Chemistry
Matter and Change
Chapter 15
Energy and Chemical Change
Click the mouse button or press the Space Bar to continue.
Energy and Chemical Change

**Section 15.1** Energy

**Section 15.2** Heat

**Section 15.3** Thermochemical Equations

**Section 15.4** Calculating Enthalpy Change

**Section 15.5** Reaction Spontaneity

Click a hyperlink or folder tab to view the corresponding slides.
Section 15.1 Energy

Objectives

• **Define** energy.

• **Distinguish** between potential and kinetic energy.

• **Relate** chemical potential energy to the heat lost or gained in chemical reactions.

• **Calculate** the amount of heat absorbed or released by a substance as its temperature changes.

Review Vocabulary

temperature: a measure of the average kinetic energy of the particles in a sample of matter
Section 15.1 Energy (cont.)

New Vocabulary

energy

law of conservation of energy

chemical potential energy

heat

calorie

joule

specific heat

MAIN Idea  Energy can change form and flow, but it is always conserved.
The Nature of Energy

- **Energy** is the ability to do work or produce heat.
- Two forms of energy exist, potential and kinetic.
- Potential energy is due to composition or position.
- Kinetic energy is energy of motion.
Energy can be converted from one form to another, but is always conserved.

In the diagram, how is electricity being produced?
The law of conservation of energy states that in any chemical reaction or physical process, energy can be converted from one form to another, but it is neither created nor destroyed—also known as the first law of thermodynamics.
The Nature of Energy (cont.)

- **Chemical potential energy** is energy stored in a substance because of its composition.

- Chemical potential energy is important in chemical reactions.

- **Heat** is energy that is in the process of flowing from a warmer object to a cooler object.

- $q$ is used to symbolize heat.
Measuring Heat

- A **calorie** is defined as the amount of energy required to raise the temperature of one gram of water one degree Celsius.

- Food is measured in Calories, or 1000 calorie (kilocalorie).

- A **joule** is the SI unit of heat and energy, equivalent to 0.2390 calories.
### Measuring Heat (cont.)

#### Table 15.1: Relationships Among Energy Units

<table>
<thead>
<tr>
<th>Relationship</th>
<th>Conversion Factors</th>
</tr>
</thead>
</table>
| 1 J = 0.2390 cal              | \[
\frac{1 \text{ J}}{0.2390 \text{ cal}} = \frac{0.2390 \text{ cal}}{1 \text{ J}}
\] |
| 1 cal = 4.184 J              | \[
\frac{1 \text{ cal}}{4.184 \text{ J}} = \frac{4.184 \text{ J}}{1 \text{ cal}}
\] |
| 1 Calorie = 1 kcal           | \[
\frac{1 \text{ Calorie}}{1000 \text{ cal}} = \frac{1000 \text{ cal}}{1 \text{ Calorie}}
\] |
Specific Heat

- The **specific heat** of any substance is the amount of heat required to raise one gram of that substance one degree Celsius.

- Some objects require more heat than others to raise their temperature.

### Table 15.2: Specific Heats at 298 K (25°C)

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific heat J/(g·°C)</th>
</tr>
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<tbody>
<tr>
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</tr>
</tbody>
</table>
• Calculating heat absorbed and released

\[ q = c \times m \times \Delta T \]

- \( q \) = heat absorbed or released
- \( c \) = specific heat of substance
- \( m \) = mass of substance in grams
- \( \Delta T \) = change in temperature in Celsius
The heat required to raise one gram of a substance by one degree Celsius is called _____.

A. joule  
B. calorie  
C. specific heat  
D. energy
Section 15.1 Assessment

Which of the following is an example of chemical potential energy?

A. the moon orbiting Earth
B. a car battery
C. a compressed spring
D. a roller coaster at the top of a hill
Click the mouse button to return to the Chapter Menu.
Section 15.2 Heat

Objectives

• Describe how a calorimeter is used to measure energy that is absorbed or released.

• Explain the meaning of enthalpy and enthalpy change in chemical reactions and processes.

Review Vocabulary

pressure: force applied per unit area
The enthalpy change for a reaction is the enthalpy of the products minus the enthalpy of the reactants.
Calorimetry

• A **calorimeter** is an insulated device used for measuring the amount of heat absorbed or released in a chemical reaction or physical process.

Chemical Energy and the Universe

• **Thermochemistry** is the study of heat changes that accompany chemical reactions and phase changes.

• The **system** is the specific part of the universe that contains the reaction or process you wish to study.
Chemical Energy and the Universe (cont.)

- The **surroundings** are everything else other than the system in the universe.

- The **universe** is defined as the system plus the surroundings.
Chemical Energy and the Universe (cont.)

• Chemists are interested in changes in energy during reactions.

• **Enthalpy** is the heat content of a system at constant pressure.

• **Enthalpy (heat) of reaction** is the change in enthalpy during a reaction symbolized as $\Delta H_{\text{rxn}}$.

\[
\Delta H_{\text{rxn}} = H_{\text{final}} - H_{\text{initial}}
\]

\[
\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}
\]
Enthalpy changes for exothermic reactions are always negative.

\[ 4 \text{Fe}_2 + 3 \text{O}_2 \rightarrow 2 \text{Fe}_2\text{O}_3 \quad \Delta H = -1625 \text{kJ} \]

Enthalpy changes for endothermic reactions are always positive.

\[ 6 \text{CO}_2 + 6 \text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2 \quad \Delta H = 2808 \text{kJ} \]
The Heat-Pack Reaction

4Fe(s) + 3O₂(g) → 2Fe₂O₃(s)

Heat to surroundings

ΔH = -1625 kJ

Exothermic Reaction

ΔH < 0
The Cold-Pack Process

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

Reactant

Products

Enthalpy

\[ \Delta H = +27 \text{ kJ} \]

Endothermic Process

\[ \Delta H > 0 \]
In thermochemistry, the specific part of the universe you are studying is called _____.

A. system
B. area
C. enthalpy
D. surroundings
What is the heat content of a system at constant pressure called?

A. heat of reaction
B. heat of the system
C. enthalpy
D. entropy
Click the mouse button to return to the Chapter Menu.
Section 15.3 Thermochemical Equations

Objectives

- **Write** thermochemical equations for chemical reactions and other processes.
- **Describe** how energy is lost or gained during changes of state.
- **Calculate** the heat absorbed or released in a chemical reaction.

Review Vocabulary

**combustion reaction**: a chemical reaction that occurs when a substance reacts with oxygen, releasing energy in the form of heat and light
Section 15.3 Thermochemical Equations (cont.)

New Vocabulary

thermochemical equation
enthalpy (heat) of combustion
molar enthalpy (heat) of vaporization
molar enthalpy (heat) of fusion

MAIN Idea Thermochemical equations express the amount of heat released or absorbed by chemical reactions.
Writing Thermochemical Equations

• A **thermochemical equation** is a balanced chemical equation that includes the physical states of all reactants and products, and energy change.

• The **enthalpy (heat) of combustion** of a substance is the enthalpy change for the complete burning of one mole of the substance.
## Table 15.3 Standard Enthalpies of Combustion

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>$\Delta H^\circ_{\text{comb}}$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sucrose (table sugar)</td>
<td>$C_{12}H_{22}O_{11}(s)$</td>
<td>$-5644$</td>
</tr>
<tr>
<td>Octane (a component of gasoline)</td>
<td>$C_8H_{18}(l)$</td>
<td>$-5471$</td>
</tr>
<tr>
<td>Glucose (a simple sugar found in fruit)</td>
<td>$C_6H_{12}O_{6}(s)$</td>
<td>$-2808$</td>
</tr>
<tr>
<td>Propane (a gaseous fuel)</td>
<td>$C_3H_8(g)$</td>
<td>$-2219$</td>
</tr>
<tr>
<td>Methane (a gaseous fuel)</td>
<td>$CH_4(g)$</td>
<td>$-891$</td>
</tr>
</tbody>
</table>
Changes of State

- **Molar enthalpy (heat) of vaporization** refers to the heat required to vaporize one mole of a liquid substance.

- **Molar enthalpy (heat) of fusion** is the amount of heat required to melt one mole of a solid substance.
### Table 15.4

#### Standard Enthalpies of Vaporization and Fusion

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>$\Delta H^o_{\text{vap}}$ (kJ/mol)</th>
<th>$\Delta H^o_{\text{fus}}$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H$_2$O</td>
<td>40.7</td>
<td>6.01</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C$_2$H$_5$OH</td>
<td>38.6</td>
<td>4.94</td>
</tr>
<tr>
<td>Methanol</td>
<td>CH$_3$OH</td>
<td>35.2</td>
<td>3.22</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>CH$_3$COOH</td>
<td>23.4</td>
<td>11.7</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH$_3$</td>
<td>23.3</td>
<td>5.66</td>
</tr>
</tbody>
</table>
Changes of State (cont.)

Phase Changes for Water

Enthalpy

$\Delta H_{\text{vap}} = +40.7 \text{ kJ}$

$\Delta H_{\text{cond}} = -40.7 \text{ kJ}$

$\Delta H_{\text{fus}} = +6.01 \text{ kJ}$

$\Delta H_{\text{solid}} = -6.01 \text{ kJ}$
Combustion Reactions

• Combustion is the reaction of a fuel with oxygen.

• Food is the fuel in combustion reactions in biological systems.
The amount of energy required to melt one mole of a solid is called _____.

A. molar enthalpy of vaporization
B. molar enthalpy of melting
C. molar enthalpy of fusion
D. molar enthalpy of condensation
A thermochemical equation specifies about changes in _____.

A. temperature
B. pressure
C. enthalpy
D. molar mass
Click the mouse button to return to the Chapter Menu.
Section 15.4 Calculating Enthalpy Change

Objectives

- **Apply** Hess’s law to calculate the enthalpy change for a reaction.
- **Explain** the basis for the table of standard enthalpies of formation.
- **Calculate** $\Delta H_{\text{rxn}}$ using thermochemical equations.
- **Determine** the enthalpy change for a reaction using standard enthalpies of formation data.

Review Vocabulary

**allotrope**: one of two or more forms of an element with different structures and properties when they are in the same state
Section 15.4 Calculating Enthalpy Change (cont.)

New Vocabulary

Hess’s law

standard enthalpy (heat) of formation

MAIN Idea The enthalpy change for a reaction can be calculated using Hess’s law.
Hess's Law

- **Hess’s law** states that if you can add two or more thermochemical equations to produce a final equation for a reaction, then the sum of the enthalpy changes for the individual reactions is the enthalpy change for the final reaction.
**Hess's Law (cont.)**

The Synthesis of Sulfur Trioxide

\[ 2S(s) + 2O_2(g) \rightarrow 2SO_2(g) \quad \Delta H = -594 \text{ kJ} \]

\[ 2SO_2(g) + O_2(g) \rightarrow 2SO_3(g) \quad \Delta H = -198 \text{ kJ} \]

Overall energy change:
\[ \Delta H = -792 \text{ kJ} \]
Standard Enthalpy (Heat) of Formation

• The **standard enthalpy (heat) of formation** is defined as the change in enthalpy that accompanies the formation of one mole of the compound in its standard state from its elements in their standard states.
Elements in their standard states have a $\Delta H_f^\circ$ of 0.0 kJ.

The formation of compounds are placed above or below elements in their standard states.
Standard Enthalpy (Heat) of Formation (cont.)

Standard Heats of Formation

\[ \Delta H_f^\circ (\text{NO}_2) = +33.2 \text{ kJ/mol} \]

\[ \Delta H_f^\circ (\text{SO}_3) = -396 \text{ kJ/mol} \]

\[ \Delta H_f^\circ (\text{N}_2, \text{O}_2, \text{S}(s)) = 0.0 \]

Chemical species: NO\(_2\)(g), N\(_2\)(g), O\(_2\)(g), S(s), SO\(_3\)(g)
Standard Enthalpy (Heat) of Formation (cont.)

- Standard enthalpies of formation can be used to calculate the enthalpies for many reactions under standard conditions by using Hess’s law.

- The summation equation

$$\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ (\text{products}) - \sum \Delta H_f^\circ (\text{reactants})$$
## Standard Enthalpy (Heat) of Formation (cont.)

### Table 15.5: Standard Enthalpies of Formation

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formation Equation</th>
<th>$\Delta H_f^\circ$(kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$H_2S(g)$</td>
<td>$H_2(g) + S(s) \rightarrow H_2S(g)$</td>
<td>$-21$</td>
</tr>
<tr>
<td>$HF(g)$</td>
<td>$\frac{1}{2}H_2(g) + \frac{1}{2}F_2(g) \rightarrow HF(g)$</td>
<td>$-273$</td>
</tr>
<tr>
<td>$SO_3(g)$</td>
<td>$S(s) + \frac{3}{2}O_2(g) \rightarrow SO_3(g)$</td>
<td>$-396$</td>
</tr>
<tr>
<td>$SF_6(g)$</td>
<td>$S(s) + 3F_2(g) \rightarrow SF_6(g)$</td>
<td>$-1220$</td>
</tr>
</tbody>
</table>
What is the enthalpy of oxygen in its standard state?

A. 0.00 kJ
B. 15.99 kJ
C. 100.0 kJ
D. 8.00 kJ
Two or more thermochemical reactions can be summed to determine the overall enthalpy changes based on what law?

A. Boyle’s law
B. Hess’s law
C. Gay-Lussac’s law
D. law of conservation of energy
Click the mouse button to return to the Chapter Menu.
Section 15.5 Reaction Spontaneity

Objectives

• **Differentiate** between spontaneous and nonspontaneous processes.

• **Explain** how changes in entropy and free energy determine the spontaneity of chemical reactions and other processes.

Review Vocabulary

**vaporization:** the energy-requiring process by which a liquid changes to a gas or vapor
Section 15.5 Reaction Spontaneity (cont.)

New Vocabulary

spontaneous process

entropy

second law of thermodynamics

free energy

MAIN Idea

Changes in enthalpy and entropy determine whether a process is spontaneous.
Spontaneous Processes

- A **spontaneous process** is a physical or chemical change that once begun, occurs with no outside intervention.

- Many spontaneous processes require some energy from the surroundings to start the process.
Spontaneous Processes (cont.)

- **Entropy** is a measure of the possible ways that the energy of a system can be distributed, and this is related to the freedom of the system’s particles to move and the number of ways they can be arranged.
Spontaneous Processes (cont.)

• The **second law of thermodynamics** states that spontaneous processes always proceed in such a way that the entropy of the universe increases.

• Entropy is sometimes considered a measure of disorder or randomness of the particles in a system.

• The more spread out the particles are, the more disorder.
Spontaneous Processes (cont.)

- Entropy changes associated with changes in state can be predicted.
- Entropy increases as a substance changes from a solid to a liquid and from a liquid to a gas.
- Dissolving a gas in a solvent always results in a decrease in entropy.
Spontaneous Processes (cont.)

• Assuming no change in physical state occurs, the entropy of a system usually increases when the number of gaseous product particles is greater than the number of gaseous reactant particles.

• With some exceptions, entropy increases when a solid or liquid dissolves in a solvent.

• The random motion of particles of a substance increases as its temperature increases.
Entropy, the Universe, and Free Energy

• In nature, the change in entropy tends to be positive when:
  − The reaction or process is exothermic, which raises the temperature of the surroundings
  − The entropy of the system increases.
Entropy, the Universe, and Free Energy (cont.)

- **Free energy** is energy that is available to do work.

- Gibbs free energy equation

  \[ \Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}} \]
• If the sign of the energy change, $\Delta G^\circ$, is positive, the reaction is nonspontaneous.

• If the sign of the energy change is negative, the reaction is spontaneous.
Entropy, the Universe, and Free Energy (cont.)

- The entropy of the universe (system + surroundings) always increases in any spontaneous process.

<table>
<thead>
<tr>
<th>Table 15.6</th>
<th>Reaction Spontaneity</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\Delta H_{\text{system}}$</td>
<td>$\Delta S_{\text{system}}$</td>
</tr>
<tr>
<td>negative</td>
<td>positive</td>
</tr>
<tr>
<td>negative</td>
<td>negative</td>
</tr>
<tr>
<td>positive</td>
<td>positive</td>
</tr>
<tr>
<td>positive</td>
<td>negative</td>
</tr>
</tbody>
</table>
Section 15.5 Assessment

What is the measure of randomness in a system?

A. enthalpy
B. entropy
C. free energy
D. Gibbs free energy
Energy that is available to do work is _____.

A. enthalpy
B. entropy
C. random molecular motion
D. free energy
Click the mouse button to return to the Chapter Menu.
Chemistry Online
Study Guide
Chapter Assessment
Standardized Test Practice
Image Bank
Concepts in Motion
Key Concepts

- Energy is the capacity to do work or produce heat.

- Chemical potential energy is energy stored in the chemical bonds of a substance by virtue of the arrangement of the atoms and molecules.

- Chemical potential energy is released or absorbed as heat during chemical processes or reactions.

\[ q = c \times m \times \Delta T \]
Key Concepts

• In thermochemistry, the universe is defined as the system plus the surroundings.

• The heat lost or gained by a system during a reaction or process carried out at constant pressure is called the change in enthalpy ($\Delta H$).

• When $\Delta H$ is positive, the reaction is endothermic. When $\Delta H$ is negative, the reaction is exothermic.
Key Concepts

- A thermochemical equation includes the physical states of the reactants and products and specifies the change in enthalpy.

- The molar enthalpy (heat) of vaporization, $\Delta H_{\text{vap}}$, is the amount of energy required to evaporate one mole of a liquid.

- The molar enthalpy (heat) of fusion, $\Delta H_{\text{fus}}$, is the amount of energy needed to melt one mole of a solid.
Key Concepts

• The enthalpy change for a reaction can be calculated by adding two or more thermochemical equations and their enthalpy changes.

• Standard enthalpies of formation of compounds are determined relative to the assigned enthalpy of formation of the elements in their standard states.

\[ \Delta H_{rxn}^\circ = \sum \Delta H_f^\circ \text{(products)} - \sum \Delta H_f^\circ \text{(reactants)} \]
Key Concepts

- Entropy is a measure of the disorder or randomness of a system.

- Spontaneous processes always result in an increase in the entropy of the universe.

- Free energy is the energy available to do work. The sign of the free energy change indicates whether the reaction is spontaneous.

\[
\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}
\]
Which represents the larger amount of energy?

A. 1 calorie
B. 1000 calories
C. 10 kilocalories
D. they are all equal
What is the universe when using a bomb-calorimeter to measure heat absorbed by a substance in a process?

A. the substance in the calorimeter
B. the calorimeter itself
C. the water in the calorimeter
D. the calorimeter and all its contents
In which example is the $\Delta H$ positive?

A. an ice cube melting

B. condensation forming on cold glass of water

C. molten lava cooling to form solid rock

D. water vapor changing directly into frost on windows
The standard enthalpy of an element in its standard state is _____.

A. 0.00 calories
B. 0.00°C
C. 0.00 kilocalories
D. 0.00 kJ
Spontaneous processes always result in:

A. an increase in entropy
B. a decrease in entropy
C. a decrease in randomness
D. an increase in enthalpy
The specific heat of ethanol is 2.4 J/g. C. How many degrees Celsius can a 50.0g sample be raised with 2400 J of energy?

A. 10°
B. 20°
C. 30°
D. 40°
What is the measure of the number of possible ways that the energy in a system can be distributed?

A. free energy
B. Hess's law
C. enthalpy
D. entropy
Energy stored in chemical bonds is _____.

A. free energy
B. kinetic energy
C. specific heat
D. chemical potential energy
The first law of thermodynamics states:

A. matter is neither created nor destroyed in chemical reactions

B. energy is neither created nor destroyed in chemical reactions

C. spontaneous processes always proceed in such a way that the entropy of the universe increases

D. free energy is random motion of particles
 energy is energy of motion.

A. Free
B. Kinetic
C. Enthalpy
D. Chemical potential
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Bomb Calorimeter

- Ignition terminals
- Thermometer
- Stirrer
- Water
- Insulation
- Sealed reaction chamber containing substance and oxygen (the bomb)
The Heat-Pack Reaction

4Fe(s) + 3O₂(g) → 2Fe₂O₃(s)

ΔH = -1625 kJ

Exothermic Reaction
ΔH < 0

Heat to surroundings
The Cold-Pack Process

**Reactant:** NH₄NO₃(s)

**Products:** NH₄⁺(aq) + NO₃⁻(aq)

ΔH = +27 kJ

Endothermic Process

ΔH > 0

Heat from surroundings
Table 15.3  Standard Enthalpies of Combustion

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>$\Delta H_{\text{comb}}^\circ$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sucrose (table sugar)</td>
<td>C$<em>{12}$H$</em>{22}$O$_{11}$(s)</td>
<td>$-5644$</td>
</tr>
<tr>
<td>Octane (a component of gasoline)</td>
<td>C$<em>{8}$H$</em>{18}$(l)</td>
<td>$-5471$</td>
</tr>
<tr>
<td>Glucose (a simple sugar found in fruit)</td>
<td>C$<em>{6}$H$</em>{12}$O$_{6}$(s)</td>
<td>$-2808$</td>
</tr>
<tr>
<td>Propane (a gaseous fuel)</td>
<td>C$<em>{3}$H$</em>{8}$(g)</td>
<td>$-2219$</td>
</tr>
<tr>
<td>Methane (a gaseous fuel)</td>
<td>CH$_{4}$(g)</td>
<td>$-891$</td>
</tr>
</tbody>
</table>
### Table 15.4

**Standard Enthalpies of Vaporization and Fusion**

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>$\Delta H_{vap}^o$ (kJ/mol)</th>
<th>$\Delta H_{fus}^o$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H$_2$O</td>
<td>40.7</td>
<td>6.01</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C$_2$H$_5$OH</td>
<td>38.6</td>
<td>4.94</td>
</tr>
<tr>
<td>Methanol</td>
<td>CH$_3$OH</td>
<td>35.2</td>
<td>3.22</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>CH$_3$COOH</td>
<td>23.4</td>
<td>11.7</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH$_3$</td>
<td>23.3</td>
<td>5.66</td>
</tr>
</tbody>
</table>
Phase Changes for Water

\[ \Delta H_{\text{vap}} = +40.7 \text{ kJ} \]

\[ \Delta H_{\text{cond}} = -40.7 \text{ kJ} \]

\[ \Delta H_{\text{fus}} = +6.01 \text{ kJ} \]

\[ \Delta H_{\text{solid}} = -6.01 \text{ kJ} \]
The Synthesis of Sulfur Trioxide

Equation c

\[ 2S(s) + 2O_2(g) \]

\[ \Delta H = -594 \text{ kJ} \]

Equation d

\[ 2SO_2(g) + O_2(g) \]

\[ \Delta H = -198 \text{ kJ} \]

Overall energy change

\[ \Delta H = -792 \text{ kJ} \]
### Table 15.5 Standard Enthalpies of Formation

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formation Equation</th>
<th>$\Delta H_f^\circ$(kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{H}_2\text{S}(g)$</td>
<td>$\text{H}_2(g) + \text{S}(s) \rightarrow \text{H}_2\text{S}(g)$</td>
<td>$-21$</td>
</tr>
<tr>
<td>$\text{HF}(g)$</td>
<td>$\frac{1}{2}\text{H}_2(g) + \frac{1}{2}\text{F}_2(g) \rightarrow \text{HF}(g)$</td>
<td>$-273$</td>
</tr>
<tr>
<td>$\text{SO}_3(g)$</td>
<td>$\text{S}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{SO}_3(g)$</td>
<td>$-396$</td>
</tr>
<tr>
<td>$\text{SF}_6(g)$</td>
<td>$\text{S}(s) + 3\text{F}_2(g) \rightarrow \text{SF}_6(g)$</td>
<td>$-1220$</td>
</tr>
</tbody>
</table>
Image Bank

O₂  He

Diagram showing molecular structures with different compositions.
### Table 15.6 Reaction Spontaneity

\[ \Delta G_{\text{system}} = \Delta H_{\text{system}} - T \Delta S_{\text{system}} \]

<table>
<thead>
<tr>
<th>( \Delta H_{\text{system}} )</th>
<th>( \Delta S_{\text{system}} )</th>
<th>( \Delta G_{\text{system}} )</th>
<th>Reaction Spontaneity</th>
</tr>
</thead>
<tbody>
<tr>
<td>negative</td>
<td>positive</td>
<td>always negative</td>
<td>always spontaneous</td>
</tr>
<tr>
<td>negative</td>
<td>negative</td>
<td>negative or positive</td>
<td>spontaneous at lower temperatures</td>
</tr>
<tr>
<td>positive</td>
<td>positive</td>
<td>negative or positive</td>
<td>spontaneous at higher temperatures</td>
</tr>
<tr>
<td>positive</td>
<td>negative</td>
<td>always positive</td>
<td>never spontaneous</td>
</tr>
</tbody>
</table>
Figure 15.5  Calorimetry

Figure 15.10  Heat Flow in Endothermic and Exothermic Reactions

Table 15.6  Reaction Spontaneity
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