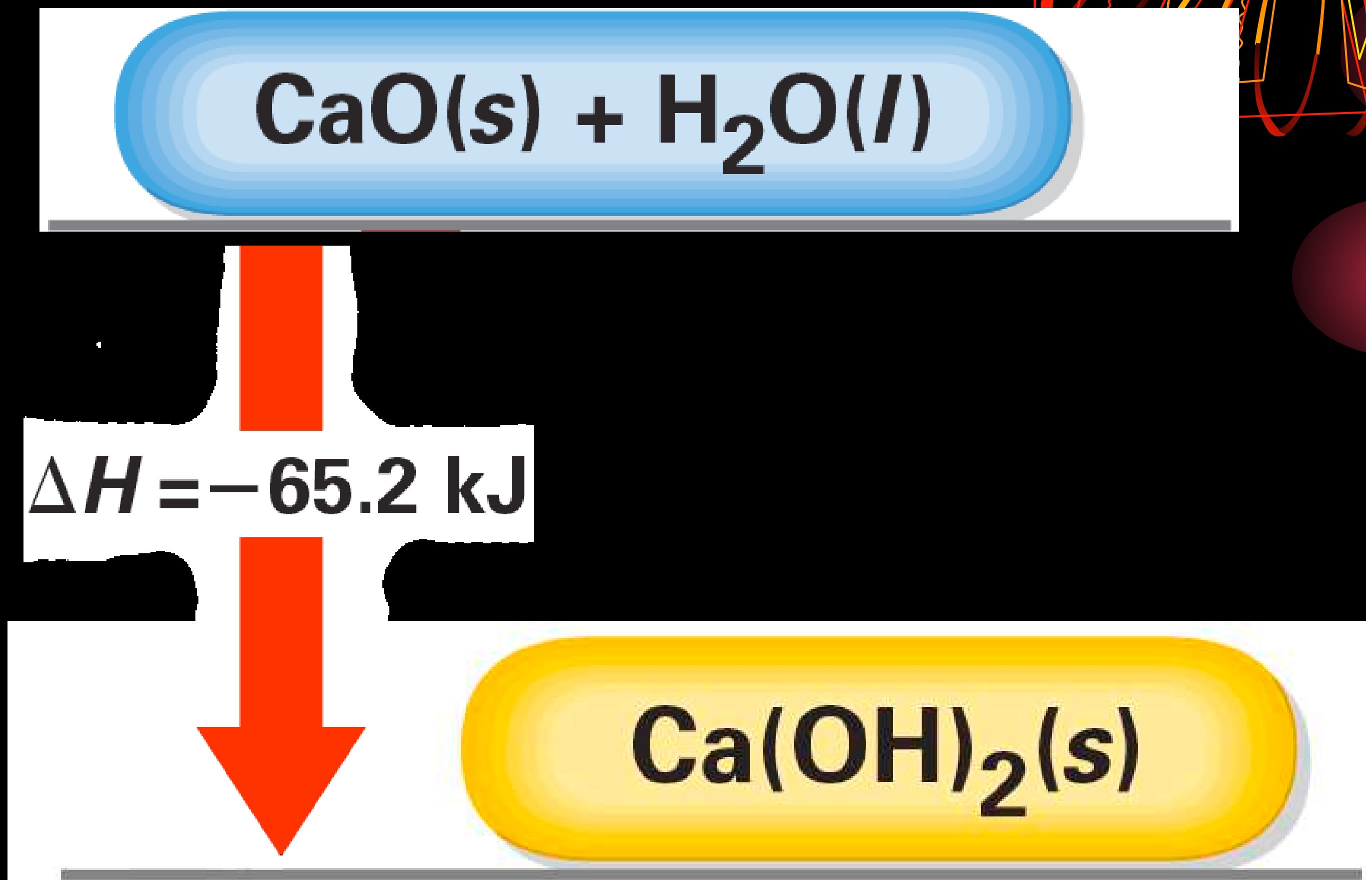
$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(s) + 65.2\ kJ$$

Thermochemical Equations

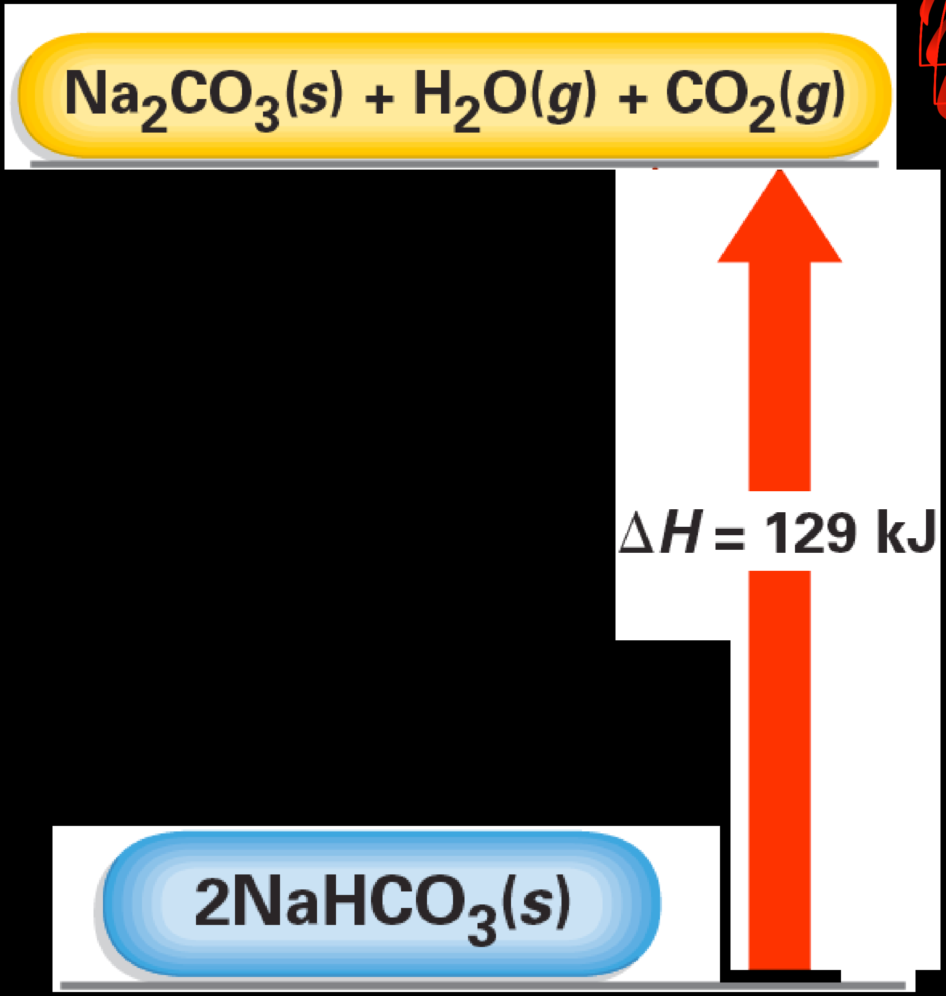
- The **heat of reaction** is the enthalpy change for the chemical equation exactly as it is written.

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(s) \qquad \Delta H = -65.2\ \text{kJ}$$

Exothermic



Endothermic



Sample Problem 17.3

Using the Heat of Reaction to Calculate Enthalpy Change

Using the thermochemical equation in Figure 17.7b on page 515, calculate the amount of heat (in kJ) required to decompose 2.24 mol $NaHCO_3(s)$.

17.3

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

Knowns

- 2.24 mol $NaHCO_3(s)$ decomposes
- $\Delta H = 129$ kJ (for 2 mol $NaHCO_3$)

Unknown

- $\Delta H = ?$ kJ

Use the thermochemical equation,

$$2NaHCO_3(s) + 129\text{ kJ} \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g),$$

to write a conversion factor relating kilojoules of heat and moles of $NaHCO_3$. Then use the conversion factor to determine ΔH for 2.24 mol $NaHCO_3$.

17.3

Calculate ***Solve for the unknown.***

The thermochemical equation indicates that 129 kJ are needed to decompose 2 mol $NaHCO_3(s)$. Use this relationship to write the following conversion factor.

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

Using dimensional analysis, solve for ΔH.

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

for Sample Problem 17.3

15. 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 Fe_2O_3 reacts with an excess of CO?

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g) + 26.3\ kJ$$

Practice

- The reaction for heat packs to treat sports injury is:
- $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ $\Delta H = -1652$ kJ
- How much heat is released when 1.00 g of Fe is reacted?

Thermochemical Equations

- The **heat of combustion** is the heat of reaction for the complete burning of one mole of a substance.

Table 17.2

Heats of Combustion at 25°C		
Substance	**Formula**	**ΔH (kJ/mol)**
Hydrogen	$H_2(g)$	−286
Carbon	$C(s)$, graphite	−394
Methane	$CH_4(g)$	−890
Acetylene	$C_2H_2(g)$	−1300
Ethanol	$C_2H_5OH(l)$	−1368
Propane	$C_3H_8(g)$	−2220
Glucose	$C_6H_{12}O_6(s)$	−2808
Octane	$C_8H_{18}(l)$	−5471
Sucrose	$C_{12}H_{22}O_{11}(s)$	−5645

17.2 Section Quiz.

- 1. The change in temperature recorded by the thermometer in a calorimeter is a measurement of

  a) the enthalpy change of the reaction in the calorimeter.

  b) the specific heat of each compound in a calorimeter.

  c) the physical states of the reactants in a colorimeter.

  d) the heat of combustion for one substance in a calorimeter.

17.2 Section Quiz.

- 2. For the reaction $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$, $\Delta H = -65.2$ kJ. This means that 65.2 kJ of heat is ____________ during the process.

  a) absorbed

  b) destroyed

  c) changed to mass

  d) released

17.2 Section Quiz.

- 3. How much heat is absorbed by 325 g of water if its temperature changes from 17.0°C to 43.5°C? The specific heat of water is 4.18 J/g°C.

  a) 2.00 kJ

  b) 3.60 kJ

  c) 36.0 kJ

  d) 360 kJ

17.2 Section Quiz.

- 4. Which of the following is a thermochemical equation for an endothermic reaction?

  a) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) + 890\ kJ$

  b) $241.8\ kJ + 2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

  c) $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)\ \ 65.2\ kJ$

  d) $2NaHCO_3(s)\ 129\ kJ \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

17.2 Section Quiz.

- 5. Oxygen is necessary for releasing energy from glucose in organisms. How many kJ of heat are produced when 2.24 mol glucose reacts with an excess of oxygen?
  - $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g) + 2808\ kJ/mol$

    a) 4.66 kJ

    b) 9.31 kJ

    c) 1048 kJ

    d) 6290 kJ

17.3 Heat in Changes of State

- During a race, an athlete can burn a lot of calories that either do work or are released as heat. This section will help you to understand how the evaporation of sweat from your skin helps to rid your body of excess heat.

Heats of Fusion and Solidification

- The **molar heat of fusion** (ΔH_{fus}) is the heat absorbed by one mole of a solid substance as it melts to a liquid at a constant temperature.
- The **molar heat of solidification** (ΔH_{solid}) is the heat lost when one mole of a liquid solidifies at a constant temperature.

Heats of Fusion and Solidification

- The quantity of heat absorbed by a melting solid is exactly the same as the quantity of heat released when the liquid solidifies; that is, $\Delta H_{fus} = -\Delta H_{solid}$.

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?

17.4

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

Knowns

- Initial and final temperatures are 0°C
- $\Delta H_{fus} = 6.01$ kJ/mol
- $\Delta H = 2.25$ kJ

Unknown

- $m_{ice} = ?$ g

Use the thermochemical equation

$$H_2O(s) + 6.01 \text{ kJ} \longrightarrow H_2O(l)$$

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.

17.4

Calculate ***Solve for the unknown.***

Express ΔH_{fus} and the molar mass of ice as conversion factors.

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

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

$$m_{ice} = 2.25 \cancel{\text{kJ}} \times \frac{1 \cancel{\text{mol ice}}}{6.01 \cancel{\text{kJ}}} \times \frac{18.0 \text{ g ice}}{1 \cancel{\text{mol ice}}}$$
$$= 6.74 \text{ g ice}$$

for Sample Problem 17.4

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.

Practice

- 1) Calculate the energy released when 15.5 g of ice freezes at 0°C.
- Calculate the energy required to melt 12.5 g of ice at 0°C and change it to water at 25°C. (Specific heat capacity is 4.18 J/g°C)

Heats of Vaporization and Condensation

- The amount of heat necessary to vaporize one mole of a given liquid is called its **molar heat of vaporization** (ΔH_{vap}).
- 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}).

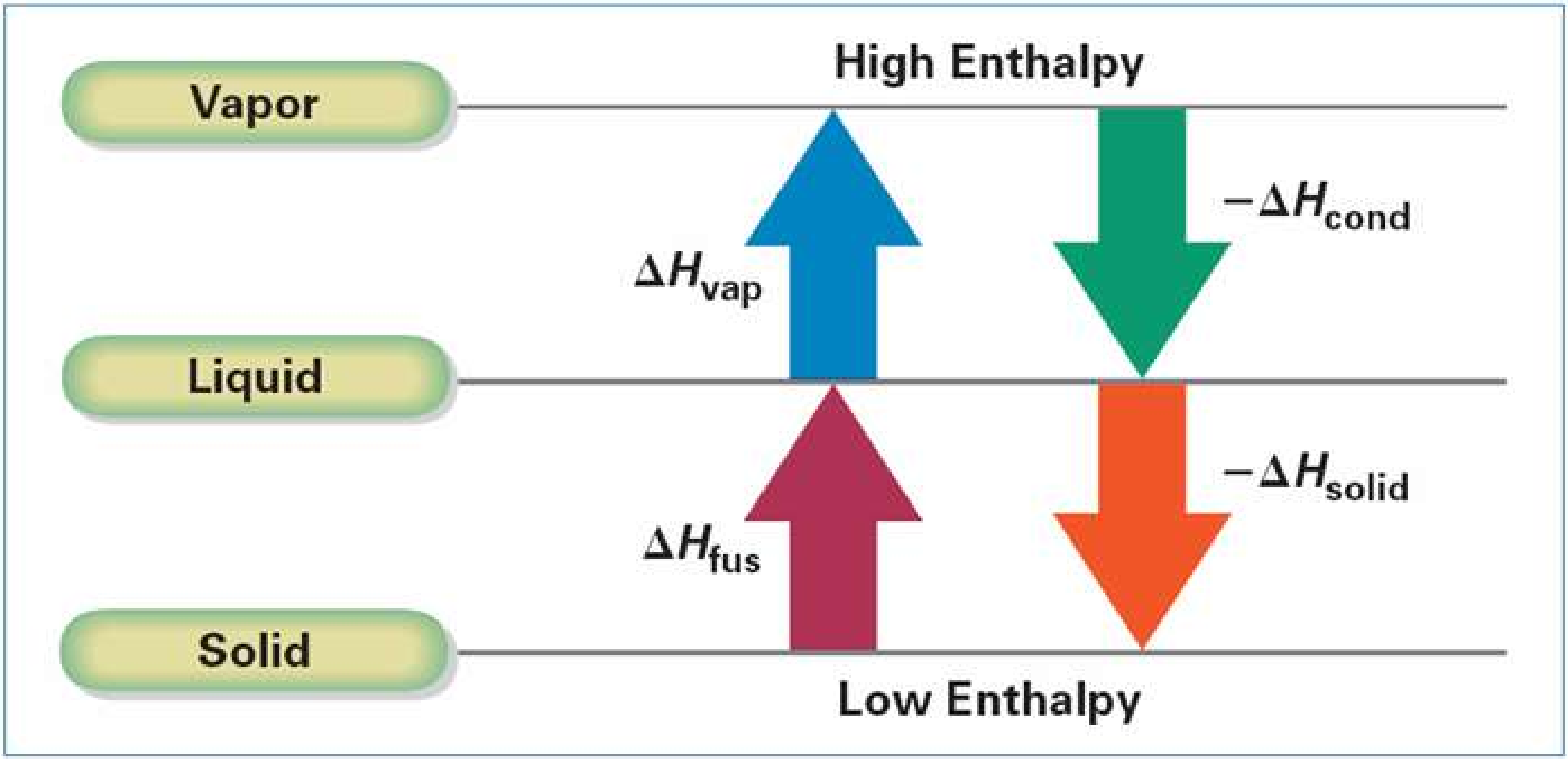
Heats of Vaporization and Condensation

- 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, $\Delta H_{vap} = -\Delta H_{cond}$.

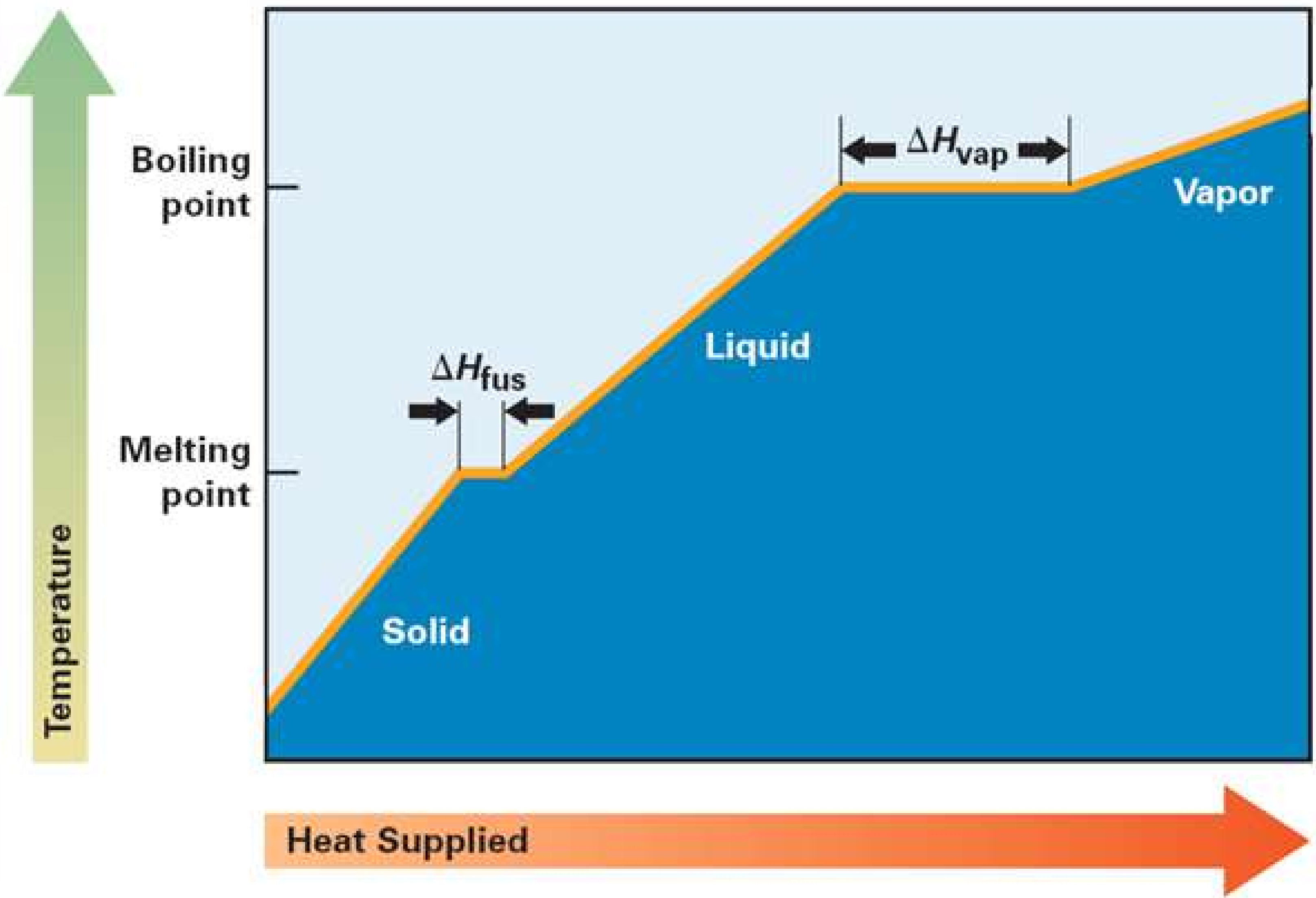
Table 17.3

Heats of Physical Change		
Substance	**ΔH_{fus} (kJ/mol)**	**ΔH_{vap} (kJ/mol)**
Ammonia (NH_3)	5.65	23.4
Ethanol (C_2H_5OH)	4.60	43.5
Hydrogen (H_2)	0.12	0.90
Methanol (CH_3OH)	3.16	35.3
Oxygen (O_2)	0.44	6.82
Water (H_2O)	6.01	40.7

- Enthalpy changes accompany changes in state.



Heating Curve for Water



Sample Problem 17.5

Using the Heat of Vaporization in Phase-Change Calculations

How much heat (in kJ) is absorbed when 24.8 g $H_2O(l)$ at 100°C and 101.3 kPa is converted to steam at 100°C?

17.5

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
- ΔH_{vap} = 40.7 kJ/mol

Unknown

- ΔH = ? kJ

Refer to the following thermochemical equation.

$$H_2O(l) + 40.7 \text{ kJ/mol} \longrightarrow H_2O(g)$$

ΔH_{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 ΔH_{vap}.

17.5

Calculate ***Solve for the unknown.***

The required conversion factors come from ΔH_{vap} and the molar mass of water.

$$\frac{1\ \text{mol}\ H_2O(l)}{18.0\ \text{g}\ H_2O(l)} \quad \text{and} \quad \frac{40.7\ \text{kJ}}{1\ \text{mol}\ H_2O(l)}$$

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

$$\Delta H = 24.8\ \cancel{\text{g}\ H_2O(l)} \times \frac{1\ \cancel{\text{mol}\ H_2O(l)}}{18.0\ \cancel{\text{g}\ H_2O(l)}} \times \frac{40.7\ \text{kJ}}{1\ \cancel{\text{mol}\ H_2O(l)}}$$

$$= 56.1\ \text{kJ}$$

for Sample Problem 17.5

24. How many kilojoules of heat are absorbed when 0.46 g of chloroethane (C_2H_5Cl, bp 12.3°C) vaporizes at its normal boiling point? The molar heat of vaporization of chloroethane is 26.4 kJ/mol.

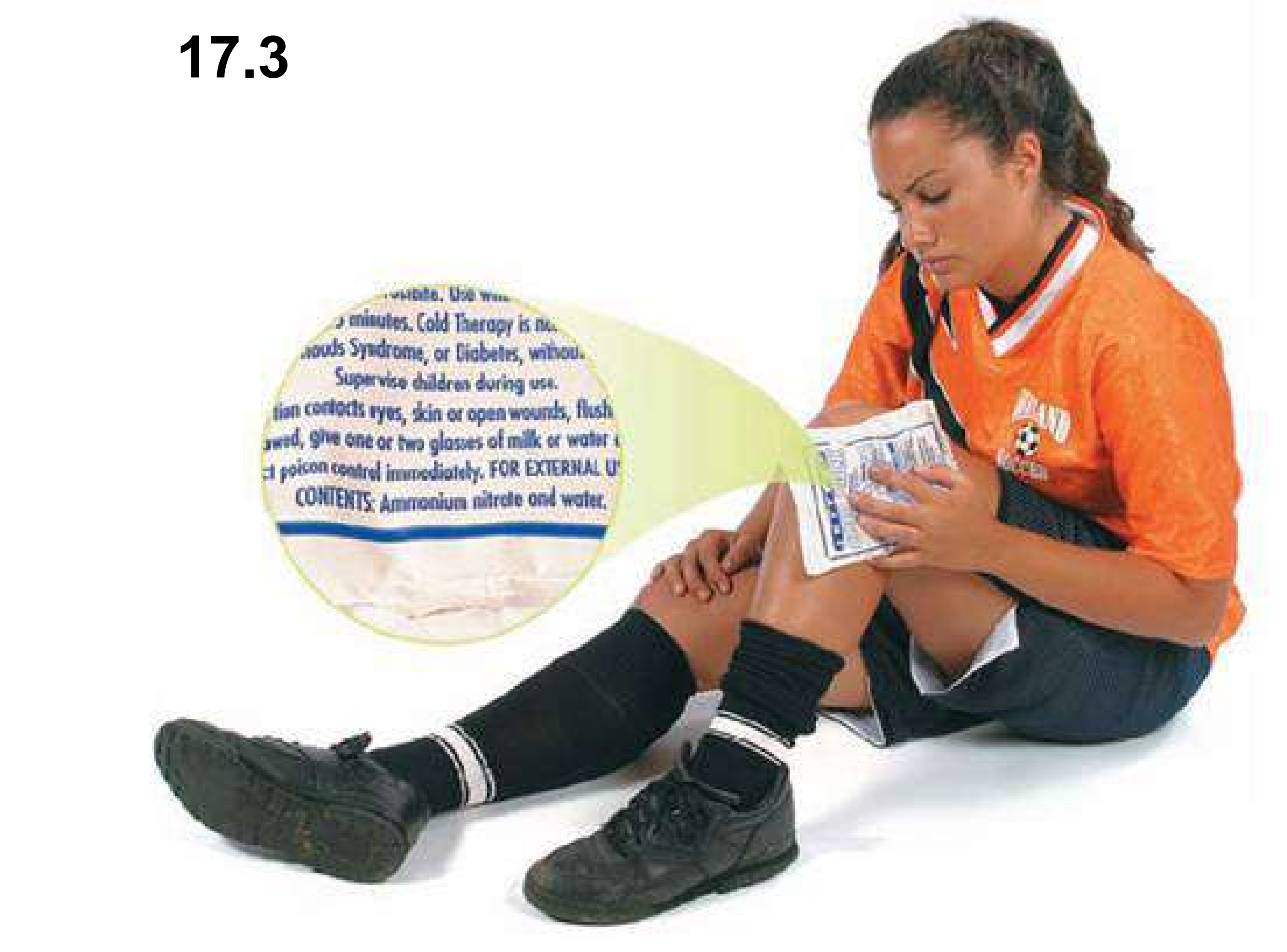
Practice

- 1) Calculate the energy required to vaporize 35 g of water at 100°C
- 2) Calculate the energy to melt 15 g of ice at 0°C, heat it to 100°C and vaporize it to steam at 100°C. The specific heat of water is 4.18 J/g°C.

Heat of Solution

- 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_{soln}).**

17.3



- When ammonium nitrate crystals and water mix inside the cold pack, heat is absorbed as the crystals dissolve.

Sample Problem 17.6

Calculating the Enthalpy Change in Solution Formation

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

17.6

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

Knowns

- $\Delta H_{soln} = -445.1 \text{ kJ/mol}$
- 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 (ΔH).

$$NaOH(s) \xrightarrow{H_2O(l)} Na^+(aq) + OH^-(aq) + 445.1 \text{ kJ/mol}$$

17.6

Calculate ***Solve for the unknown.***

Multiply the number of moles of NaOH by ΔH_{soln}.

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

for Sample Problem 17.6

26. How many moles of $NH_4NO_3(s)$ must be dissolved in water so that 88.0 kJ of heat is absorbed from the water?

17.3 Section Quiz.

1. The molar heat of condensation of a substance is the same, in magnitude, as its molar heat of
   a) formation.
   b) fusion.
   c) solidification.
   d) vaporization.

17.3 Section Quiz

2. The heat of condensation of ethanol (C_2H_5OH) is –43.5 kJ/mol. As C_2H_5OH condenses, the temperature of the surroundings

a) stays the same.

b) may increase or decrease.

c) increases.

d) decreases.

17.3 Section Quiz

3. Calculate the amount of heat absorbed to liquefy 15.0 g of methanol (CH_3OH) at its melting point. The molar heat of fusion for methanol is 3.16 kJ/mol.

a) 1.48 kJ

b) 47.4 kJ

c) 1.52×10^3 kJ

d) 4.75 kJ

17.3 Section Quiz

4. How much heat (in kJ) is released when 50 g of $NH_4NO_3(s)$, 0.510 moles, are dissolved in water? Δ S_{soln} = –25.7 kJ/mol

   a) 12.85 kJ

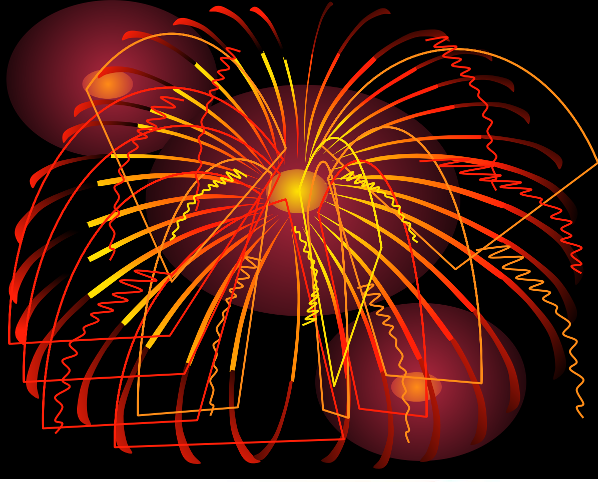
   b) 13.1 kJ

   c) 25.7 kJ

   d) 1285 kJ

17.4 Calculating Heats of Reaction

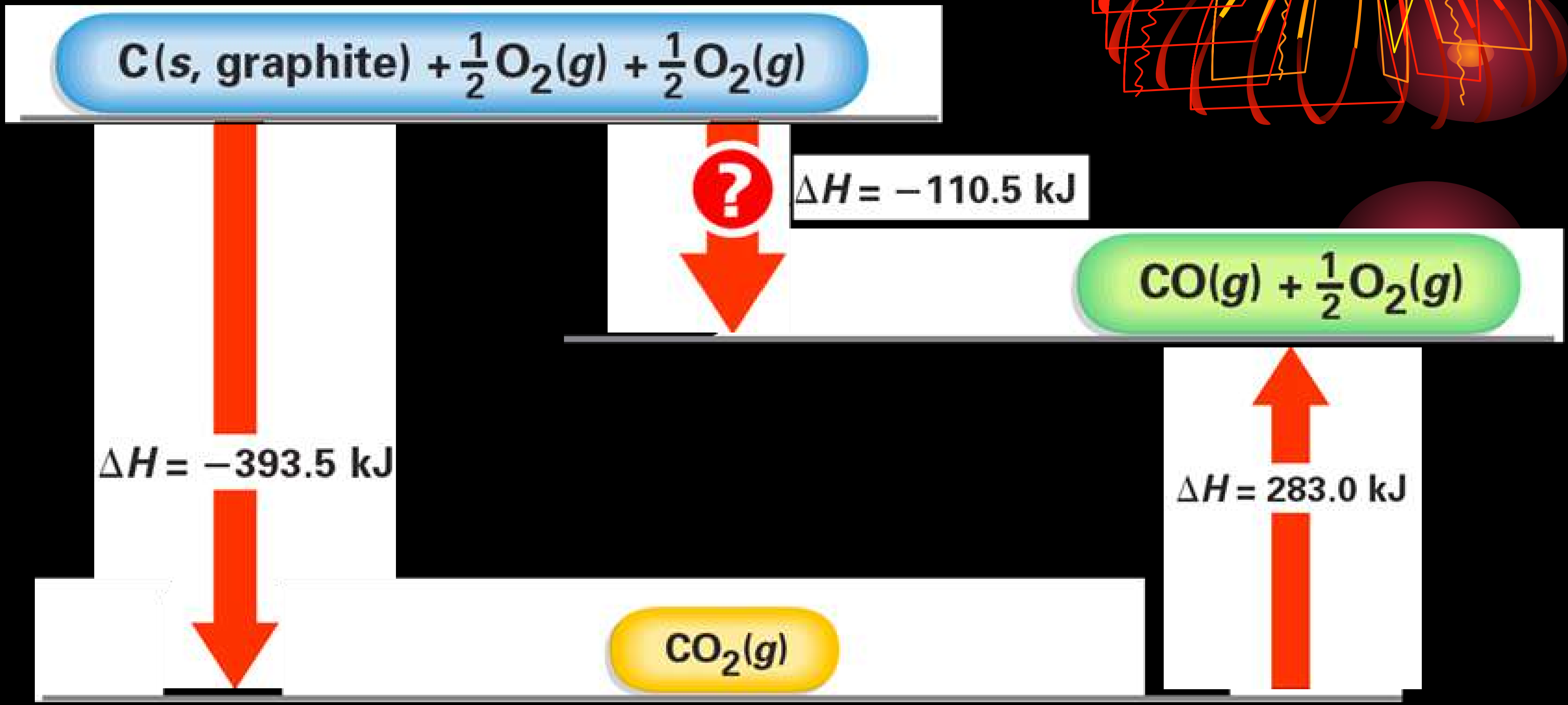
- Emeralds are composed of the elements chromium, aluminum, silicon, oxygen, and beryllium. What if you wanted to determine the heat of reaction without actually breaking the gems down to their component elements? You will see how you can calculate heats of reaction from known thermochemical equations and enthalpy data.



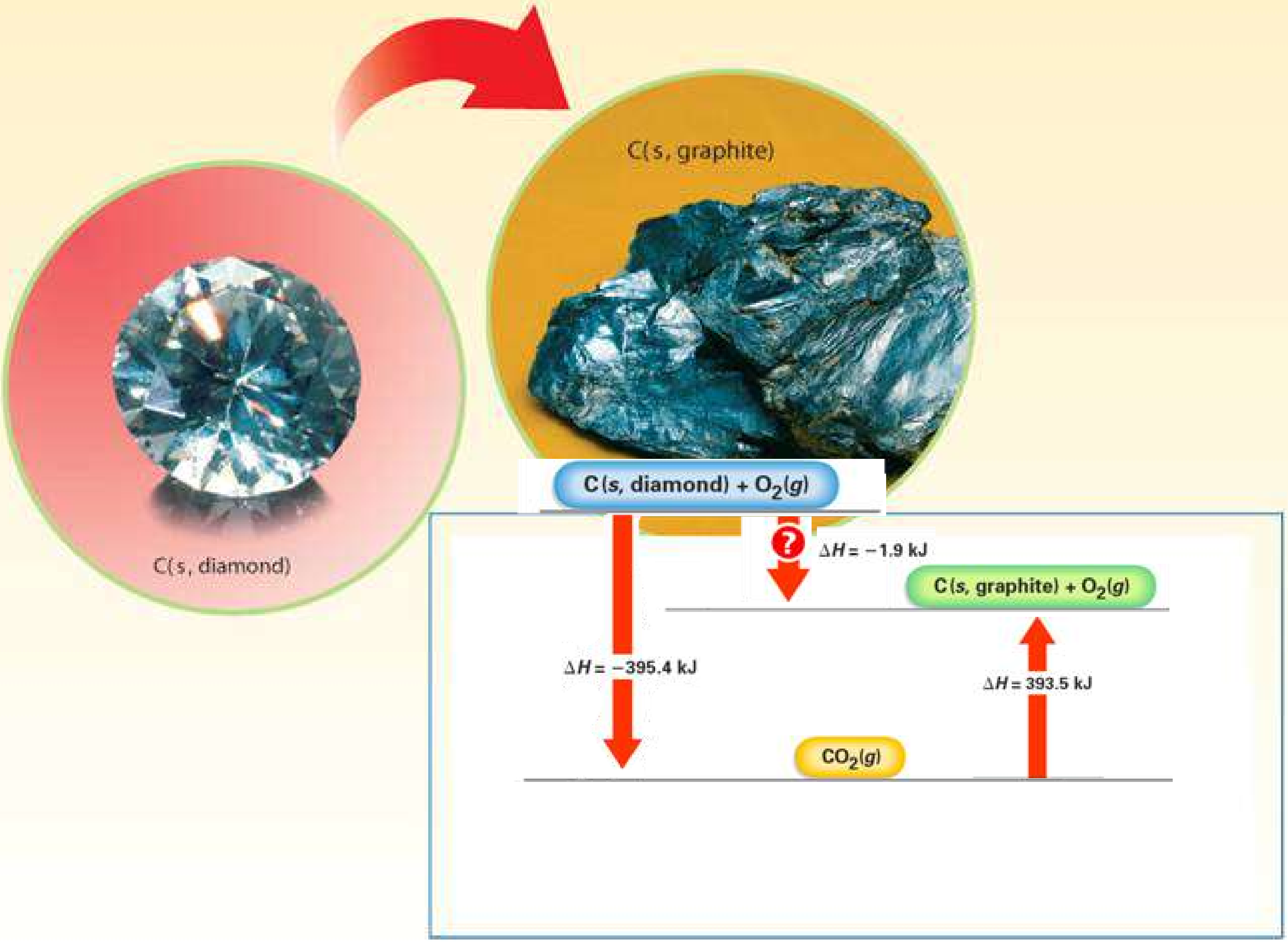
Hess's Law

- Hess's law allows you to determine the heat of reaction indirectly.
  - Hess's law of heat summation **states that 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.**

Hess's Law


C(s, graphite) + ½O₂(g) + ½O₂(g)
?
ΔH = −110.5 kJ
CO(g) + ½O₂(g)
ΔH = −393.5 kJ
ΔH = 283.0 kJ
CO₂(g)

Hess's Law



C(s, diamond)
C(s, graphite)
C (s, diamond) + O2(g)
ΔH = −1.9 kJ
C (s, graphite) + O2(g)
ΔH = −395.4 kJ
ΔH = 393.5 kJ
CO2(g)

Standard Heats of Formation

- Standard Heats of Formation
  - For a reaction that occurs at standard conditions, you can calculate the heat of reaction by using standard heats of formation.

Standard Heats of Formation

- The **standard heat of formation** (ΔH_f^0) 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.

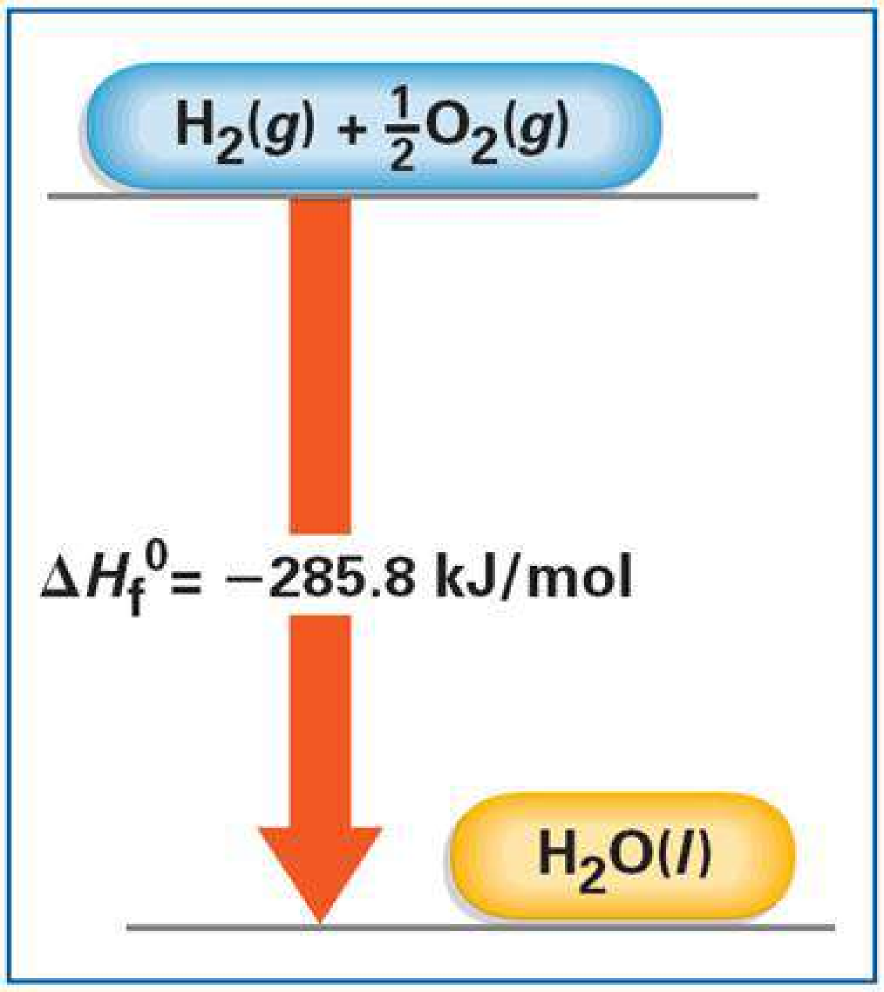
$$\Delta H^0 = \Delta H_f^0(\text{products}) - \Delta H_f^0(\text{reactants})$$

Table 17.4

Standard Heats of Formation (ΔH_f^0) at 25°C and 101.3 kPa			
Substance	ΔH_f^0 (kJ/mol)	Substance	ΔH_f^0 (kJ/mol)
$Al_2O_3(s)$	−1676.0	$H_2O_2(l)$	−187.8
$Br_2(g)$	30.91	$I_2(g)$	62.4
$Br_2(l)$	0.0	$I_2(s)$	0.0
C(*s*, diamond)	1.9	$N_2(g)$	0.0
C(*s*, graphite)	0.0	$NH_3(g)$	−46.19
$CH_4(g)$	−74.86	$NO(g)$	90.37
$CO(g)$	−110.5	$NO_2(g)$	33.85
$CO_2(g)$	−393.5	$NaCl(s)$	−411.2
$CaCO_3(s)$	−1207.0	$O_2(g)$	0.0
$CaO(s)$	−635.1	$O_3(g)$	142.0
$Cl_2(g)$	0.0	P(*s*, white)	0.0
$Fe(s)$	0.0	P(*s*, red)	−18.4
$Fe_2O_3(s)$	−822.1	S(*s*, rhombic)	0.0
$H_2(g)$	0.0	S(*s*, monoclinic)	0.30
$H_2O(g)$	−241.8	$SO_2(g)$	−296.8
$H_2O(l)$	−285.8	$SO_3(g)$	−395.7

Standard Heats of Formation

- **The Standard Heat of Formation of Water**

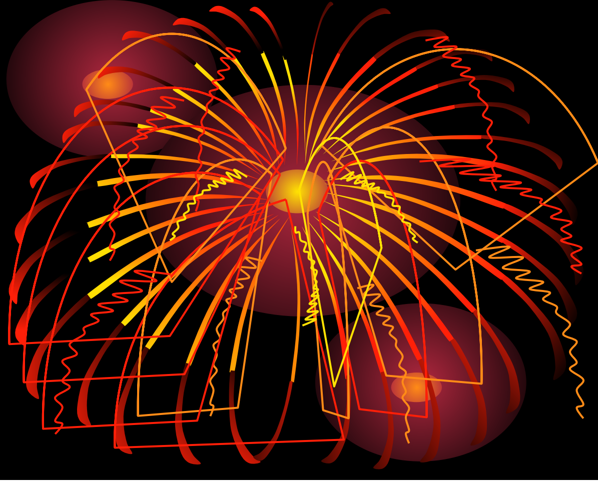


Sample Problem 17.7

Calculating the Standard Heat of Reaction

What is the standard heat of reaction (ΔH^0) for the reaction of $CO(g)$ with $O_2(g)$ to form $CO_2(g)$?

17.7



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

Knowns

(from Table 17.4)

- $\Delta H_f^0 O_2(g) = 0$ kJ/mol (free element)
- $\Delta H_f^0 CO(g) = -110.5$ kJ/mol
- $\Delta H_f^0 CO_2(g) = -393.5$ kJ/mol

Unknown

- $\Delta H^0 = ?$ kJ

Balance the equation of the reaction of $CO(g)$ with $O_2(g)$ to form $CO_2(g)$. Then determine ΔH^0 using the standard heats of formation of the reactants and products.

Calculate ***Solve for the unknown.***

First, write the balanced equation.

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$

Next, find and add the ΔH_f^0 of all of the reactants, taking into account the number of moles of each.

$$\Delta H_f^0(\text{reactants}) = 2\ \cancel{\text{mol CO}(g)} = \frac{-110.5\ \text{kJ}}{1\ \cancel{\text{mol CO}(g)}} + 0\ \text{kJ}$$

$$= -221.0\ \text{kJ}$$

Then, find the ΔH_f^0 of the product in a similar way.

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

$$= -787.0\ \text{kJ}$$

Finally, solve for the unknown

$$\Delta H^0 = \Delta H_f^0(\text{products}) - \Delta H_f^0(\text{reactants})$$

$$\Delta H^0 = (-787.0\ \text{kJ}) - (-221.0\ \text{kJ})$$

$$\Delta H^0 = -566.0\ \text{kJ}$$

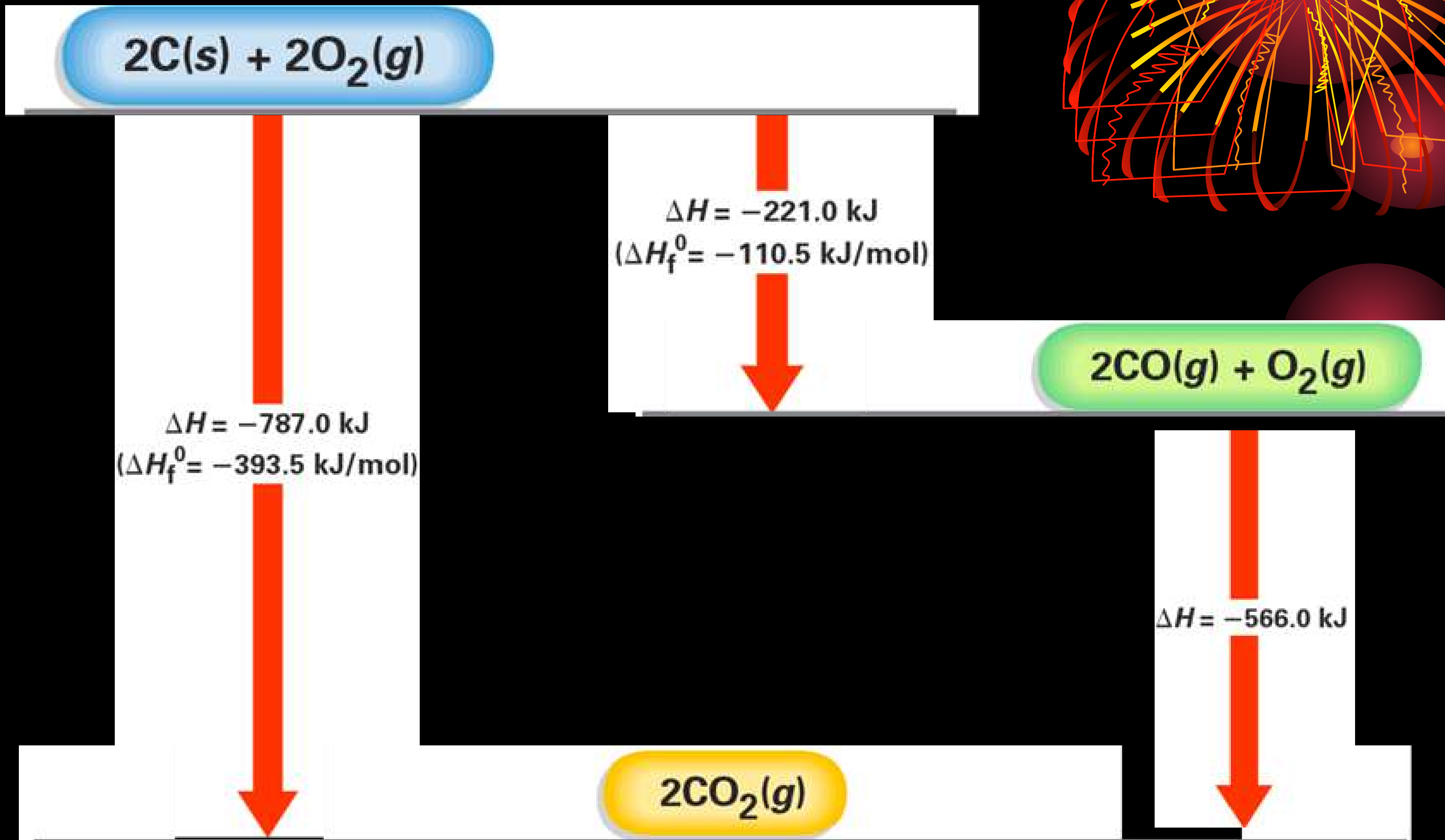
32. Calculate ΔH^0 for the following reactions.

a. $Br_2(g) \longrightarrow Br_2(l)$

b. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

c. $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

Standard Heats of Formation



17.4 Section Quiz.

- 1. According to Hess's law, it is possible to calculate an unknown heat of reaction by using
  a) heats of fusion for each of the compounds in the reaction.
  b) two other reactions with known heats of reaction.
  c) specific heat capacities for each compound in the reaction.
  d) density for each compound in the reaction.

17.4 Section Quiz.

- 2. The heat of formation of $Cl_2(g)$ at 25°C is

  a) the same as that of H_2O at 25°C.

  b) larger than that of Fe(*s*) at 25°C.

  c) undefined.

  d) zero.

17.4 Section Quiz.

- 3. Calculate ΔH^0 for $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$. Standard heats of formation: $NH_3(g) = -45.9$ kJ/mol, $HCl(g) = -92.3$ kJ/mol, $NH_4Cl(s) = -314.4$ kJ/mol

- a) 176.2 kJ

  b) –360.8 kJ

  c) –176.2 kJ

  d) –268 kJ.