

Chapter Summary and Review

Key Learning Outcomes

CHAPTER OBJECTIVES	ASSESSMENT
Calculate Changes in Internal Energy from Heat and Work (9.3)	<ul style="list-style-type: none"> Example 9.1 For Practice 9.1 Exercises 33, 34, 35, 36
Find Heat from Temperature Changes (9.4)	<ul style="list-style-type: none"> Example 9.2 For Practice 9.2 For More Practice 9.2 Exercises 39, 40
Determine Quantities in Thermal Energy Transfer (9.4)	<ul style="list-style-type: none"> Example 9.3 For Practice 9.3 Exercises 57, 58, 59, 60, 61, 62
Calculate Work from Volume Changes (9.4)	<ul style="list-style-type: none"> Example 9.4 For Practice 9.4 For More Practice 9.4 Exercises 43, 44
Use Bomb Calorimetry to Calculate ΔE_{rxn} (9.5)	<ul style="list-style-type: none"> Example 9.5 For Practice 9.5 For More Practice 9.5 Exercises 65, 66
Distinguish between Endothermic and Exothermic Processes (9.6)	<ul style="list-style-type: none"> Example 9.6 For Practice 9.6 Exercises 49, 50
Determine Heat from ΔH and Stoichiometry (9.6)	<ul style="list-style-type: none"> Example 9.7 For Practice 9.7 For More Practice 9.7 Exercises 51, 52, 53, 54, 55, 56
Find ΔH_{rxn} Using Calorimetry (9.7)	<ul style="list-style-type: none"> Example 9.8 For Practice 9.8 Exercises 67, 68
Find ΔH_{rxn} Using Hess's Law (9.8)	<ul style="list-style-type: none"> Example 9.9 For Practice 9.9 For More Practice 9.9 Exercises 71, 72, 73, 74
Find ΔH_{rxn} Using Bond Energies (9.9)	<ul style="list-style-type: none"> Example 9.10 For Practice 9.10 For More Practice 9.10 Exercises 75, 76, 77, 78
Find ΔH_{rxn} Using Standard Enthalpies of Formation (9.10)	<ul style="list-style-type: none"> Examples 9.11, 9.12, 9.13 For Practice 9.11, 9.12, 9.13 Exercises 79, 80, 81, 82, 83, 84, 85, 86
Predict Relative Lattice Energies (9.11)	<ul style="list-style-type: none"> Example 9.14 For Practice 9.14 For More Practice 9.14 Exercises 89, 90, 91, 92

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Key Terms

Section 9.1

thermochemistry

Section 9.2

energy

work

heat (q)

kinetic energy

thermal energy

potential energy

chemical energy

law of conservation of energy

system

surroundings

Section 9.3

thermodynamics

first law of thermodynamics

internal energy (E)

state function

Section 9.4

thermal equilibrium

heat capacity (C)

specific heat capacity (C_s)

molar heat capacity

pressure-volume work

Section 9.5

calorimetry

bomb calorimeter

Section 9.6

enthalpy (H)

endothermic reaction

exothermic reaction

enthalpy (heat) of reaction (ΔH_{rxn})

Section 9.7

coffee-cup calorimeter

Section 9.8

Hess's law

Section 9.10

standard state

standard enthalpy change (ΔH°)

standard enthalpy (or heat) of formation (ΔH_f°)

standard heat of formation

Section 9.11

Born–Haber cycle

Key Concepts

The Nature of Energy and Thermodynamics (9.1–9.3)

- Energy, which is measured in the SI unit of joules (J), is the capacity to do work.
- Work is the result of a force acting through a distance.
- Many different kinds of energy exist, including kinetic energy, thermal energy, potential energy, and chemical energy, a type of potential energy associated with the relative positions of electrons and nuclei in atoms and molecules.
- The first law of thermodynamics states that energy can be converted from one form to another, but the total amount of energy is always conserved.
- The internal energy (E) of a system is the sum of all of its kinetic and potential energy. Internal energy is a state function, which means that it depends only on the state of the system and not on the pathway by which it got to that state.
- A chemical system exchanges energy with its surroundings through heat (the transfer of thermal energy caused by a temperature difference) or work. The total change in internal energy is the sum of these two quantities.
- The change in internal energy (ΔE) that occurs during a chemical reaction is the sum of the heat (q) exchanged and the work (w) done: $\Delta E = q + w$.

Heat and Work (9.4)

- We quantify heat with the equation $q = m \times C_s \times \Delta T$. In this expression, C_s is the specific heat capacity, the amount of heat required to change the temperature of 1 g of the substance 1 °C. Compared to most substances, water has a very high heat capacity—it takes a lot of heat to change its temperature.
- The type of work most characteristic of chemical reactions is pressure–volume work, which occurs when a gas expands against an external pressure. Pressure–volume work can be quantified with the equation $w = -P\Delta V$.

Enthalpy (9.6)

- The heat evolved in a chemical reaction occurring at constant pressure is the change in enthalpy (ΔH) for the reaction. Like internal energy, enthalpy is a state function.
- An endothermic reaction has a positive enthalpy of reaction; an exothermic reaction has a negative enthalpy of reaction.
- The enthalpy of reaction can be used to determine stoichiometrically the heat evolved when a specific amount of reactant reacts.

Calorimetry (9.5, 9.7)

- Calorimetry is a method of measuring ΔE or ΔH for a reaction.
- In bomb calorimetry, the reaction is carried out under conditions of *constant volume*, so the energy is released only as heat ($\Delta E = q_v$). The temperature change of the calorimeter can therefore be used to calculate ΔE for the reaction.
- When a reaction takes place at *constant pressure*, energy may be released both as heat and as work.
- In coffee-cup calorimetry, a reaction is carried out under atmospheric pressure in a solution, so $q = \Delta H$. The

temperature change of the solution is then used to calculate ΔH for the reaction.

Calculating ΔH_{rxn} (9.8–9.10)

- We can calculate the enthalpy of reaction (ΔH_{rxn}) from known thermochemical data using the following relationships: (a) when a reaction is multiplied by a factor, ΔH_{rxn} is multiplied by the same factor; (b) when a reaction is reversed, ΔH_{rxn} changes sign; and (c) if a chemical reaction can be expressed as a sum of two or more steps, ΔH_{rxn} is the sum of the ΔH values for the individual steps (Hess's law). We can use these relationships together to determine the enthalpy change of an unknown reaction from reactions with known enthalpy changes.
- We can also calculate the enthalpy of reaction (ΔH_{rxn}) from the bond energies of the reactants and products. The ΔH_{rxn} is the sum of the ΔH values for bonds broken (which are positive) plus the sum of the ΔH values for the bond formed (which are negative).
- A third way to calculate ΔH_{rxn} from known thermochemical data involves using tabulated standard enthalpies of formation for the reactants and products of the reaction. These are usually tabulated for substances in their standard states, and the enthalpy of reaction is called the standard enthalpy of reaction ($\Delta H_{\text{rxn}}^\circ$). We can obtain $\Delta H_{\text{rxn}}^\circ$ by subtracting the sum of the enthalpies of formation of the reactants multiplied by their stoichiometric coefficients from the sum of the enthalpies of formation of the products multiplied by their stoichiometric coefficients.

Lattice Energy (9.11)

- The formation of most ionic compounds is exothermic because of lattice energy, the energy released when metal cations and nonmetal anions coalesce to form the solid. The smaller the radius of the ions and the greater their charge, the more exothermic the lattice energy.

Key Equations and Relationships

Change in Internal Energy (ΔE) of a Chemical System (9.3)

$$\Delta E = E_{\text{products}} - E_{\text{reactants}}$$

Energy Flow between System and Surroundings (9.3)

$$\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$$

Relationship between Internal Energy (ΔE), Heat (q), and Work (w) (9.3)

$$\Delta E = q + w$$

Relationship between Heat (q), Temperature (T), and Heat Capacity (C) (9.4)

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

Relationship between Heat (q), Mass (m), Temperature (T), and Specific Heat Capacity of a Substance (C_s) (9.4)

$$q = m \times C_s \times \Delta T$$

Relationship between Work (w), Force (F), and Distance (D) (9.4)

$$w = F \times D$$

Relationship between Work (w), Pressure (P), and Change in Volume (ΔV) (9.4)

$$w = -P\Delta V$$

Change in Internal Energy (ΔE) of System at Constant Volume (9.5)

$$\Delta E = q_v$$

Heat of a Bomb Calorimeter (q_{cal}) (9.5□)

$$q_{\text{cal}} = C_{\text{cal}} \times \Delta T$$

Heat Exchange between a Calorimeter and a Reaction (9.5□)

$$q_{\text{cal}} = -q_{\text{rxn}}$$

Relationship between Enthalpy (ΔH), Internal Energy (ΔE), Pressure (P), and Volume (V) (9.6□)

$$\begin{aligned}\Delta H &= \Delta E + P\Delta V \\ \Delta H &= q_p\end{aligned}$$

Enthalpy Change of a Reaction (ΔH_{rxn}): Relationship of Bond Energies (9.9□)

$$\Delta H_{\text{rxn}} = \Sigma (\Delta H \text{ s bonds broken}) + \Sigma (\Delta H \text{ s bonds formed})$$

Relationship between Enthalpy of a Reaction ($\Delta H_{\text{rxn}}^\circ$) and the Heats of Formation (ΔH_f°) (9.10□)

$$\Delta H_{\text{rxn}}^\circ = \Sigma n_p \Delta H_f^\circ (\text{products}) - \Sigma n_r \Delta H_f^\circ (\text{reactants})$$

Not for Distribution

Not for Distribution