

Chapter Summary and Review

Key Learning Outcomes

CHAPTER OBJECTIVES	ASSESSMENT
Convert between Pressure Units (10.3)	<ul style="list-style-type: none"> Example 10.1 For Practice 10.1 For More Practice 10.1 Exercises 25, 26, 27, 28
Relate Volume and Pressure: Boyle's Law (10.4)	<ul style="list-style-type: none"> Example 10.2 For Practice 10.2 Exercises 31, 32
Relate Volume and Temperature: Charles's Law (10.4)	<ul style="list-style-type: none"> Example 10.3 For Practice 10.3 Exercises 33, 34
Relate Volume and Moles: Avogadro's Law (10.4)	<ul style="list-style-type: none"> Example 10.4 For Practice 10.4 Exercises 35, 36
Determine P, V, n, or T Using the Ideal Gas Law (10.5)	<ul style="list-style-type: none"> Examples 10.5, 10.6 For Practice 10.5, 10.6 For More Practice 10.6 Exercises 37, 38, 39, 40, 41, 42, 43, 44, 45, 46, 51, 52
Relate the Density of a Gas to Its Molar Mass (10.6)	<ul style="list-style-type: none"> Example 10.7 For Practice 10.7 For More Practice 10.7 Exercises 55, 56
Calculate the Molar Mass of a Gas with the Ideal Gas Law (10.6)	<ul style="list-style-type: none"> Example 10.8 For Practice 10.8 Exercises 57, 58, 59, 60
Calculate Total Pressure, Partial Pressures, and Mole Fractions of Gases in a Mixture (10.7)	<ul style="list-style-type: none"> Examples 10.9, 10.10, 10.11 For Practice 10.9, 10.10, 10.11 Exercises 61, 62, 65, 67, 68, 70
Calculate the Root Mean Square Velocity of a Gas (10.8)	<ul style="list-style-type: none"> Example 10.12 For Practice 10.12 Exercises 73, 74
Calculate the Effusion Rate or the Ratio of Effusion Rates of Two Gases (10.9)	<ul style="list-style-type: none"> Example 10.13 For Practice 10.13 Exercises 75, 76, 77, 78
Relate the Amounts of Reactants and Products in Gaseous Reactions: Stoichiometry (10.10)	<ul style="list-style-type: none"> Examples 10.14, 10.15 For Practice 10.14, 10.15 For More Practice 10.14 Exercises 81, 82, 83, 84, 85, 86, 87

Key Terms

Section 10.1

pressure

Section 10.2

kinetic molecular theory

Section 10.3

millimeter of mercury (mmHg)

barometer

torr

atmosphere (atm)

pascal (Pa)

manometer

Section 10.4

Boyle's law

Charles's law

Avogadro's law

Section 10.5

ideal gas law

ideal gas

ideal gas constant

Section 10.6

molar volume

standard temperature and pressure (STP)

Section 10.7

partial pressure (P_n)

Dalton's law of partial pressures

mole fraction (χ_n)

hypoxia

oxygen toxicity

nitrogen narcosis

vapor pressure

Section 10.9

mean free path

diffusion

effusion

Graham's law of effusion

Section 10.11

van der Waals equation

Key Concepts

Kinetic Molecular Theory (10.2)

- Kinetic molecular theory is a quantitative model of gas behavior. The theory has three main assumptions: (1) the gas particles are negligibly small; (2) the average kinetic energy of gas particles is proportional to the temperature in kelvins; and (3) the collision of one gas particle with another is completely elastic (the particles do not stick together). The gas laws all follow from the kinetic molecular theory.

Pressure (10.3)

- Gas pressure is the force per unit area that results from gas particles colliding with the surfaces around them. We use a variety of units to measure pressure, including mmHg, torr, Pa, psi, in Hg, and atm.

The Simple Gas Laws (10.4)

- The simple gas laws express relationships between pairs of variables when other variables are held constant. These variables correspond to the four basic properties of a gas sample: pressure (P), volume (V), temperature (T), and amount in moles (n).
- Boyle's law states that the volume of a gas is inversely proportional to its pressure.
- Charles's law states that the volume of a gas is directly proportional to its temperature.
- Avogadro's law states that the volume of a gas is directly proportional to the amount (in moles).

The Ideal Gas Law (10.5, 10.6)

- The ideal gas law, $PV = nRT$ relates the relationship among all four gas variables and contains the simple gas laws within it.
- We can use the ideal gas law to find one of the four variables if we know the other three. We can use it to calculate the molar volume of an ideal gas, which is 22.4 L at STP, and to calculate the density and molar mass of a gas.

Mixtures of Gases and Partial Pressures (10.7)

- In a mixture of gases, each gas acts independently of the others so that any overall property of the mixture is the sum of the properties of the individual components.
- The pressure due to any individual component is its partial pressure.

Molecular Velocities and Effusion (10.8, 10.9)

- We can use kinetic molecular theory to derive the expression for the root mean square velocity of gas particles. This velocity is inversely proportional to the molar mass of the gas, and therefore—at a given temperature—smaller gas particles (on average) move more quickly than larger ones.
- The kinetic molecular theory also allows us to predict the mean free path of a gas particle (the distance it travels between collisions) and relative rates of diffusion or effusion.

Gas Stoichiometry (10.10)

- In reactions involving gaseous reactants and products, we often report quantities in volumes at specified pressures and temperatures. We can convert these quantities to amounts (in moles) using the ideal gas law. Then we can use the stoichiometric coefficients from the balanced equation to determine the stoichiometric amounts of other reactants or products.
- The general form for these types of calculations is: volume A \rightarrow amount A (in moles) \rightarrow amount B (in moles) \rightarrow quantity of B (in desired units).
- In cases where the reaction is carried out at STP, we can use the molar volume at STP (22.4 L = 1 mol) to convert between volume in liters and amount in moles.

Real Gases (10.11)

- Real gases, unlike ideal gases, do not always fit the assumptions of kinetic molecular theory.
- These assumptions tend to break down at high pressures, where the volume is higher than predicted for an ideal gas because the particles are no longer negligibly small compared to the space between them.
- The assumptions also break down at low temperatures where the pressure is lower than predicted because the attraction between molecules combined with low kinetic energies causes partially inelastic collisions.
- We can use the van der Waals equation to predict gas properties under nonideal conditions.

Key Equations and Relationships

Relationship between Pressure (P), Force (F), and Area (A) (10.3)

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

Boyle's Law: Relationship between Pressure (P) and Volume (V) (10.4)

$$V \propto \frac{1}{P}$$

$$P_1 V_1 = P_2 V_2$$

Charles's Law: Relationship between Volume (V) and Temperature (T) (10.4)

$$V \propto T \text{ (in K)}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Avogadro's Law: Relationship between Volume (V) and Amount in Moles (n) (10.4)

$$V \propto n$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Ideal Gas Law: Relationship between Volume (V), Pressure (P), Temperature (T), and Amount (n) (10.5)

$$PV = nRT$$

Density of a Gas (10.6)

$$d = \frac{PM}{RT}$$

Dalton's Law: Relationship between Partial Pressures P_n in Mixture of Gases and Total Pressure (P_{total}) (10.7)

$$P_{\text{total}} = P_a + P_b + P_c + \dots$$

$$P_a = \frac{n_a RT}{V} \quad P_b = \frac{n_b RT}{V} \quad P_c = \frac{n_c RT}{V}$$

Mole Fraction (χ_a) (10.7)

$$\chi_a = \frac{n_a}{n_{\text{total}}}$$

$$P_a = \chi_a P_{\text{total}}$$

Average Kinetic Energy (KE_{avg}) (10.8)

$$\text{KE}_{\text{avg}} = \frac{3}{2} RT$$

Relationship between Root Mean Square Velocity (u_{rms}) and Temperature (T) (10.8)

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$$

Relationship of Effusion Rates of Two Different Gases (10.9)

$$\frac{\text{rate A}}{\text{rate B}} = \sqrt{\frac{M_B}{M_A}}$$

Van der Waals Equation: The Effects of Volume and Intermolecular Forces on Nonideal Gas Behavior (10.11)

$$\left[P + a(n/V)^2 \right] \times (V - nb) = nRT$$

Not for Distribution

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