



CHEMISTRY

TENTH EDITION

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Chapter 6

Thermochemistry

Chapter 6

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Section 6.1

The Nature of Energy

Energy

- Capacity to do work or to produce heat
- **Law of conservation of energy:** Energy can be converted from one form to another but can be neither created nor destroyed
 - Energy of the universe is constant

Section 6.1

The Nature of Energy

Classification of Energy

- **Potential energy** (PE): Energy due to position or composition
 - Can result from attractive and repulsive forces
- **Kinetic energy** (KE): Energy due to the motion of an object
 - Depends on the mass of the object (m) and its velocity (v)

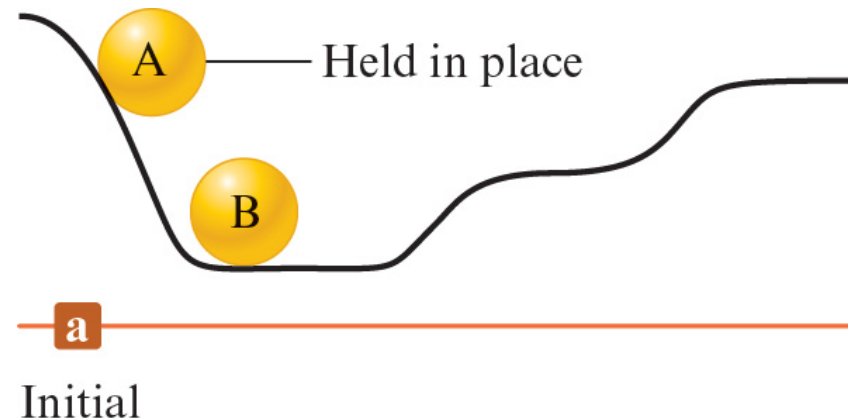
$$\text{KE} = \frac{1}{2}mv^2$$

Section 6.1

The Nature of Energy

Conversion of Energy

- Consider the following image:

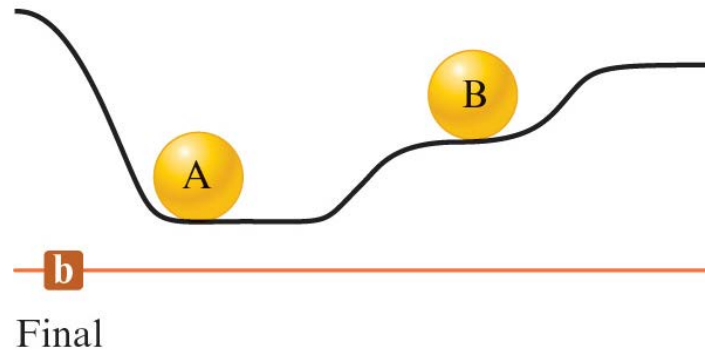


- Because of its higher initial position, ball A has more PE than ball B

Section 6.1

The Nature of Energy

Conversion of Energy (continued 1)



- PE of A is changed to KE as the ball rolls down the hill
- Part of this KE is then transferred to B, causing it to be raised to a higher final position
 - Thus, PE of B has been increased

Section 6.1

The Nature of Energy

Conversion of Energy (continued 2)

- Since the final position of B is lower than the original position of A, some of the energy is still unaccounted for
- Part of the original energy stored as PE in A has been transferred through work to B, thereby increasing B's PE

Section 6.1

The Nature of Energy

Methods of Transferring Energy

- **Heat:** Transfer of energy between two objects due to a temperature difference
 - Temperature is a property that reflects random motion of particles in a substance
- **Work:** Force acting over a distance

Section 6.1

The Nature of Energy

Pathway

- Specific conditions that define the path by which energy is transferred
- Energy change is independent of the pathway
- Work and heat are dependent on the pathway

Section 6.1

The Nature of Energy

Energy as a State Function

- **State function (state property)**: Property that does not depend in any way on the system's past or future
 - Value depends on the characteristics of the present state
 - While transitioning from one state to another, the change in state property is independent of the pathway taken between the two states

Section 6.1

The Nature of Energy

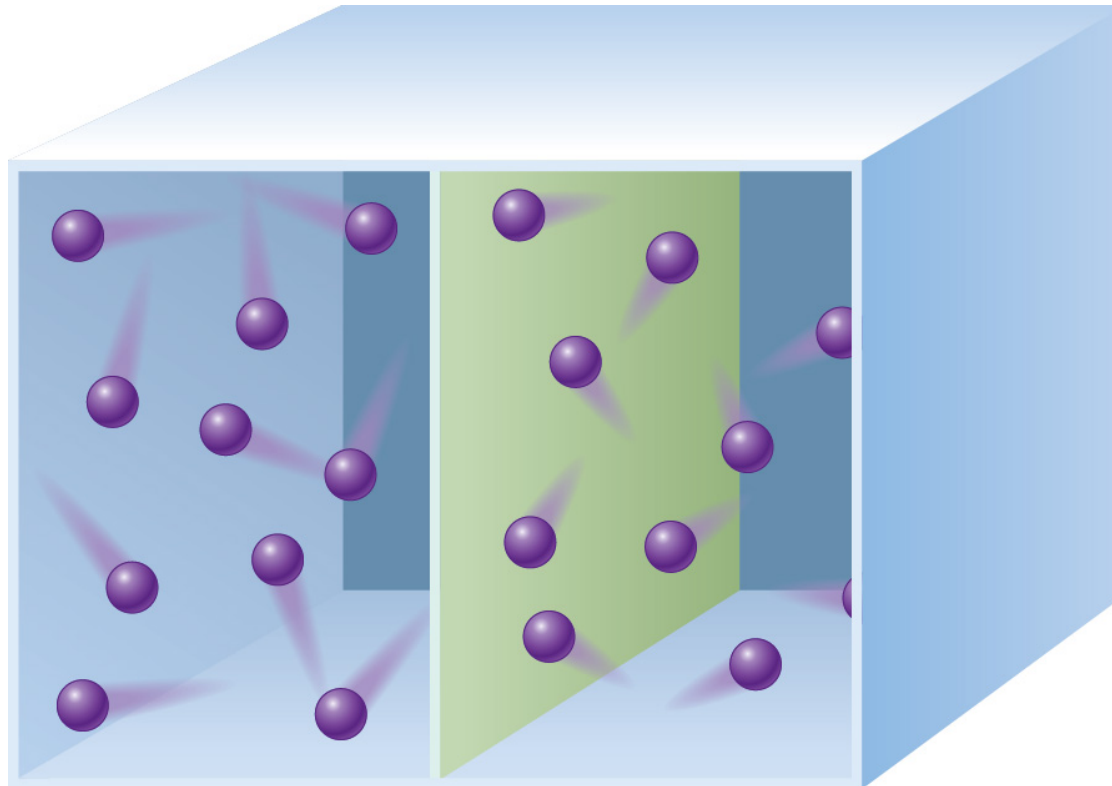
The Process of Heat

- Energy is always transferred from the hot body to the cold body
 - Temperature is a measure of the average KE of particles in solids, liquids, and gases
 - Energy is transferred as the initially faster particles decrease in motion and the initially slower molecules increase in motion
 - Particles will have the same average KE and thus the same temperature

Section 6.1

The Nature of Energy

Figure 6.2 - Particles at Different Temperatures in Adjoining Chambers



Section 6.1

The Nature of Energy

Parts of the Universe

- **System:** Part of the universe on which one wishes to focus his/her attention
 - Example - System can be defined as the reactants and products of a reaction
- **Surroundings:** Include everything else in the universe
 - Example - Surroundings consist of anything else other than the reactants and products

Section 6.1

The Nature of Energy

Types of Reactions

- **Exothermic:** Reaction that results in the evolution of heat
 - Energy flows out of the system
 - Example - Combustion of methane
- **Endothermic:** Reaction that results in the absorption of energy from the surroundings
 - Heat flows into a system
 - Example - Formation of nitric oxide

Section 6.1

The Nature of Energy

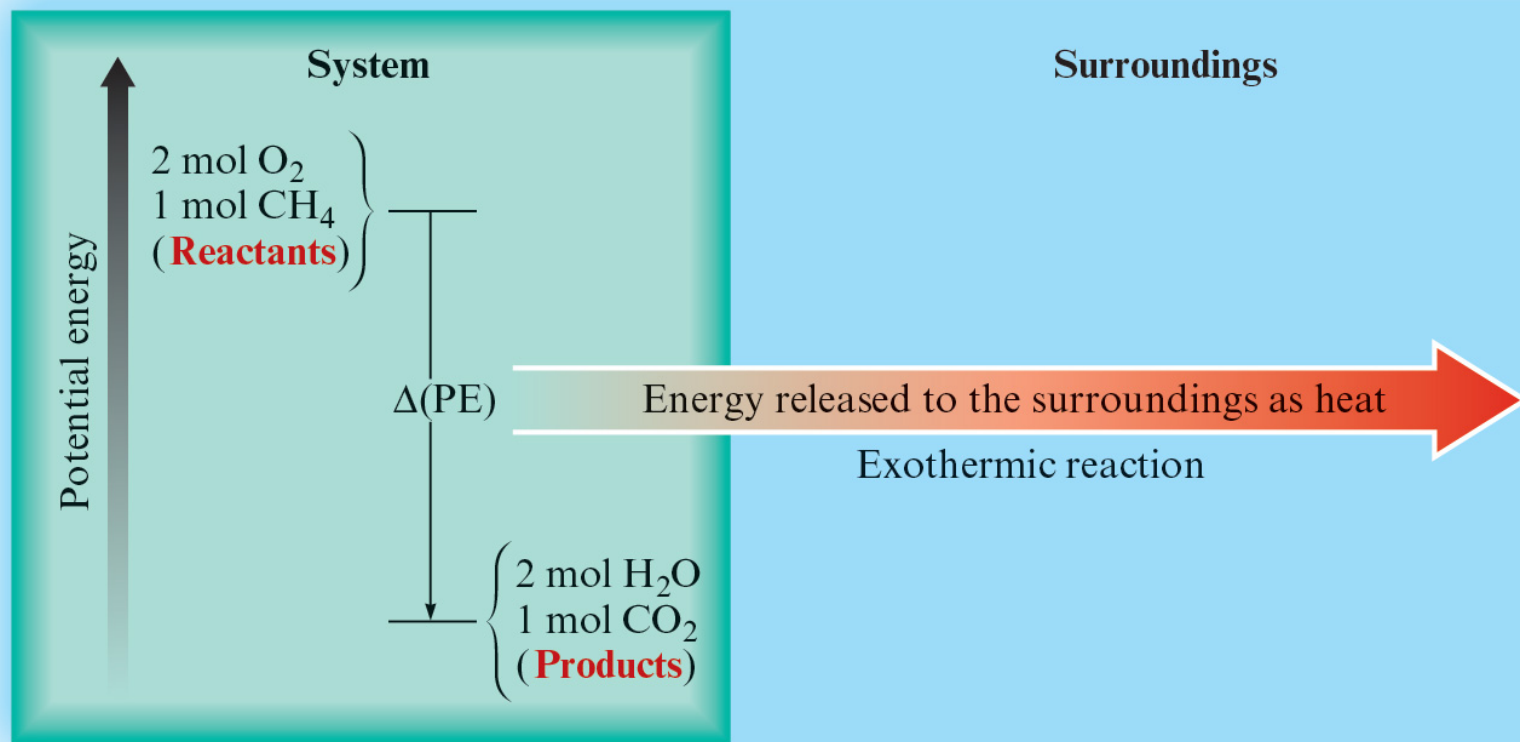
Reaction Mechanism

- Energy gained by the surroundings must be equal to the energy lost by the system
- Exothermic reactions
 - PE stored in chemical bonds is converted to thermal energy via heat
 - Bonds in the products are stronger than those of the reactants
 - Net result - Quantity of energy $\Delta(\text{PE})$ is transferred to the surroundings through heat

Section 6.1

The Nature of Energy

Figure 6.3 - Energy Diagram for the Combustion of Methane, an Exothermic Process



Section 6.1

The Nature of Energy

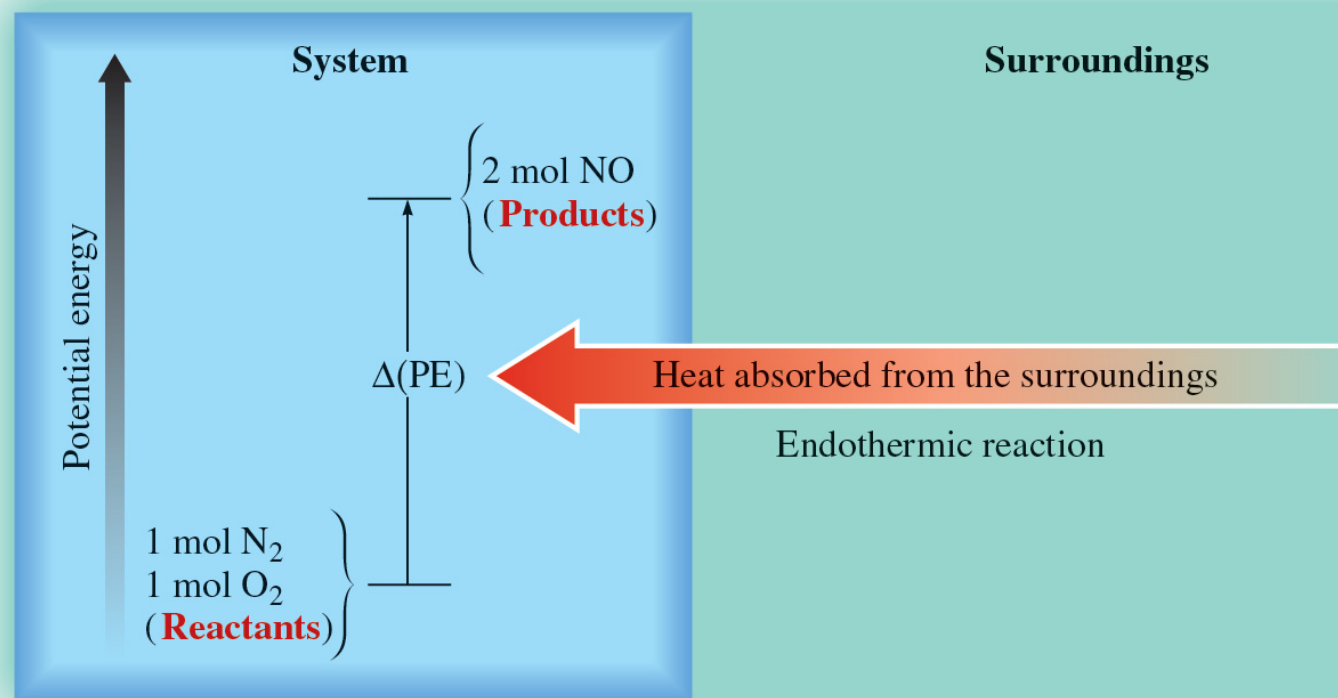
Reaction Mechanism (continued)

- Endothermic reactions
 - Energy that flows into the system as heat is used to increase the PE of the system
 - Products have higher PE than reactants

Section 6.1

The Nature of Energy

Figure 6.4 - Energy Diagram for the Formation of Nitric Oxide, an Endothermic Process



Section 6.1

The Nature of Energy

Thermodynamics

- Study of energy and its interconversions
- **First law of thermodynamics:** Energy of the universe is constant
 - Known as the law of conservation of energy

Section 6.1

The Nature of Energy

Internal Energy (E)

- Sum of kinetic and potential energies of all particles in the system
- E of a system can be changed by flow of work, heat, or both

$$\Delta E = q + w$$

- ΔE - Change in the system's internal energy
- q - Heat
- w - Work

Section 6.1

The Nature of Energy

Parts of Thermodynamic Quantities

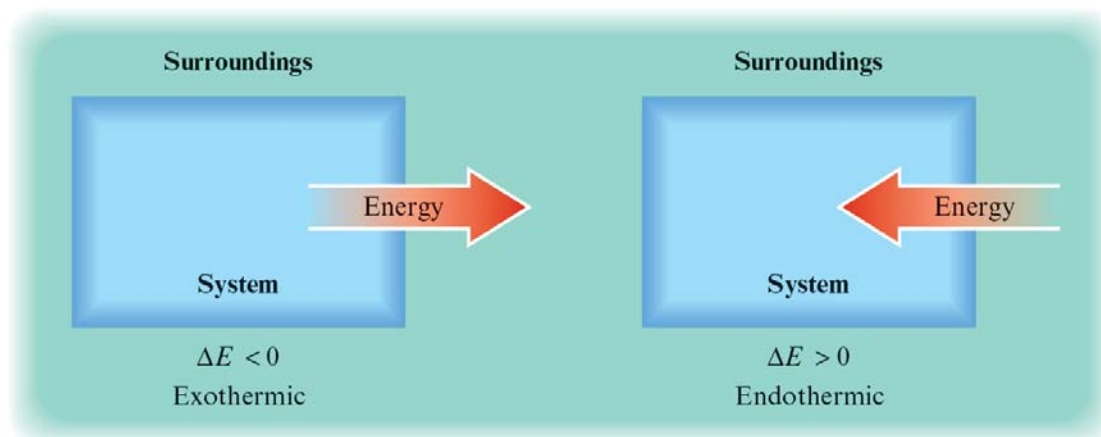
- Number - Gives the magnitude of change
- Sign - Indicates the direction of flow and reflects the system's point of view
 - In an endothermic system, $q = +x$
 - Positive sign indicates that the system's energy is increasing
 - In an exothermic system, $q = -x$
 - Negative sign indicates that the system's energy is decreasing

Section 6.1

The Nature of Energy

Parts of Thermodynamic Quantities (continued)

- Conventions that apply to the flow of work
 - When a system does work on the surroundings, w is negative
 - When the surroundings do work on the system, w is positive



Section 6.1

The Nature of Energy

Interactive Example 6.1 - Internal Energy

- Calculate ΔE for a system undergoing an endothermic process in which 15.6 kJ of heat flows and where 1.4 kJ of work is done on the system

Section 6.1

The Nature of Energy

Interactive Example 6.1 - Solution

- We use the following equation:

$$\Delta E = q + w$$

- $q = + 15.6 \text{ kJ}$, since the process is endothermic
- $w = + 1.4 \text{ kJ}$, since work is done on the system

$$\Delta E = 15.6 \text{ kJ} + 1.4 \text{ kJ} = 17.0 \text{ kJ}$$

- Thus, the system has gained 17.0 kJ of energy

Section 6.1

The Nature of Energy

Work

- Types of work associated with a chemical process
 - Work done by a gas through expansion
 - Work done to a gas through compression
- Example
 - In an automobile engine, heat from the combustion of gasoline expands the gases in the cylinder to push back the piston
 - This motion is then translated to the motion of the car

Section 6.1

The Nature of Energy

Deriving the Equation for Work

- Consider a gas confined to a cylindrical container with a movable piston
 - F is the force acting on the piston of area A
 - Pressure is defined as force per unit area

$$P = \frac{F}{A}$$

Section 6.1

The Nature of Energy

Deriving the Equation for Work (continued 1)

- Work is defined as force applied over a distance
 - If the piston moves a distance of Δh , work done is

$$\text{Work} = \text{force} \times \text{distance} = F \times \Delta h$$

- Since $P = F/A$ or $F = P \times A$,

$$\text{Work} = F \times \Delta h = P \times A \times \Delta h$$

- Volume of the cylinder equals the area of the piston times the height of the cylinder

Section 6.1

The Nature of Energy

Deriving the Equation for Work (continued 2)

- Change in volume ΔV resulting from the piston moving a distance Δh is

$$\Delta V = \text{final volume} - \text{initial volume} = A \times \Delta h$$

- Substituting $\Delta V = A \times \Delta h$ into the expression for work gives the magnitude of the work required to expand a gas ΔV against a pressure P

$$\text{Work} = P \times A \times \Delta h = P\Delta V$$

Section 6.1

The Nature of Energy

Deriving the Equation for Work (continued 3)

- Since the system is doing work on the surroundings, the sign of work should be negative
 - For an expanding gas, ΔV is a positive quantity because the V is increasing

$$w = -P\Delta V$$

- When a gas is compressed, ΔV is a negative quantity because V decreases, which makes w a positive quantity

Section 6.1

The Nature of Energy

Critical Thinking

- You are calculating ΔE in a chemistry problem
 - What if you confuse the system and the surroundings?
 - How would this affect the magnitude of the answer you calculate? The sign?

Section 6.1

The Nature of Energy

Interactive Example 6.2 - PV Work

- Calculate the work associated with the expansion of a gas from 46 L to 64 L at a constant external pressure of 15 atm

Section 6.1

The Nature of Energy

Interactive Example 6.2 - Solution

- For a gas at constant pressure,

$$w = -P\Delta V$$

- In this case $P = 15 \text{ atm}$ and $\Delta V = 64 - 46 = 18 \text{ L}$

$$w = -15 \text{ atm} \times 18 \text{ L} = -270 \text{ L} \cdot \text{atm}$$

- Note that since the gas expands, it does work on its surroundings

- Reality check

- Energy flows out of the gas, so w is a negative quantity

Section 6.1

The Nature of Energy

Interactive Example 6.3 - Internal Energy, Heat, and Work

- A balloon is being inflated to its full extent by heating the air inside it
 - In the final stages of this process, the volume of the balloon changes from $4.00 \times 10^6 \text{ L}$ to $4.50 \times 10^6 \text{ L}$ by the addition of $1.3 \times 10^8 \text{ J}$ of energy as heat
 - Assuming that the balloon expands against a constant pressure of 1.0 atm, calculate ΔE for the process
 - To convert between $\text{L} \cdot \text{atm}$ and J, use $1 \text{ L} \cdot \text{atm} = 101.3 \text{ J}$

Section 6.1

The Nature of Energy

Interactive Example 6.3 - Solution

- Where are we going?
 - To calculate ΔE
 - What do we know?
 - $V_1 = 4.00 \times 10^6 \text{ L}$
 - $q = +1.3 \times 10^8 \text{ J}$
 - $P = 1.0 \text{ atm}$
 - $1 \text{ L} \cdot \text{atm} = 101.3 \text{ J}$
 - $V_2 = 4.50 \times 10^6 \text{ L}$

Section 6.1

The Nature of Energy

Interactive Example 6.3 - Solution (continued 1)

- What do we need?

$$\Delta E = q + w$$

- How do we get there?

- What is the work done on the gas?

$$w = -P\Delta V$$

- What is ΔV ?

$$\Delta V = V_2 - V_1 = 4.50 \times 10^6 \text{ L} - 4.00 \times 10^6 \text{ L} = 5.0 \times 10^5 \text{ L}$$

Section 6.1

The Nature of Energy

Interactive Example 6.3 - Solution (continued 2)

- What is the work?

$$w = -P\Delta V = -1.0 \text{ atm} \times 5.0 \times 10^5 \text{ L} = -5.0 \times 10^5 \text{ L} \cdot \text{atm}$$

- The negative sign makes sense because the gas is expanding and doing work on the surroundings
- To calculate ΔE , we must sum q and w
 - However, since q is given in units of J and w is given in units of L · atm, we must change the work to units of joules

$$w = -5.0 \times 10^5 \cancel{\text{L} \cdot \text{atm}} \times \frac{101.3 \text{ J}}{\cancel{\text{L} \cdot \text{atm}}} = -5.1 \times 10^7 \text{ J}$$

Section 6.1

The Nature of Energy

Interactive Example 6.3 - Solution (continued 3)

- Then,

$$\Delta E = q + w = \left(+1.3 \times 10^8 \text{ J}\right) + \left(-5.1 \times 10^7 \text{ J}\right) = 8 \times 10^7 \text{ J}$$

- Reality check
 - Since more energy is added through heating than the gas expends doing work, there is a net increase in the internal energy of the gas in the balloon
 - Hence ΔE is positive



Section 6.1

The Nature of Energy

example

- Label the following process as exothermic or endothermic: Ice increases in temperature when you touch it
 - a. Exothermic
 - b. Endothermic



Section 6.1

The Nature of Energy

Join In (6)

- Gas A_2 reacts with gas B_2 to form gas AB
 - The bond energy of AB is much greater than the bond energy of either A_2 or B_2
 - Is the reaction for the formation of AB exothermic or endothermic?
 - a. Exothermic
 - b. Endothermic

Section 6.2

Enthalpy and Calorimetry

Enthalpy (H)

- A state function that is defined as:

$$H = E + PV$$

- E - Internal energy of the system
- P - Pressure of the system
- V - Volume of the system

Section 6.2

Enthalpy and Calorimetry

Change in Enthalpy (ΔH)

- For a process carried out at constant pressure and where the only work allowed is that from a volume change:

$$\Delta H = q_P$$

- q_P is the heat at constant pressure
- At constant pressure, the change in enthalpy ΔH of the system is equal to the energy flow as heat

Section 6.2

Enthalpy and Calorimetry

Enthalpy and PV Work

- For a chemical reaction, the enthalpy change is given by the following equation:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

- When $H_{\text{products}} > H_{\text{reactants}}$, ΔH is positive
 - Heat is absorbed by the system, and the reaction is endothermic
- When $H_{\text{products}} < H_{\text{reactants}}$, ΔH is negative
 - Overall decrease in enthalpy is achieved by the generation of heat, and the reaction is exothermic

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.4 - Enthalpy

- When 1 mole of methane (CH_4) is burned at constant pressure, 890 kJ of energy is released as heat
 - Calculate ΔH for a process in which a 5.8-g sample of methane is burned at constant pressure

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.4 - Solution

- Where are we going?
 - To calculate ΔH
 - What do we know?
 - $q_p = \Delta H = -890 \text{ kJ/mol CH}_4$
 - Mass = 5.8 g CH₄
 - Molar mass CH₄ = 16.04 g

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.4 - Solution (continued 1)

- How do we get there?
 - What are the moles of CH₄ burned?

$$5.8 \cancel{\text{ g CH}_4} \times \frac{1 \text{ mol CH}_4}{16.04 \cancel{\text{ g CH}_4}} = 0.36 \text{ mol CH}_4$$

- What is ΔH ?

$$\Delta H = 0.36 \cancel{\text{ mol CH}_4} \times \frac{-890 \text{ kJ}}{\cancel{\text{ mol CH}_4}} = -320 \text{ kJ}$$

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.4 - Solution (continued 2)

- Thus, when a 5.8-g sample of CH_4 is burned at constant pressure,

$$\Delta H = \text{heat flow} = -320 \text{ kJ}$$

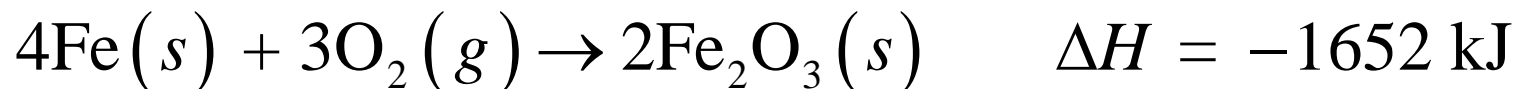
- Reality check
 - In this case, a 5.8-g sample of CH_4 is burned
 - Since this amount is smaller than 1 mole, less than 890 kJ will be released as heat

Section 6.2

Enthalpy and Calorimetry

Exercise

- The overall reaction in a commercial heat pack can be represented as



- a. How much heat is released when 4.00 moles of iron are reacted with excess O_2 ?

1650 kJ is released

Section 6.2

Enthalpy and Calorimetry

Exercise (continued)

- b. How much heat is released when 1.00 mole of Fe_2O_3 is produced?

826 kJ released

- b. How much heat is released when 1.00 g iron is reacted with excess O_2 ?

7.39 kJ released

- b. How much heat is released when 10.0 g Fe and 2.00 g O_2 are reacted?

34.4 kJ released

Section 6.2

Enthalpy and Calorimetry

Calorimetry

- Science of measuring heat
 - Based on observations of temperature change when a body absorbs or discharges energy as heat
- **Calorimeter**: Device used experimentally to determine the heat associated with a chemical reaction

Section 6.2

Enthalpy and Calorimetry

Heat Capacity (C)

$$C = \frac{\text{heat absorbed}}{\text{increase in temperature}}$$

- **Specific heat capacity:** Energy required to raise the temperature of one gram of a substance by one degree Celsius
 - Units - J/° C · g or J/K · g

Section 6.2

Enthalpy and Calorimetry

Table 6.1 - The Specific Heat Capacities of Some Common Substances

Substance	Specific Heat Capacity ($\text{J}/^\circ\text{C} \cdot \text{g}$)
$\text{H}_2\text{O}(l)$	4.18
$\text{H}_2\text{O}(s)$	2.03
$\text{Al}(s)$	0.89
$\text{Fe}(s)$	0.45
$\text{Hg}(l)$	0.14
$\text{C}(s)$	0.71

Section 6.2

Enthalpy and Calorimetry

Heat Capacity (C) (continued)

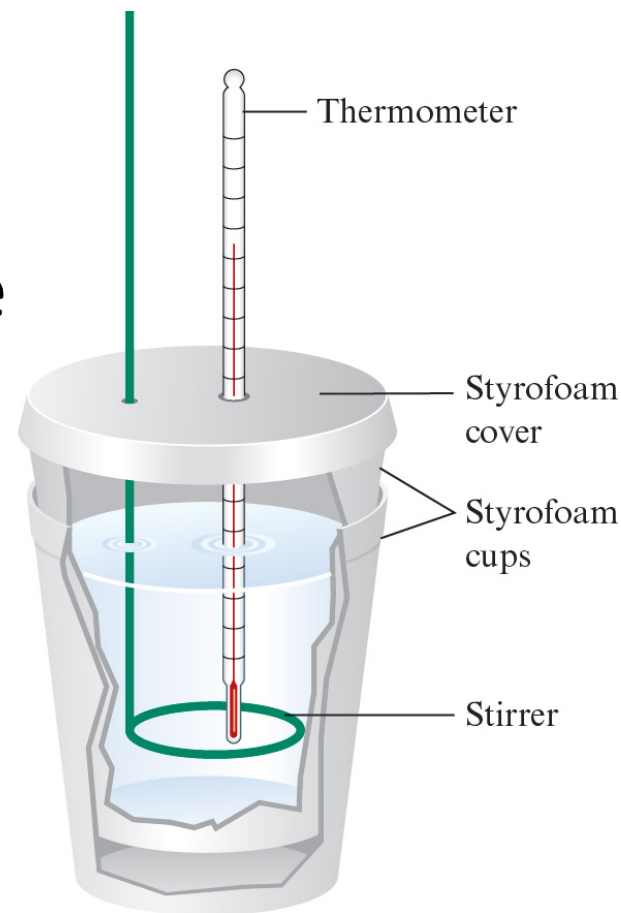
- **Molar heat capacity:** Energy required to raise the temperature of one mole of a substance by one degree Celsius
 - Units - $\text{J}/^{\circ}\text{C} \cdot \text{mol}$ or $\text{J}/\text{K} \cdot \text{mol}$
- Heat capacities of metals are different from that of water
 - It takes less energy to change the temperature of a gram of a metal by 1°C than for a gram of water

Section 6.2

Enthalpy and Calorimetry

Coffee-Cup Calorimeter

- Contains two nested Styrofoam cups with a cover through which a stirrer and thermometer can be inserted
 - Outer cup is used to provide extra insulation
 - Inner cup holds the solution in which the reaction occurs



Section 6.2

Enthalpy and Calorimetry

Constant-Pressure Calorimetry

- Atmospheric pressure remains constant during the process
- Used to determine changes in enthalpy for reactions that occur in solution
 - $\Delta H = q_p$

Section 6.2

Enthalpy and Calorimetry

Constant-Pressure Calorimetry (continued 1)

- If two reactants at the same temperature are mixed and the resulting solution gets warmer, this means the reaction taking place is exothermic
 - Endothermic reaction cools the solution

Section 6.2

Enthalpy and Calorimetry

Constant-Pressure Calorimetry: Determining Change in Enthalpy for a Neutralization Reaction

Energy (as heat) released by the reaction

= energy (as heat) absorbed by the solution

= specific heat capacity \times mass of solution \times increase in temperature

$$= s \times m \times \Delta T$$

- When twice the amount of solution has been mixed, twice as much heat would be produced
 - Heat of a reaction is an extensive property
 - Depends directly on the amount of substance

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.5 - Constant-Pressure Calorimetry

- When 1.00 L of 1.00 M $\text{Ba}(\text{NO}_3)_2$ solution at 25.0°C is mixed with 1.00 L of 1.00 M Na_2SO_4 solution at 25.0°C in a calorimeter, the white solid BaSO_4 forms, and the temperature of the mixture increases to 28.1°C

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.5 - Constant-Pressure Calorimetry (continued)

- Assume that:
 - The calorimeter absorbs only a negligible quantity of heat
 - The specific heat capacity of the solution is $4.18 \text{ J/}^{\circ} \text{ C} \cdot \text{g}$
 - The density of the final solution is 1.0 g/mL
- Calculate the enthalpy change per mole of BaSO_4 formed

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.5 - Solution

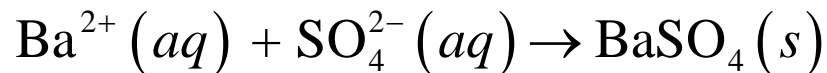
- Where are we going?
 - To calculate ΔH per mole of BaSO_4 formed
 - What do we know?
 - 1.00 L of 1.00 M $\text{Ba}(\text{NO}_3)_2$
 - 1.00 L of 1.00 M Na_2SO_4
 - $T_{\text{initial}} = 25.0^\circ \text{ C}$ and $T_{\text{final}} = 28.1^\circ \text{ C}$
 - Heat capacity of solution = $4.18 \text{ J/}^\circ \text{ C} \cdot \text{g}$
 - Density of final solution = 1.0 g/mL

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.5 - Solution (continued 1)

- What do we need?
 - Net ionic equation for the reaction
 - The ions present before any reaction occurs are Ba^{2+} , NO_3^- , Na^+ , and SO_4^{2-}
 - The Na^+ and NO_3^- ions are spectator ions, since NaNO_3 is very soluble in water and will not precipitate under these conditions
 - The net ionic equation for the reaction is:



Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.5 - Solution (continued 2)

- How do we get there?
 - What is ΔH ?
 - Since the temperature increases, formation of solid BaSO_4 must be exothermic; ΔH is negative

Heat evolved by the reaction

= heat absorbed by the solution

= specific heat capacity \times mass of solution \times increase in temperature

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.5 - Solution (continued 3)

- What is the mass of the final solution?

$$\text{Mass of solution} = 2.00 \cancel{\text{ L}} \times \frac{1000 \cancel{\text{ mL}}}{1 \cancel{\text{ L}}} \times \frac{1.0 \text{ g}}{\cancel{\text{ mL}}} = 2.0 \times 10^3 \text{ g}$$

- What is the temperature increase?

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 28.1^{\circ}\text{C} - 25.0^{\circ}\text{C} = 3.1^{\circ}\text{C}$$

- How much heat is evolved by the reaction?

$$\text{Heat evolved} = (4.18 \text{ J}/^{\circ}\cancel{\text{C}} \cdot \cancel{\text{g}})(2.0 \times 10^3 \cancel{\text{ g}})(3.1^{\circ}\cancel{\text{C}}) = 2.6 \times 10^4 \text{ J}$$

Section 6.2

Enthalpy and Calorimetry

Interactive Example 6.5 - Solution (continued 4)

- Thus,

$$q = q_P = \Delta H = -2.6 \times 10^4 \text{ J}$$

- What is ΔH per mole of BaSO_4 formed?
 - Since 1.0 L of 1.0 M $\text{Ba}(\text{NO}_3)_2$ contains 1 mole of Ba^{2+} ions and 1.0 L of 1.0 M Na_2SO_4 contains 1.0 mole of SO_4^{2-} ions, 1.0 mole of solid BaSO_4 is formed in this experiment
 - Thus the enthalpy change per mole of BaSO_4 formed is

$$\Delta H = -2.6 \times 10^4 \text{ J/mol} = -26 \text{ kJ/mol}$$

Section 6.2

Enthalpy and Calorimetry

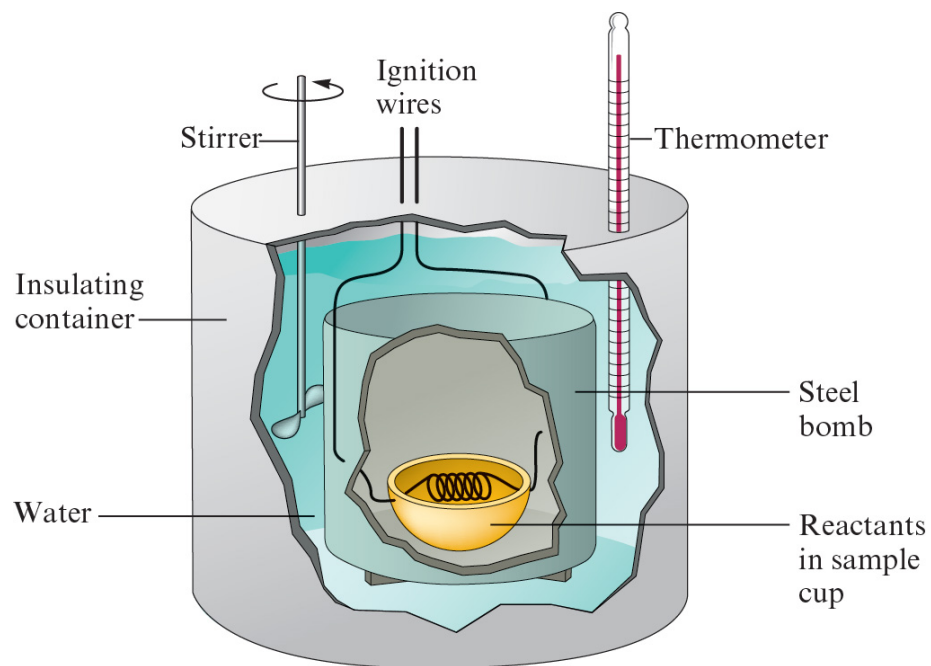
Constant-Volume Calorimetry

- Used to study the energy changes in reactions under conditions of constant volume
 - No work is done because V must change for pressure–volume work to be performed
- Device used - Bomb calorimeter
 - Weighed reactants are placed within a rigid steel container and ignited
 - Energy change is determined by the increase in temperature of the water and other calorimeter parts

Section 6.2

Enthalpy and Calorimetry

Figure 6.7 - A Bomb Calorimeter



Section 6.2

Enthalpy and Calorimetry

Constant-Volume Calorimetry (continued 1)

- For a constant-volume process, $\Delta V = 0$
 - Therefore, $w = -P\Delta V = 0$

$$\Delta E = q + w = q = q_v \quad (\text{constant volume})$$

- Energy released by the reaction
 - = temperature increase \times energy required to change the temperature by 1°C
 - = $\Delta T \times$ heat capacity of calorimeter

Section 6.2

Enthalpy and Calorimetry

Constant-Volume Calorimetry (continued 2)

- Since no work is done in this case, ΔE is equal to the heat

$$\Delta E = q + w = q \quad \text{since } w = 0$$

Section 6.2

Enthalpy and Calorimetry

Example 6.6 - Constant-Volume Calorimetry

- It has been suggested that hydrogen gas obtained by the decomposition of water might be a substitute for natural gas (principally methane)
 - To compare the energies of combustion of these fuels, the following experiment was carried out using a bomb calorimeter with a heat capacity of $11.3 \text{ kJ/}^{\circ} \text{C}$

Section 6.2

Enthalpy and Calorimetry

Example 6.6 - Constant-Volume Calorimetry (continued)

- When a 1.50-g sample of methane gas was burned with excess oxygen in the calorimeter, the temperature increased by 7.3°C
- When a 1.15-g sample of hydrogen gas was burned with excess oxygen, the temperature increase was 14.3°C
 - Compare the energies of combustion (per gram) for hydrogen and methane

Section 6.2

Enthalpy and Calorimetry

Example 6.6 - Solution

- Where are we going?
 - To calculate ΔH of combustion per gram for H_2 and CH_4
 - What do we know?
 - $1.50 \text{ g } CH_4 \Rightarrow \Delta T = 7.3^\circ \text{ C}$
 - $1.15 \text{ g } H_2 \Rightarrow \Delta T = 14.3^\circ \text{ C}$
 - Heat capacity of calorimeter = $11.3 \text{ kJ/}^\circ \text{ C}$
 - What do we need?
 - $\Delta E = \Delta T \times \text{heat capacity of calorimeter}$

Section 6.2

Enthalpy and Calorimetry

Example 6.6 - Solution (continued 1)

- How do we get there?
 - What is the energy released for each combustion?
 - For CH₄, we calculate the energy of combustion for methane using the heat capacity of the calorimeter (11.3 kJ/°C) and the observed temperature increase of 7.3°C

$$\begin{aligned}\text{Energy released in the combustion of 1.5 g CH}_4 &= (11.3 \text{ kJ/}^\circ\text{C})(7.3^\circ\text{C}) \\ &= 83 \text{ kJ}\end{aligned}$$

Section 6.2

Enthalpy and Calorimetry

Example 6.6 - Solution (continued 2)

$$\text{Energy released in the combustion of 1 g CH}_4 = \frac{83 \text{ kJ}}{1.5 \text{ g}} = 55 \text{ kJ/g}$$

■ For H₂,

$$\begin{aligned} \text{Energy released in the combustion of 1.15 g H}_2 &= (11.3 \text{ kJ/}^\circ\text{C})(14.3^\circ\text{C}) \\ &= 162 \text{ kJ} \end{aligned}$$

$$\text{Energy released in the combustion of 1 g H}_2 = \frac{162 \text{ kJ}}{1.15 \text{ g}} = 141 \text{ kJ/g}$$

Section 6.2

Enthalpy and Calorimetry

Example 6.6 - Solution (continued 3)

- How do the energies of combustion compare?
 - The energy released in the combustion of 1 g hydrogen is approximately 2.5 times that for 1 g methane, indicating that hydrogen gas is a potentially useful fuel



Section 6.2

Enthalpy and Calorimetry

Join In (11)

- If 5.0 kJ of energy is added to a 15.5-g sample of water at $10.^{\circ}\text{C}$, the water is:
 - a. boiling
 - b. completely vaporized
 - c. frozen solid
 - d. still a liquid



Section 6.2

Enthalpy and Calorimetry

Join In (12)

- A 50.0-g sample of water at 80°C is added to a 50.0-g sample of water at 20°C
 - The final temperature of the water should be:
 - a. between 20°C and 50°C
 - b. 50°C
 - c. between 50°C and 80°C

Section 6.3

Hess's Law

Hess's Law

- In going from a particular set of reactants to a particular set of products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps

Section 6.3

Hess's Law

Characteristics of Enthalpy Changes

- If a reaction is reversed, the sign of ΔH is also reversed
- Magnitude of ΔH is directly proportional to the quantities of reactants and products in a reaction
 - If the coefficients in a balanced reaction are multiplied by an integer, the value of ΔH is multiplied by the same integer

Section 6.3

Hess's Law

Critical Thinking

- What if Hess's law were not true?
 - What are some possible repercussions this would have?

Section 6.3

Hess's Law

Problem-Solving Strategy - Hess's Law

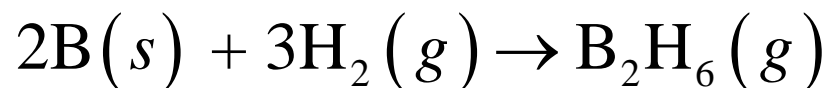
- Work backward from the required reaction
 - Use the reactants and products to decide how to manipulate the other given reactions at your disposal
- Reverse any reactions as needed to give the required reactants and products
- Multiply reactions to give the correct numbers of reactants and products

Section 6.3

Hess's Law

Interactive Example 6.8 - Hess's Law II

- Diborane (B_2H_6) is a highly reactive boron hydride that was once considered as a possible rocket fuel for the U.S. space program
 - Calculate ΔH for the synthesis of diborane from its elements, according to the following equation:



Section 6.3

Hess's Law

Interactive Example 6.8 - Hess's Law II (continued)

- Use the following data:

Reaction	ΔH
a. $2\text{B}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{B}_2\text{O}_3(s)$	-1273 kJ
b. $\text{B}_2\text{H}_6(g) + 3\text{O}_2(g) \rightarrow \text{B}_2\text{O}_3(s) + 3\text{H}_2\text{O}(g)$	-2035 kJ
c. $\text{H}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l)$	-286 kJ
d. $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$	44 kJ

Section 6.3

Hess's Law

Interactive Example 6.8 - Solution

- To obtain ΔH for the required reaction, we must somehow combine equations (a), (b), (c), and (d) to produce that reaction and add the corresponding ΔH values
 - This can best be done by focusing on the reactants and products of the required reaction
 - The reactants are $\text{B}(s)$ and $\text{H}_2(g)$, and the product is $\text{B}_2\text{H}_6(g)$

Section 6.3

Hess's Law

Interactive Example 6.8 - Solution (continued 1)

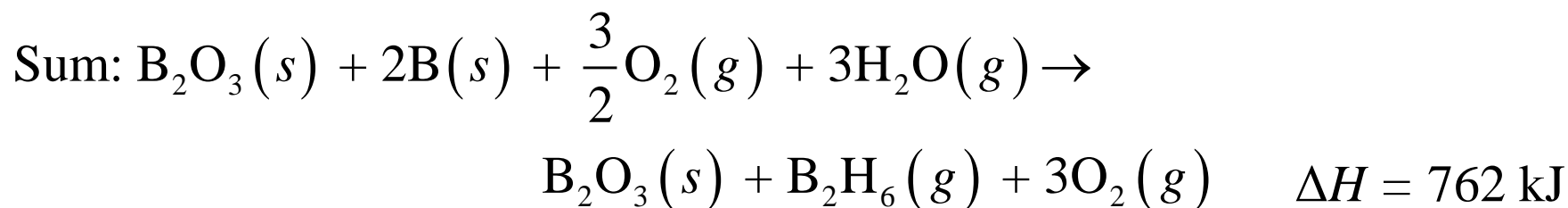
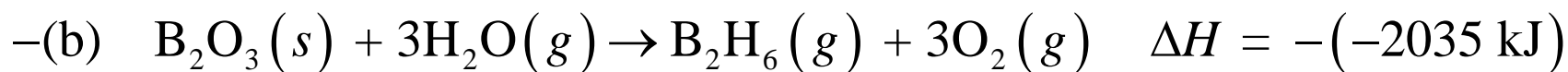
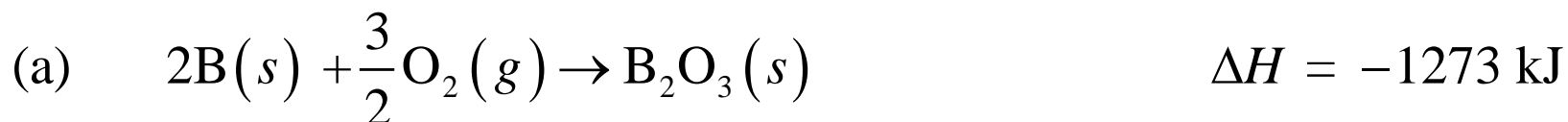
- How can we obtain the correct equation?
 - Reaction (a) has $B(s)$ as a reactant, as needed in the required equation
 - Thus reaction (a) will be used as it is
 - Reaction (b) has $B_2H_6(g)$ as a reactant, but this substance is needed as a product
 - Thus reaction (b) must be reversed, and the sign of ΔH must be changed accordingly

Section 6.3

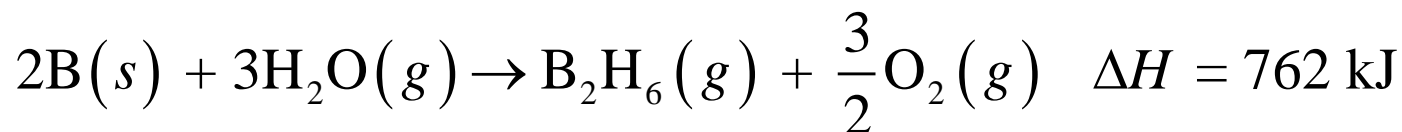
Hess's Law

Interactive Example 6.8 - Solution (continued 2)

- Up to this point we have:



- Deleting the species that occur on both sides gives



Section 6.3

Hess's Law

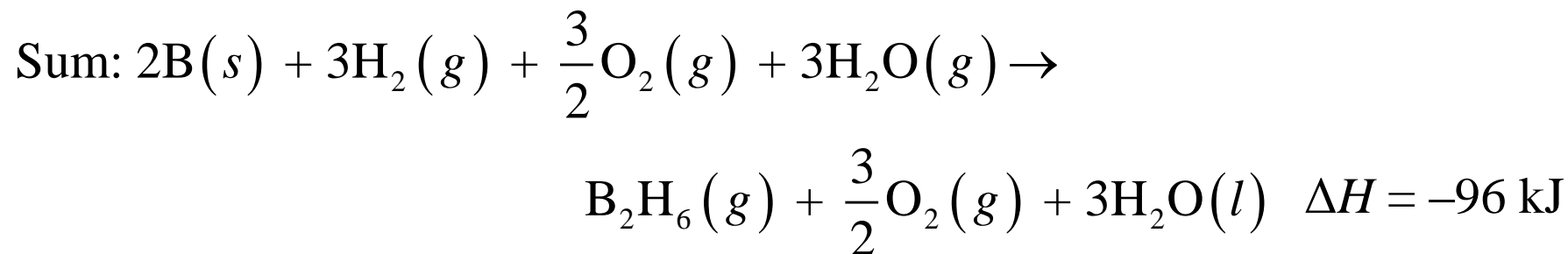
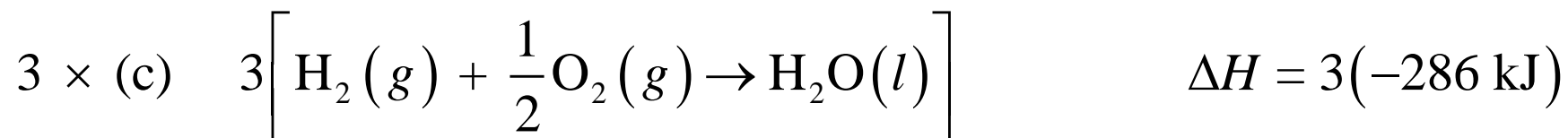
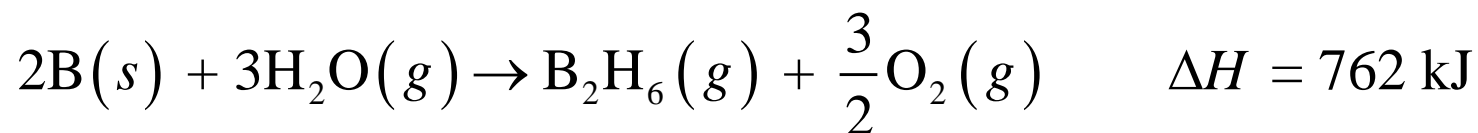
Interactive Example 6.8 - Solution (continued 3)

- We are closer to the required reaction, but we still need to remove $\text{H}_2\text{O}(g)$ and $\text{O}_2(g)$ and introduce $\text{H}_2(g)$ as a reactant
 - We can do this using reactions (c) and (d)
 - Multiply reaction (c) and its ΔH value by 3 and add the result to the preceding equation

Section 6.3

Hess's Law

Interactive Example 6.8 - Solution (continued 4)

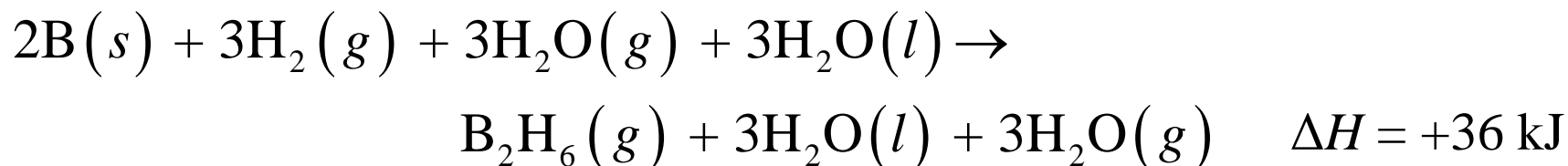
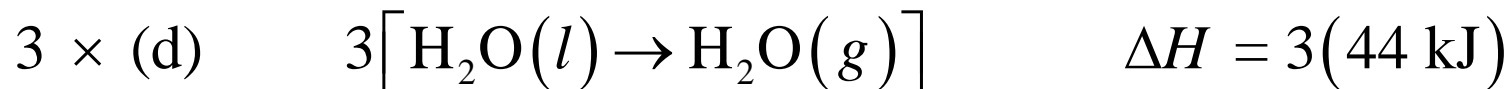


Section 6.3

Hess's Law

Interactive Example 6.8 - Solution (continued 5)

- We can cancel the $3/2 \text{ O}_2(g)$ on both sides, but we cannot cancel the H_2O because it is gaseous on one side and liquid on the other
 - This can be solved by adding reaction (d), multiplied by 3:

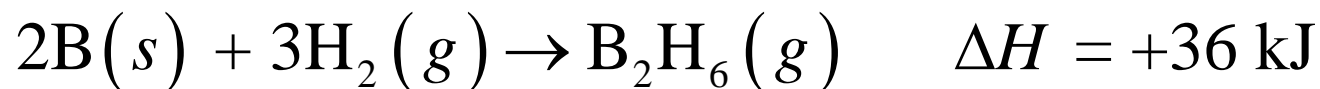


Section 6.3

Hess's Law

Interactive Example 6.8 - Solution (continued 6)

- This gives the reaction required by the problem



- Thus ΔH for the synthesis of 1 mole of diborane from the elements is +36 kJ

Section 6.3

Hess's Law

Join In (15)

- Given the following equation, which of the following statement(s) is (are) true?



- I. The reaction is exothermic
- II. When 0.500 mol sulfur is reacted, 148 kJ of energy is released
- III. When 32.0 g of sulfur is burned, $2.96 \times 10^5 \text{ J}$ of energy is released

Section 6.3

Hess's Law

Homework

1. Which has the greater kinetic energy, an object with a mass of 2.0 kg and a velocity of 1.0 m/s or an object with a mass of 1.0 kg and a velocity of 2.0 m/s?
2. For the process $\text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{O}(g)$ at 298 K and 1.0 atm, ΔH is more positive than ΔE by 2.5 kJ/mol. What does the 2.5 kJ/mol quantity represent?

Section 6.3

Hess's Law

Thank you