

Homework 11: Gas Laws and Kinetic Molecular Theory

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October 2025

1 Quiz Details

- **When:** Tuesday, October 28th, 7:00 AM - 7:30 AM.
- **Format:** Closed notes.
- **Allowed Materials:** A scientific calculator. Phones are not permitted.

2 Key Topics

The quiz will cover all aspects of gas behavior discussed in Chapter 11. You must have a strong command of the following 13 topics:

1. Pressure Unit Conversions
2. The Empirical Gas Laws (Boyle's, Charles's, Gay-Lussac's)
3. Proportionality of Gas Parameters
4. The Combined Gas Law
5. Standard Temperature and Pressure (STP)
6. The Ideal Gas Law ($PV = nRT$)
7. Molar Mass and Density Calculations
8. Gas Stoichiometry
9. Partial Pressures, Total Pressure, and Mole Fraction
10. Gas Collection Over Water
11. Kinetic Molecular Theory: Speed vs. Molar Mass
12. Diffusion and Effusion Rates
13. Calculating Ratios of Molecular Speeds (Graham's Law)

3 Phased Study Plan

3.1 Phase 1: Foundation (Friday/Saturday)

- **Core Concepts Review:** Read through `lecture_slides.pdf`, focusing on slides 7-23 (Gas Laws), slides 31-43 (Gas Mixtures and Collection), and slides 45-50 (KMT and Graham's Law).
- **Memorize Key Formulas:**

- **Boyle’s Law:** $P_1V_1 = P_2V_2$ (constant n, T)
- **Charles’s Law:** $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (constant n, P)
- **Gay-Lussac’s Law:** $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ (constant n, V)
- **Combined Gas Law:** $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ (constant n)
- **Ideal Gas Law:** $PV = nRT$ (where $R = 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$)
- **Molar Mass from Ideal Gas Law:** $M = \frac{dRT}{P}$ where d is density.
- **Dalton’s Law:** $P_{total} = P_A + P_B + P_C + \dots$
- **Mole Fraction and Partial Pressure:** $P_A = \chi_A P_{total}$ where $\chi_A = \frac{n_A}{n_{total}}$
- **Graham’s Law of Effusion:** $\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{M_B}{M_A}}$

- **Conceptual Understanding:**

- Temperature (T) must **ALWAYS** be in Kelvin (K) for all gas law calculations. $K = {}^\circ\text{C} + 273.15$.
- Standard Temperature and Pressure (STP) is defined as exactly 273.15 K (0°C) and 1 atm. At STP, 1 mole of any ideal gas occupies 22.4 L.
- Lighter gases (smaller molar mass) move faster, diffuse faster, and effuse faster than heavier gases at the same temperature.

3.2 Phase 2: Practice & Application (Sunday)

- **Problem Drills:** Redo problems from the homework solutions. Do not just read them.
 - Gas Law Problems.pdf: #1-7
 - Molar mass, density, gas stoichiometry, gas mixtures.pdf: #1-7
 - Kinetic molecular theory.pdf: #1-7
- **Unit Conversion Practice:** Master these conversions.
 - Pressure: mmHg to atm (760 mmHg = 1 atm), torr to atm (760 torr = 1 atm).
 - Temperature: Celsius to Kelvin ($K = {}^\circ\text{C} + 273.15$).

3.3 Phase 3: Final Mastery (Monday)

- **Final Review & Prep:** Briefly review all formulas and key concepts. Explain the KMT graphs to yourself one last time. Get a good night’s sleep.

4 Explanations of Practice Problems

5 Topic 1: Pressure Unit Conversions

5.1 Key Principles

- The standard unit for pressure in the Ideal Gas Law is the atmosphere (atm). Most problems will require you to convert other units, like mmHg or torr, into atm.
- **Memorization Required:** You must know the key conversion factors.
 - 1 atm = 760 mmHg
 - 1 atm = 760 torr
 - Therefore, 1 mmHg = 1 torr.

5.2 Worked Examples

5.2.1 Example 1

Problem: Convert 684 mmHg to atm. **Source:** lecture_slides.pdf, slide 6, Problem A. **Analysis:** This is a direct conversion using the factor 1 atm = 760 mmHg. We will set up the dimensional analysis so that mmHg cancels out. **Step-by-Step Solution:**

$$684 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.900 \text{ atm}$$

Answer: 0.900 atm.

5.2.2 Example 2

Problem: Convert 788 torr to atm. **Source:** lecture_slides.pdf, slide 6, Problem C. **Analysis:** This is a direct conversion using the factor 1 atm = 760 torr. **Step-by-Step Solution:**

$$788 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.0368 \text{ atm}$$

Answer: 1.04 atm.

5.2.3 Example 3

Problem: Convert 0.977 atm to torr. **Source:** lecture_slides.pdf, slide 6, Problem B. **Analysis:** This is a conversion from atm to torr. We use the same conversion factor, but inverted. **Step-by-Step Solution:**

$$0.977 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 742.52 \text{ torr}$$

Answer: 743 torr.

5.2.4 Example 4

Problem: The pressure on planet Quibblefritz is calculated to be 718 torr. What is this pressure in atm?

Source: Adapted from Gas Law Problems.pdf, Problem 6. **Analysis:** A simple conversion from torr to atm. **Step-by-Step Solution:**

$$718 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.9447 \text{ atm}$$

Answer: 0.945 atm.

6 Topic 2: The Empirical Gas Laws

6.1 Key Principles

- These laws relate two gas properties while holding the other two constant. They are used to compare the initial (1) and final (2) states of a gas.
- **Boyle's Law** ($P_1V_1 = P_2V_2$): At constant temperature and moles, pressure and volume are inversely proportional. As pressure increases, volume decreases.
- **Charles's Law** ($\frac{V_1}{T_1} = \frac{V_2}{T_2}$): At constant pressure and moles, volume and absolute temperature are directly proportional. As temperature increases, volume increases.
- **Gay-Lussac's Law** ($\frac{P_1}{T_1} = \frac{P_2}{T_2}$): At constant volume and moles, pressure and absolute temperature are directly proportional. As temperature increases, pressure increases.

6.2 Worked Examples

6.2.1 Example 1 (Gay-Lussac's Law)

Problem: A gas sample in a sealed flask has a pressure of 792.84 torr at a temperature of 21.26°C. What is the pressure if the flask is heated to 75.35°C? **Source:** Gas Law Problems.pdf, Page 1, Problem 1. **Analysis:** The flask is "sealed" (constant n) and "rigid" (constant V). Pressure and temperature are changing. This is a Gay-Lussac's Law problem. We must convert temperatures to Kelvin first. **Step-by-Step Solution:**

1. Identify variables: $P_1 = 792.84$ torr, $T_1 = 21.26^\circ\text{C}$, $P_2 = ?$, $T_2 = 75.35^\circ\text{C}$.
2. Convert temperatures to Kelvin:
 - $T_1 = 21.26 + 273.15 = 294.41$ K
 - $T_2 = 75.35 + 273.15 = 348.50$ K
3. Rearrange Gay-Lussac's Law to solve for P_2 : $P_2 = \frac{P_1 T_2}{T_1}$.
4. Substitute values and calculate:

$$P_2 = \frac{(792.84 \text{ torr})(348.50 \text{ K})}{294.41 \text{ K}} = 938.50 \text{ torr}$$

Answer: 938.50 torr.

6.2.2 Example 2 (Boyle's Law)

Problem: A gas sample in a syringe has a volume of 35.26 mL at 0.9854 atm. The plunger is depressed, reducing the volume to 22.33 mL. What is the new pressure, assuming constant temperature? **Source:** Gas Law Problems.pdf, Page 1, Problem 2. **Analysis:** Temperature and moles are constant, while pressure and volume change. This is a Boyle's Law problem. **Step-by-Step Solution:**

1. Identify variables: $P_1 = 0.9854$ atm, $V_1 = 35.26$ mL, $P_2 = ?$, $V_2 = 22.33$ mL.
2. Rearrange Boyle's Law to solve for P_2 : $P_2 = \frac{P_1 V_1}{V_2}$.
3. Substitute values and calculate:

$$P_2 = \frac{(0.9854 \text{ atm})(35.26 \text{ mL})}{22.33 \text{ mL}} = 1.556 \text{ atm}$$

Answer: 1.556 atm.

6.2.3 Example 3 (Charles's Law)

Problem: A sample of nitrogen in a flexible container has a volume of 42.748 L at -24.36°C. The temperature is increased to 8.97°C. Assuming pressure remains constant, what is the new volume? **Source:** Gas Law Problems.pdf, Page 1, Problem 3. **Analysis:** A "flexible container" at constant pressure implies that volume can change. Moles are also constant. Volume and temperature are changing. This is a Charles's Law problem. **Step-by-Step Solution:**

1. Identify variables: $V_1 = 42.748$ L, $T_1 = -24.36^\circ\text{C}$, $V_2 = ?$, $T_2 = 8.97^\circ\text{C}$.
2. Convert temperatures to Kelvin:
 - $T_1 = -24.36 + 273.15 = 248.79$ K
 - $T_2 = 8.97 + 273.15 = 282.12$ K
3. Rearrange Charles's Law to solve for V_2 : $V_2 = \frac{V_1 T_2}{T_1}$.

4. Substitute values and calculate:

$$V_2 = \frac{(42.748 \text{ L})(282.12 \text{ K})}{248.79 \text{ K}} = 48.456 \text{ L}$$

Answer: The provided key shows 14.456 L, which corresponds to an initial volume of 12.748 L, likely a typo in my source file. Based on the problem as written, the answer is **48.456 L**.

6.2.4 Example 4 (Boyle's Law)

Problem: A sample of air occupies 3.8 L when the pressure is 1.2 atm. What volume does it occupy at 6.6 atm, assuming constant temperature? **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.32a. **Analysis:** Constant temperature and amount of gas, with changing pressure and volume. This is a Boyle's Law problem. **Step-by-Step Solution:**

1. Identify variables: $P_1 = 1.2 \text{ atm}$, $V_1 = 3.8 \text{ L}$, $P_2 = 6.6 \text{ atm}$, $V_2 = ?$.
2. Rearrange Boyle's Law to solve for V_2 : $V_2 = \frac{P_1 V_1}{P_2}$.
3. Substitute values and calculate:

$$V_2 = \frac{(1.2 \text{ atm})(3.8 \text{ L})}{6.6 \text{ atm}} = 0.69 \text{ L}$$

Answer: 0.69 L.

7 Topic 3: Proportionality of Gas Parameters

7.1 Key Principles

- The empirical gas laws are based on proportionalities between variables.
- **Direct Proportionality ($X \propto Y$):** As Y increases, X increases by the same factor (e.g., volume and temperature in Charles's Law). The graph is a straight line through the origin.
- **Inverse Proportionality ($X \propto 1/Y$):** As Y increases, X decreases by the same factor (e.g., pressure and volume in Boyle's Law). The graph is a hyperbola.
- These relationships are key to conceptually understanding gas behavior without calculations.

7.2 Worked Examples

7.2.1 Example 1

Problem: Explain why a helium weather balloon expands as it rises in the air, assuming the temperature remains constant. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.29. **Analysis:** This question asks for a conceptual explanation based on gas law proportionalities. As a balloon rises, the external atmospheric pressure decreases. We need to explain how this affects the balloon's volume.

Step-by-Step Solution:

1. Identify the relationship: At constant temperature and moles of gas, pressure and volume are related by Boyle's Law.
2. State the proportionality: Boyle's Law states that pressure and volume are inversely proportional ($V \propto 1/P$).
3. Apply to the scenario: As the balloon rises, the external pressure decreases. To maintain pressure equilibrium, the internal pressure must also decrease. Because of the inverse relationship, a decrease in pressure causes an increase in volume.

Answer: As the balloon rises, the atmospheric pressure decreases. According to Boyle's Law, pressure and volume are inversely proportional, so the volume of the helium inside the balloon expands to match the lower external pressure.

7.2.2 Example 2

Problem: The absolute temperature of a gas in a flexible container is doubled at constant pressure and amount of gas. What happens to the volume? **Source:** Adapted from Questions and Problems (textbook).pdf, Chapter 11, Problem 11.40(2). **Analysis:** This requires understanding the direct proportionality from Charles's Law. **Step-by-Step Solution:**

1. Identify the relationship: At constant pressure and moles, volume and temperature are related by Charles's Law.
2. State the proportionality: Charles's Law states that volume and absolute temperature are directly proportional ($V \propto T$).
3. Apply to the scenario: If the absolute temperature is doubled, the volume must also double to maintain the direct proportionality.

Answer: The volume will double.

7.2.3 Example 3

Problem: The pressure on a gas in a piston is tripled at constant temperature and amount of gas. What happens to the volume? **Source:** Adapted from Questions and Problems (textbook).pdf, Chapter 11, Problem 11.40(1). **Analysis:** This requires understanding the inverse proportionality from Boyle's Law. **Step-by-Step Solution:**

1. Identify the relationship: At constant temperature and moles, pressure and volume are related by Boyle's Law.
2. State the proportionality: Pressure and volume are inversely proportional ($V \propto 1/P$).
3. Apply to the scenario: If the pressure is tripled, the volume must decrease by a factor of three (be reduced to one-third of its original value) to maintain the inverse relationship.

Answer: The volume is reduced to one-third of its original value.

7.2.4 Example 4

Problem: If a sealed, rigid container of gas is heated, what happens to the pressure inside? **Source:** Conceptual Question. **Analysis:** A "sealed, rigid container" means moles (n) and volume (V) are constant. We are heating it, so temperature (T) increases. We need to describe the effect on pressure (P). **Step-by-Step Solution:**

1. Identify the relationship: At constant volume and moles, pressure and temperature are related by Gay-Lussac's Law.
2. State the proportionality: Pressure and absolute temperature are directly proportional ($P \propto T$).
3. Apply to the scenario: As the temperature increases, the pressure must also increase.

Answer: The pressure increases.

8 Topic 4: The Combined Gas Law

8.1 Key Principles

- The Combined Gas Law merges Boyle's, Charles's, and Gay-Lussac's laws into a single, more versatile equation.
- It is used when comparing two states (initial and final) of a gas where the number of moles (n) is constant, but pressure, volume, and temperature may all change.

- The formula is:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

- If any variable is held constant (e.g., $T_1 = T_2$), it can be canceled from both sides, simplifying the equation to one of the empirical laws (in this case, Boyle's Law).

8.2 Worked Examples

8.2.1 Example 1

Problem: A 1.500 L balloon is filled with helium at 305.9 K and 745.9 mmHg. The balloon rises to an altitude where the pressure is 154.2 mmHg and the temperature is 250.6 K. What is the volume of the balloon at this altitude? **Source:** Gas Law Problems.pdf, Page 2, Problem 5. **Analysis:** The amount of gas (n) is constant, but P, V, and T all change. This requires the Combined Gas Law. Pressures are both in mmHg, so they will cancel and do not need to be converted to atm. **Step-by-Step Solution:**

- Identify variables:

- Initial (1): $P_1 = 745.9$ mmHg, $V_1 = 1.500$ L, $T_1 = 305.9$ K
- Final (2): $P_2 = 154.2$ mmHg, $V_2 = ?$, $T_2 = 250.6$ K

- Rearrange the Combined Gas Law to solve for V_2 : $V_2 = \frac{P_1V_1T_2}{T_1P_2}$.

- Substitute values and calculate:

$$V_2 = \frac{(745.9 \text{ mmHg})(1.500 \text{ L})(250.6 \text{ K})}{(305.9 \text{ K})(154.2 \text{ mmHg})} = 5.944 \text{ L}$$

Answer: 5.944 L.

8.2.2 Example 2

Problem: A gas-filled balloon having a volume of 2.50 L at 1.2 atm and 20°C is allowed to rise to the stratosphere, where the temperature and pressure are -23°C and 3.00×10^{-3} atm, respectively. Calculate the final volume of the balloon. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.48. **Analysis:** This is another Combined Gas Law problem. We must convert both temperatures to Kelvin before solving for the final volume, V_2 . **Step-by-Step Solution:**

- Identify variables:

- Initial (1): $P_1 = 1.2$ atm, $V_1 = 2.50$ L, $T_1 = 20^\circ\text{C}$
- Final (2): $P_2 = 3.00 \times 10^{-3}$ atm, $V_2 = ?$, $T_2 = -23^\circ\text{C}$

- Convert temperatures to Kelvin:

- $T_1 = 20 + 273.15 = 293.15$ K
- $T_2 = -23 + 273.15 = 250.15$ K

- Rearrange the Combined Gas Law for V_2 : $V_2 = \frac{P_1V_1T_2}{T_1P_2}$.

- Substitute values and calculate:

$$V_2 = \frac{(1.2 \text{ atm})(2.50 \text{ L})(250.15 \text{ K})}{(293.15 \text{ K})(3.00 \times 10^{-3} \text{ atm})} = 853.5 \text{ L}$$

Answer: 8.5×10^2 L.

8.2.3 Example 3

Problem: A gas at 572 mmHg and 35.0°C occupies a volume of 6.15 L. Calculate its volume at STP.
Source: Questions and Problems (textbook).pdf, Chapter 11, Problem 11.52. **Analysis:** This problem requires us to know the conditions for STP. We will use the Combined Gas Law, converting all units to be consistent. **Step-by-Step Solution:**

1. Identify variables:

- Initial (1): $P_1 = 572 \text{ mmHg}$, $V_1 = 6.15 \text{ L}$, $T_1 = 35.0^\circ\text{C}$
- Final (2, STP): $P_2 = 1 \text{ atm}$, $V_2 = ?$, $T_2 = 0^\circ\text{C}$

2. Convert all units to be consistent. Let's use atm and K.

- $P_1 = 572 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.7526 \text{ atm}$
- $T_1 = 35.0 + 273.15 = 308.15 \text{ K}$
- $P_2 = 1.00 \text{ atm}$
- $T_2 = 0 + 273.15 = 273.15 \text{ K}$

3. Rearrange for V_2 : $V_2 = \frac{P_1 V_1 T_2}{T_1 P_2}$.

4. Substitute values:

$$V_2 = \frac{(0.7526 \text{ atm})(6.15 \text{ L})(273.15 \text{ K})}{(308.15 \text{ K})(1.00 \text{ atm})} = 4.10 \text{ L}$$

Answer: 4.10 L.

8.2.4 Example 4

Problem: An ideal gas originally at 0.85 atm and 66°C was allowed to expand until its final volume, pressure, and temperature were 94 mL, 0.60 atm, and 45°C, respectively. What was its initial volume?

Source: Questions and Problems (textbook).pdf, Chapter 11, Problem 11.50. **Analysis:** A Combined Gas Law problem where the unknown is the initial volume, V_1 . **Step-by-Step Solution:**

1. Identify variables:

- Initial (1): $P_1 = 0.85 \text{ atm}$, $V_1 = ?$, $T_1 = 66^\circ\text{C}$
- Final (2): $P_2 = 0.60 \text{ atm}$, $V_2 = 94 \text{ mL}$, $T_2 = 45^\circ\text{C}$

2. Convert temperatures to Kelvin:

- $T_1 = 66 + 273.15 = 339.15 \text{ K}$
- $T_2 = 45 + 273.15 = 318.15 \text{ K}$

3. Rearrange the Combined Gas Law for V_1 : $V_1 = \frac{P_2 V_2 T_1}{T_2 P_1}$.

4. Substitute values:

$$V_1 = \frac{(0.60 \text{ atm})(94 \text{ mL})(339.15 \text{ K})}{(318.15 \text{ K})(0.85 \text{ atm})} = 70.7 \text{ mL}$$

Answer: 71 mL.

9 Topic 5: Standard Temperature and Pressure (STP)

9.1 Key Principles

- **Memorization Required:** STP is a specific set of reference conditions for gases.
 - Standard Temperature = 0°C or 273.15 K
 - Standard Pressure = 1 atm (exactly)
- At STP, the volume of 1 mole of any ideal gas is 22.4 L. This is called the **molar volume**.
- This molar volume can be used as a direct conversion factor between moles and liters, but **only at STP**.
- For problems not at STP, you must use the Ideal Gas Law ($PV = nRT$).

9.2 Worked Examples

9.2.1 Example 1

Problem: At STP, what is the density of a gas (in g/L) if the molar mass of this gas is 66.1 g/mole?

Source: Molar mass, density, gas stoichiometry, gas mixtures.pdf, Problem 2. **Analysis:** We can use the molar mass and molar volume at STP to find the density. Density is mass/volume. For 1 mole of gas at STP, the mass is the molar mass (66.1 g) and the volume is the molar volume (22.4 L). **Step-by-Step Solution:**

1. Identify the mass of 1 mole of the gas: $m = 66.1$ g.
2. Identify the volume of 1 mole of the gas at STP: $V = 22.4$ L.
3. Calculate density:

$$d = \frac{\text{mass}}{\text{volume}} = \frac{66.1 \text{ g}}{22.4 \text{ L}} = 2.95 \text{ g/L}$$

Answer: 2.95 g/L.

9.2.2 Example 2

Problem: Calculate the volume of 110.4 g CH₄ gas (in L) at STP. **Source:** Molar mass, density, gas stoichiometry, gas mixtures.pdf, Problem 3. **Analysis:** This is a two-step conversion: first convert grams to moles, then use the molar volume at STP to convert moles to liters. **Step-by-Step Solution:**

1. The molar mass of CH₄ is $12.01 + 4(1.008) = 16.042$ g/mol.
2. Set up the dimensional analysis:

$$110.4 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.042 \text{ g CH}_4} \times \frac{22.4 \text{ L}}{1 \text{ mol CH}_4}$$

3. Calculate the result:

$$= 6.882 \text{ mol CH}_4 \times 22.4 \frac{\text{L}}{\text{mol}} = 154.15 \text{ L}$$

Answer: 154.3 L CH₄.

9.2.3 Example 3

Problem: At STP, 0.280 L of a gas weighs 0.400 g. Calculate the molar mass of the gas. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.54. **Analysis:** We can use the molar volume at STP to find out how many moles are in 0.280 L. Then, we can use the definition of molar mass (grams/mole). **Step-by-Step Solution:**

1. Convert the volume at STP to moles:

$$n = 0.280 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.0125 \text{ mol}$$

2. Calculate the molar mass:

$$M = \frac{\text{mass}}{\text{moles}} = \frac{0.400 \text{ g}}{0.0125 \text{ mol}} = 32.0 \text{ g/mol}$$

Answer: 32.0 g/mol.

9.2.4 Example 4

Problem: Calculate the volume (in liters) of 124.3 g of CO₂ at STP. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.51. **Analysis:** This is the same process as Example 2. **Step-by-Step Solution:**

1. The molar mass of CO₂ is 12.01 + 2(16.00) = 44.01 g/mol.
2. Set up the dimensional analysis:

$$124.3 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{22.41 \text{ L}}{1 \text{ mol CO}_2}$$

(Using the more precise molar volume of 22.41 L/mol)

3. Calculate the result:

$$= 2.824 \text{ mol CO}_2 \times 22.41 \frac{\text{L}}{\text{mol}} = 63.29 \text{ L}$$

Answer: 63.29 L.

10 Topic 6: The Ideal Gas Law

10.1 Key Principles

- The Ideal Gas Law, $PV = nRT$, relates all four gas variables (Pressure, Volume, moles, Temperature) for a gas in a single state. It is the most powerful and versatile of the gas laws.
- **Memorization Required:** You must know the formula and the value and units of the ideal gas constant, R .
 - $R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$
- For the units to cancel correctly, the variables in the equation must have units that match the units in R:
 - P must be in atmospheres (atm).
 - V must be in Liters (L).
 - n must be in moles (mol).
 - T must be in Kelvin (K).

10.2 Worked Examples

10.2.1 Example 1

Problem: A 2.2854 g sample of propane (C_3H_8 ; 44.0972 g/mol) is placed in a 1.5763 L flask at 375.91 K. What is the pressure of this sample? **Source:** Gas Law Problems.pdf, Page 3, Problem 7. **Analysis:** This problem describes a single state of a gas, so we use the Ideal Gas Law. We are asked for pressure (P). We have V and T. We need to find n by converting the given mass to moles. **Step-by-Step Solution:**

1. Convert mass to moles (n):

$$n = 2.2854 \text{ g } C_3H_8 \times \frac{1 \text{ mol } C_3H_8}{44.0972 \text{ g } C_3H_8} = 0.051826 \text{ mol}$$

2. Rearrange the Ideal Gas Law to solve for P : $P = \frac{nRT}{V}$.

3. Substitute all values into the equation:

$$P = \frac{(0.051826 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(375.91 \text{ K})}{1.5763 \text{ L}} = 1.014 \text{ atm}$$

Answer: 1.014 atm.

10.2.2 Example 2

Problem: A 50.0-L cylinder of nitrogen, N_2 , has a pressure of 17.1 atm at 23°C. What is the mass of nitrogen in the cylinder? **Source:** lecture_slides.pdf, slide 22. **Analysis:** This requires a two-step process. First, use the Ideal Gas Law to solve for the number of moles, n . Second, convert the moles of N_2 to grams using its molar mass. **Step-by-Step Solution:**

1. Convert temperature to Kelvin: $T = 23 + 273.15 = 296.15 \text{ K}$.
2. Rearrange the Ideal Gas Law to solve for n : $n = \frac{PV}{RT}$.
3. Substitute values to find moles:

$$n = \frac{(17.1 \text{ atm})(50.0 \text{ L})}{(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(296.15 \text{ K})} = 35.18 \text{ mol}$$

4. Convert moles to grams. The molar mass of N_2 is $2 \times 14.01 = 28.02 \text{ g/mol}$.

$$35.18 \text{ mol } N_2 \times \frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2} = 985.8 \text{ g}$$

Answer: 986 g N_2 .

10.2.3 Example 3

Problem: What volume will 9.8 moles of sulfur hexafluoride (SF_6) gas occupy if the temperature and pressure of the gas are 105°C and 9.4 atm, respectively? **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.46. **Analysis:** A direct application of the Ideal Gas Law, solving for volume (V). **Step-by-Step Solution:**

1. Convert temperature to Kelvin: $T = 105 + 273.15 = 378.15 \text{ K}$.
2. Rearrange the Ideal Gas Law for V : $V = \frac{nRT}{P}$.
3. Substitute values:

$$V = \frac{(9.8 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(378.15 \text{ K})}{9.4 \text{ atm}} = 32.35 \text{ L}$$

Answer: 32 L.

10.2.4 Example 4

Problem: A 3.684 g sample of Freon-22 (CHF_2Cl ; 134.94 g/mol is a typo, it should be 86.47 g/mol) is placed in a 0.4792 L flask and the pressure is 1.539 atm. What is the Celsius temperature of the sample?

Source: Gas Law Problems.pdf, Page 3, Problem 8. **Analysis:** We will use the Ideal Gas Law to solve for Temperature. We must first convert the mass of Freon-22 to moles. The final answer must be converted from Kelvin to Celsius. (Using correct Molar Mass for CHFCl_2 of 86.47 g/mol). **Step-by-Step Solution:**

1. Convert mass to moles (n):

$$n = 3.684 \text{ g } \text{CHF}_2\text{Cl} \times \frac{1 \text{ mol}}{86.47 \text{ g}} = 0.04260 \text{ mol}$$

2. Rearrange the Ideal Gas Law for T : $T = \frac{PV}{nR}$.

3. Substitute values:

$$T = \frac{(1.539 \text{ atm})(0.4792 \text{ L})}{(0.04260 \text{ mol})(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})} = 211.2 \text{ K}$$

4. Convert Kelvin to Celsius: ${}^{\circ}\text{C} = K - 273.15 = 211.2 - 273.15 = -61.95 {}^{\circ}\text{C}$.

Answer: The source document has a calculation error due to the wrong molar mass. The correct answer is **-62.0 °C**.

11 Topic 7: Molar Mass and Density Calculations

11.1 Key Principles

- The Ideal Gas Law can be combined with the definitions of density ($d = m/V$) and moles ($n = m/M$) to derive two very useful equations.
- **Memorization Required: The Formulas**
 - To find Molar Mass from density: $M = \frac{dRT}{P}$
 - To find density from Molar Mass: $d = \frac{MP}{RT}$
- These equations provide a direct link between the macroscopic property of density and the microscopic property of molar mass for a gas.
- As always, ensure all variables are in units consistent with R (atm, L, K).

11.2 Worked Examples

11.2.1 Example 1

Problem: Determine the molar mass of an unknown gas in a container at -50.0 °C and 6.00 atm pressure. The density of this gas is 14.5 g/L. **Source:** Molar mass, density, gas stoichiometry, gas mixtures.pdf, Problem 1. **Analysis:** This is a direct application of the formula $M = \frac{dRT}{P}$. We just need to convert the temperature to Kelvin first. **Step-by-Step Solution:**

1. Convert temperature to Kelvin: $T = -50.0 + 273.15 = 223.15 \text{ K}$.
2. List the known variables: $d = 14.5 \text{ g/L}$, $P = 6.00 \text{ atm}$, $T = 223.15 \text{ K}$.
3. Substitute into the formula:

$$M = \frac{(14.5 \frac{\text{g}}{\text{L}})(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(223.15 \text{ K})}{6.00 \text{ atm}} = 44.3 \text{ g/mol}$$

Answer: 44.3 g/mol.

11.2.2 Example 2

Problem: Calculate the density of hydrogen bromide (HBr) gas in g/L at 733 mmHg and 46°C. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.59. **Analysis:** This requires using the formula $d = \frac{MP}{RT}$. We need to find the molar mass of HBr and convert pressure and temperature to the correct units. **Step-by-Step Solution:**

1. Find the molar mass of HBr: $M = 1.008 + 79.90 = 80.91 \text{ g/mol}$.
2. Convert temperature to Kelvin: $T = 46 + 273.15 = 319.15 \text{ K}$.
3. Convert pressure to atm: $P = 733 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.9645 \text{ atm}$.
4. Substitute into the formula:

$$d = \frac{(80.91 \frac{\text{g}}{\text{mol}})(0.9645 \text{ atm})}{(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(319.15 \text{ K})} = 2.98 \text{ g/L}$$

Answer: 2.98 g/L.

11.2.3 Example 3

Problem: A 2.10-L vessel contains 4.65 g of a gas at 1.00 atm and 27.0°C. What is the molar mass of the gas? **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.58. **Analysis:** We can solve this in two ways: (1) Calculate the density first, then use $M = dRT/P$. (2) Use $PV = nRT$ to find moles, then use $M = m/n$. Let's use the first method. **Step-by-Step Solution:**

1. Calculate the density: $d = \frac{m}{V} = \frac{4.65 \text{ g}}{2.10 \text{ L}} = 2.214 \text{ g/L}$.
2. Convert temperature to Kelvin: $T = 27.0 + 273.15 = 300.15 \text{ K}$.
3. Use the molar mass formula:

$$M = \frac{dRT}{P} = \frac{(2.214 \frac{\text{g}}{\text{L}})(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(300.15 \text{ K})}{1.00 \text{ atm}} = 54.5 \text{ g/mol}$$

Answer: 54.5 g/mol.

11.2.4 Example 4

Problem: A 1.16 g sample of an unknown compound is placed in a 0.250 L flask and heated to 563 K. Upon complete vaporization, the pressure is 1.03 atm. What is the molar mass of the compound? **Source:** Gas Law Problems.pdf, Page 4, Problem 10. **Analysis:** This is another molar mass problem. We can use the modified ideal gas law $M = \frac{mRT}{PV}$ which is derived by substituting $n = m/M$ into $PV = nRT$. **Step-by-Step Solution:**

1. Identify variables: $m = 1.16 \text{ g}$, $R = 0.08206$, $T = 563 \text{ K}$, $P = 1.03 \text{ atm}$, $V = 0.250 \text{ L}$.
2. Substitute into the formula $M = \frac{mRT}{PV}$:

$$M = \frac{(1.16 \text{ g})(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(563 \text{ K})}{(1.03 \text{ atm})(0.250 \text{ L})} = 208 \text{ g/mol}$$

Answer: 208 g/mol.

12 Topic 8: Gas Stoichiometry

12.1 Key Principles

- Gas stoichiometry problems involve chemical reactions where at least one reactant or product is a gas.
- The Ideal Gas Law is the bridge between the macroscopic properties of a gas (P, V, T) and the stoichiometric amount (moles, n).
- The central path is: Quantity of A → Moles of A $\xrightarrow{\text{Mole Ratio}}$ Moles of B → Quantity of B.
- If you are given P, V, and T of a gas, you use $PV = nRT$ to find its moles. If you need to find P, V, or T of a gas, you use stoichiometry to find its moles first, then use $PV = nRT$.
- At STP, you can use the molar volume shortcut ($1 \text{ mol} = 22.4 \text{ L}$).

12.2 Worked Examples

12.2.1 Example 1 (STP)

Problem: How many liters of Cl_2 at STP is required to react with 50.0 grams of Na? The reaction is: $_2\text{Na(s)} + \text{Cl}_2\text{(g)} \longrightarrow _2\text{NaCl(s)}$. **Source:** lecture_slides.pdf, slide 28. **Analysis:** This is a standard stoichiometry problem, with the final step involving a gas at STP. Path: g Na → mol Na → mol Cl_2 → L Cl_2 . **Step-by-Step Solution:**

1. Set up the dimensional analysis:

$$50.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol Na}} \times \frac{22.4 \text{ L Cl}_2}{1 \text{ mol Cl}_2}$$

2. Calculate the result:

$$= 2.175 \text{ mol Na} \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol Na}} \times 22.4 \frac{\text{L}}{\text{mol}} = 1.0875 \text{ mol Cl}_2 \times 22.4 \frac{\text{L}}{\text{mol}} = 24.36 \text{ L}$$

Answer: 24.4 L Cl_2 .

12.2.2 Example 2 (Non-STP)

Problem: What volume of oxygen gas is produced when 22.2 grams of potassium chlorate is decomposed at 26.0°C and a pressure of 735 mmHg? The reaction is $_2\text{KClO}_3\text{(s)} \longrightarrow _2\text{KCl(s)} + _3\text{O}_2\text{(g)}$. Molar mass of KClO_3 is 122.55 g/mol. **Source:** lecture_slides.pdf, slide 30. **Analysis:** This is a gas stoichiometry problem not at STP. Path: g KClO_3 → mol KClO_3 → mol O_2 . Then, use $PV = nRT$ to find the volume of O_2 . **Step-by-Step Solution:**

1. Convert mass of KClO_3 to moles of O_2 :

$$22.2 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} = 0.2717 \text{ mol O}_2$$

2. Convert T and P to the correct units for the Ideal Gas Law:

- $T = 26.0 + 273.15 = 299.15 \text{ K}$
- $P = 735 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.9671 \text{ atm}$

3. Rearrange $PV = nRT$ to solve for V: $V = \frac{nRT}{P}$.

4. Substitute values:

$$V = \frac{(0.2717 \text{ mol})(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(299.15 \text{ K})}{0.9671 \text{ atm}} = 6.90 \text{ L}$$

Answer: 6.90 L O_2 .

12.2.3 Example 3

Problem: What volume of NH₃ (in L) must be used at 244 torr and 35°C to produce 2.3 kg of HCl gas? Reaction: $_2\text{NH}_3(\text{g}) + _3\text{Cl}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + _6\text{HCl}(\text{g})$. **Source:** Molar mass, density, gas stoichiometry, gas mixtures.pdf, Problem 6. **Analysis:** This is a product-to-reactant stoichiometry problem. Path: kg HCl → g HCl → mol HCl → mol NH₃. Then use $PV = nRT$ to find the volume of NH₃. **Step-by-Step Solution:**

1. Convert mass of HCl to moles of NH₃. Molar mass of HCl is 36.46 g/mol.

$$2.3 \text{ kg HCl} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{2 \text{ mol NH}_3}{6 \text{ mol HCl}} = 21.0 \text{ mol NH}_3$$

2. Convert T and P to the correct units:

- $T = 35 + 273.15 = 308.15 \text{ K}$
- $P = 244 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.321 \text{ atm}$

3. Use $V = \frac{nRT}{P}$ to find the volume of NH₃:

$$V = \frac{(21.0 \text{ mol})(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(308.15 \text{ K})}{0.321 \text{ atm}} = 1653 \text{ L}$$

Answer: $1.7 \times 10^3 \text{ L NH}_3$.

12.2.4 Example 4

Problem: If 9.0 L of NO is combined with excess O₂ at STP, what is the volume in liters of the NO₂ produced? Reaction: $_2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow _2\text{NO}_2(\text{g})$. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.86. **Analysis:** According to Avogadro's law, at constant temperature and pressure (like STP), the ratio of volumes is equal to the ratio of moles. We can use the stoichiometric coefficients directly on the volumes. **Step-by-Step Solution:**

1. Identify the mole ratio from the balanced equation: 2 mol NO produces 2 mol NO₂. This is a 1:1 ratio.
2. Apply this ratio to the volumes:

$$9.0 \text{ L NO} \times \frac{2 \text{ L NO}_2}{2 \text{ L NO}} = 9.0 \text{ L NO}_2$$

Answer: 9.0 L.

13 Topic 9: Partial Pressures & Mole Fraction

13.1 Key Principles

- **Dalton's Law of Partial Pressures:** The total pressure of a gas mixture is the sum of the partial pressures of each individual gas. $P_{total} = P_A + P_B + \dots$
- **Partial Pressure (P_A):** The pressure a single gas in the mixture would exert if it were the only gas in the container. It can be calculated using the ideal gas law for that gas alone: $P_A V = n_A R T$.
- **Mole Fraction (χ_A):** The ratio of the moles of one component to the total moles in the mixture. It is a dimensionless quantity. $\chi_A = \frac{n_A}{n_{total}}$.
- The partial pressure of a gas is equal to its mole fraction multiplied by the total pressure: $P_A = \chi_A P_{total}$.

13.2 Worked Examples

13.2.1 Example 1

Problem: A 10.0 L flask contains 1.03 g O₂ and 0.572 g CO₂ at 291.15 K. a) What are the partial pressures of O₂ and CO₂? b) What is the total pressure? c) What is the mole fraction of O₂? **Source:** lecture_slides.pdf, slides 38-40.

Analysis: First, we must convert the mass of each gas to moles. Then, we can use the Ideal Gas Law for each gas individually to find its partial pressure. The total pressure is the sum of the partial pressures. The mole fraction can be found from either the ratio of moles or the ratio of pressures.

Step-by-Step Solution:

1. Convert masses to moles:

- Moles O₂ (n_{O_2}): $1.03 \text{ g} \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.03218 \text{ mol}$
- Moles CO₂ (n_{CO_2}): $0.572 \text{ g} \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 0.01300 \text{ mol}$

2. a) Calculate partial pressures using $P = nRT/V$:

- $P_{O_2} = \frac{(0.03218 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(291.15 \text{ K})}{10.0 \text{ L}} = 0.0769 \text{ atm}$
- $P_{CO_2} = \frac{(0.01300 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(291.15 \text{ K})}{10.0 \text{ L}} = 0.0311 \text{ atm}$

3. b) Calculate total pressure:

$$P_{total} = P_{O_2} + P_{CO_2} = 0.0769 \text{ atm} + 0.0311 \text{ atm} = 0.1080 \text{ atm}$$

4. c) Calculate mole fraction of O₂:

$$\chi_{O_2} = \frac{P_{O_2}}{P_{total}} = \frac{0.0769 \text{ atm}}{0.1080 \text{ atm}} = 0.712$$

(Alternatively, using moles: $n_{total} = 0.03218 + 0.01300 = 0.04518 \text{ mol}$. $\chi_{O_2} = \frac{0.03218}{0.04518} = 0.712$)

Answer: a) $P_{O_2} = 0.0769 \text{ atm}$, $P_{CO_2} = 0.0311 \text{ atm}$; b) $P_{total} = 0.1080 \text{ atm}$; c) $\chi_{O_2} = 0.712$.

13.2.2 Example 2

Problem: A cylinder has a volume of 20.0 L and contains 1813 g of CH₄ and 336 g of O₂ at 22.0°C. Calculate the partial pressure of each gas and the total pressure. **Source:** Molar mass, density, gas stoichiometry, gas mixtures.pdf, Problem 7a. **Analysis:** Same procedure as Example 1: convert mass to moles for each gas, then use the Ideal Gas Law to find each partial pressure, and finally sum them for the total pressure.

Step-by-Step Solution:

1. Convert masses to moles:

- Moles CH₄: $1813 \text{ g} \times \frac{1 \text{ mol}}{16.04 \text{ g}} = 113.0 \text{ mol}$
- Moles O₂: $336 \text{ g} \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 10.50 \text{ mol}$

2. Convert temperature to Kelvin: $T = 22.0 + 273.15 = 295.15 \text{ K}$.

3. Calculate partial pressures:

- $P_{CH_4} = \frac{(113.0 \text{ mol})(0.08206)(295.15 \text{ K})}{20.0 \text{ L}} = 136.9 \text{ atm}$
- $P_{O_2} = \frac{(10.50 \text{ mol})(0.08206)(295.15 \text{ K})}{20.0 \text{ L}} = 12.7 \text{ atm}$

4. Calculate total pressure:

$$P_{total} = P_{CH_4} + P_{O_2} = 136.9 \text{ atm} + 12.7 \text{ atm} = 149.6 \text{ atm}$$

Answer: $P_{CH_4} = 136.9 \text{ atm}$, $P_{O_2} = 12.7 \text{ atm}$, $P_{total} = 149.6 \text{ atm}$.

13.2.3 Example 3

Problem: A mixture of gases contains 0.31 mol CH₄, 0.25 mol C₂H₆, and 0.29 mol C₃H₈. The total pressure is 1.50 atm. Calculate the partial pressures of the gases. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.76. **Analysis:** We are given moles and total pressure. The most direct method is to calculate the mole fraction of each gas and then use $P_A = \chi_A P_{total}$. **Step-by-Step Solution:**

1. **Calculate total moles:** $n_{total} = 0.31 + 0.25 + 0.29 = 0.85$ mol.

2. **Calculate mole fractions:**

$$\bullet \chi_{CH_4} = \frac{0.31}{0.85} = 0.365$$

$$\bullet \chi_{C_2H_6} = \frac{0.25}{0.85} = 0.294$$

$$\bullet \chi_{C_3H_8} = \frac{0.29}{0.85} = 0.341$$

3. **Calculate partial pressures:**

$$\bullet P_{CH_4} = 0.365 \times 1.50 \text{ atm} = 0.547 \text{ atm}$$

$$\bullet P_{C_2H_6} = 0.294 \times 1.50 \text{ atm} = 0.441 \text{ atm}$$

$$\bullet P_{C_3H_8} = 0.341 \times 1.50 \text{ atm} = 0.512 \text{ atm}$$

Answer: $P_{CH_4} \approx 0.55$ atm, $P_{C_2H_6} \approx 0.44$ atm, $P_{C_3H_8} \approx 0.51$ atm.

13.2.4 Example 4

Problem: A 2.5-L flask at 15°C contains a mixture of N₂, He, and Ne at partial pressures of 0.32 atm for N₂, 0.15 atm for He, and 0.42 atm for Ne. Calculate the total pressure. **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.77a. **Analysis:** This is a direct application of Dalton's Law of Partial Pressures. **Step-by-Step Solution:**

1. **State Dalton's Law:** $P_{total} = P_{N_2} + P_{He} + P_{Ne}$.

2. **Substitute values and sum:**

$$P_{total} = 0.32 \text{ atm} + 0.15 \text{ atm} + 0.42 \text{ atm} = 0.89 \text{ atm}$$

Answer: 0.89 atm.

14 Topic 10: Gas Collection Over Water

14.1 Key Principles

- When a gas is collected by bubbling it through water, the collected gas is not pure. It is a mixture of the desired gas and water vapor.
- The total pressure measured inside the collection tube (P_{total}) is the sum of the partial pressure of the collected gas (P_{gas}) and the partial pressure of the water vapor (P_{H_2O}).
- $P_{total} = P_{gas} + P_{H_2O}$
- The vapor pressure of water (P_{H_2O}) depends only on the temperature. You will be given this value or a table to find it (like in *lecture_slides.pdf, slide 42*).
- The Key Step:** Before using the collected gas pressure in an ideal gas law calculation, you MUST first subtract the vapor pressure of water to find the partial pressure of the dry gas.

$$P_{gas} = P_{total} - P_{H_2O}$$

14.2 Worked Examples

14.2.1 Example 1

Problem: If 0.0875 L of nitrogen gas was collected over water at 296.15 K (23°C) and 727 mmHg, how many grams of NH₄NO₂ has decomposed? Reaction: NH₄NO₂(s) → N₂(g) + H₂O(l). The vapor pressure of water at 23°C is 21.1 mmHg (Note: slide has 22.1, likely a typo, using 21.1 from standard tables). **Source:** lecture_slides.pdf, slides 43-44. **Analysis:** This is a gas stoichiometry problem that involves collecting a gas over water. First, we must find the partial pressure of the dry N₂. Then use PV = nRT to find the moles of N₂. Finally, use stoichiometry to find the mass of NH₄NO₂. **Step-by-Step Solution:**

1. Find the partial pressure of N₂:

$$P_{N_2} = P_{total} - P_{H_2O} = 727 \text{ mmHg} - 21.1 \text{ mmHg} = 705.9 \text{ mmHg}$$

2. Convert P_{N₂} to atm:

$$705.9 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.9288 \text{ atm}$$

3. Use the Ideal Gas Law to find moles of N₂: $n = \frac{PV}{RT}$.

$$n_{N_2} = \frac{(0.9288 \text{ atm})(0.0875 \text{ L})}{(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(296.15 \text{ K})} = 0.00334 \text{ mol N}_2$$

4. Use stoichiometry to find grams of NH₄NO₂: The mole ratio is 1:1. The molar mass of NH₄NO₂ is 64.06 g/mol.

$$0.00334 \text{ mol N}_2 \times \frac{1 \text{ mol NH}_4\text{NO}_2}{1 \text{ mol N}_2} \times \frac{64.06 \text{ g NH}_4\text{NO}_2}{1 \text{ mol NH}_4\text{NO}_2} = 0.214 \text{ g}$$

Answer: 0.214 g NH₄NO₂.

14.2.2 Example 2

Problem: The hydrogen gas from the reaction Zn(s) + 2HCl(aq) → ZnCl₂(aq) + H₂(g) is collected over water at 25.0°C. The volume of the gas is 7.80 L, and the total pressure is 0.980 atm. Calculate the amount of zinc metal in grams consumed. (Vapor pressure of water at 25°C = 23.8 mmHg). **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.97. **Analysis:** Similar to the previous problem. Find partial pressure of H₂, use ideal gas law to find moles H₂, then use stoichiometry to find grams of Zn. **Step-by-Step Solution:**

1. Convert pressures to be in the same units (atm):

$$P_{H_2O} = 23.8 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.0313 \text{ atm}$$

2. Find the partial pressure of H₂:

$$P_{H_2} = P_{total} - P_{H_2O} = 0.980 \text{ atm} - 0.0313 \text{ atm} = 0.9487 \text{ atm}$$

3. Convert temperature to Kelvin: $T = 25.0 + 273.15 = 298.15 \text{ K}$.

4. Use the Ideal Gas Law to find moles of H₂:

$$n_{H_2} = \frac{(0.9487 \text{ atm})(7.80 \text{ L})}{(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298.15 \text{ K})} = 0.303 \text{ mol H}_2$$

5. Use stoichiometry to find grams of Zn: The mole ratio is 1:1. Molar mass of Zn is 65.39 g/mol.

$$0.303 \text{ mol H}_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 19.8 \text{ g}$$

Answer: 19.8 g Zn.

14.2.3 Example 3

Problem: A piece of sodium metal reacts completely with water: $_2\text{Na(s)} + _2\text{H}_2\text{O(l)} \longrightarrow _2\text{NaOH(aq)} + \text{H}_2\text{(g)}$. The hydrogen gas is collected over water at 25.0°C. The volume is 246 mL at 1.00 atm. Calculate grams of sodium used. (Vapor pressure of water at 25°C = 0.0313 atm). **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.96. **Analysis:** Another problem following the same pattern. **Step-by-Step Solution:**

1. Find the partial pressure of H₂:

$$P_{\text{H}_2} = P_{\text{total}} - P_{\text{H}_2\text{O}} = 1.00 \text{ atm} - 0.0313 \text{ atm} = 0.9687 \text{ atm}$$

2. Convert volume to L and temperature to K: V = 0.246 L, T = 298.15 K.

3. Use the Ideal Gas Law to find moles of H₂:

$$n_{\text{H}_2} = \frac{(0.9687 \text{ atm})(0.246 \text{ L})}{(0.08206)(298.15 \text{ K})} = 0.00974 \text{ mol H}_2$$

4. Use stoichiometry to find grams of Na: The mole ratio is 2 mol Na to 1 mol H₂. Molar mass of Na is 22.99 g/mol.

$$0.00974 \text{ mol H}_2 \times \frac{2 \text{ mol Na}}{1 \text{ mol H}_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 0.448 \text{ g}$$

Answer: 0.45 g Na.

14.2.4 Example 4

Problem: A mixture of helium and neon gases is collected over water at 28.0°C and 745 mmHg. If the partial pressure of helium is 368 mmHg, what is the partial pressure of neon? (Vapor pressure of water at 28°C = 28.3 mmHg). **Source:** Questions and Problems (textbook).pdf, Chapter 11, Problem 11.79.

Analysis: This problem uses Dalton's Law for a collected gas mixture. The total pressure is the sum of the partial pressures of He, Ne, and water vapor. **Step-by-Step Solution:**

1. State Dalton's Law for this mixture: $P_{\text{total}} = P_{\text{He}} + P_{\text{Ne}} + P_{\text{H}_2\text{O}}$.

2. Rearrange to solve for P_{Ne}: $P_{\text{Ne}} = P_{\text{total}} - P_{\text{He}} - P_{\text{H}_2\text{O}}$.

3. Substitute the given values:

$$P_{\text{Ne}} = 745 \text{ mmHg} - 368 \text{ mmHg} - 28.3 \text{ mmHg} = 348.7 \text{ mmHg}$$

Answer: 349 mmHg.

15 Topic 11: KMT: Speed vs. Molar Mass

15.1 Key Principles

- The Kinetic Molecular Theory (KMT) states that at a given temperature, all gases in a mixture have the **same average kinetic energy**.
- Kinetic energy is defined as $KE = \frac{1}{2}mv^2$. For this to be constant across different gases, a gas with a larger mass (m) must have a smaller average velocity (v), and vice versa.
- This means there is an **inverse relationship** between molar mass and molecular speed: lighter gases move faster, heavier gases move slower.
- On a velocity distribution graph for different gases at the same temperature:
 - The peak for the **heavier gas** will be further to the left (lower speed) and taller.
 - The peak for the **lighter gas** will be further to the right (higher speed) and shorter/flatter.

15.2 Worked Examples

15.2.1 Example 1

Problem: Consider the graph comparing the velocity distribution of PH_3 and Kr at the same temperature. Identify which curve represents Kr. **Source:** Kinetic molecular theory.pdf, Problem 2a. **Analysis:** We need to compare the molar masses of PH_3 and Kr. The heavier gas will be the slower one (peak to the left). **Step-by-Step Solution:**

1. Calculate molar masses:
 - $M(\text{PH}_3) = 30.97 + 3(1.008) \approx 34.0 \text{ g/mol.}$
 - $M(\text{Kr}) = 83.80 \text{ g/mol.}$
2. Compare: Kr is significantly heavier than PH_3 .
3. Relate to the graph: The heavier gas (Kr) will have a lower average speed. Curve A has its peak at a lower velocity than Curve B. Therefore, Curve A represents Kr.

Answer: Curve A.

15.2.2 Example 2

Problem: Using the same graph as Example 1, which curve represents the gas with a higher velocity? **Source:** Kinetic molecular theory.pdf, Problem 2d. **Analysis:** The gas with the higher velocity will be the lighter gas. Its curve will be shifted to the right. **Step-by-Step Solution:**

1. From the previous example, we know PH_3 is lighter than Kr.
2. Lighter gases move faster.
3. Curve B has its peak at a higher velocity (further right) than Curve A.
4. Therefore, Curve B represents the gas with the higher velocity.

Answer: Curve B.

15.2.3 Example 3

Problem: The velocity distribution for four different molecules (A, B, C, D) at the same temperature is shown. Which gas has the largest molar mass? **Source:** Kinetic molecular theory.pdf, Problem 3a. **Analysis:** The gas with the largest molar mass will be the heaviest, and therefore the slowest. We need to find the curve with the lowest average speed (peak furthest to the left). **Step-by-Step Solution:**

1. The relationship is: largest molar mass \iff heaviest gas \iff slowest average speed.
2. On the graph, the horizontal axis is molecular speed. Lower speed is to the left.
3. Curve A has its peak further to the left than any other curve, indicating it has the lowest average speed.
4. Therefore, Gas A has the largest molar mass.

Answer: Gas A.

15.2.4 Example 4

Problem: True or false: Gas "A" has a smaller molar mass than gas "D". **Source:** Kinetic molecular theory.pdf, Problem 3c. **Analysis:** We compare the positions of the peaks for A and D. The peak further to the right corresponds to a higher speed and thus a smaller molar mass. **Step-by-Step Solution:**

1. Curve D's peak is much further to the right than Curve A's peak.
2. This means Gas D has a much higher average speed than Gas A.
3. Since speed is inversely proportional to the square root of molar mass, the faster gas (D) must have a smaller molar mass than the slower gas (A).
4. The statement that Gas "A" has a smaller molar mass than gas "D" is therefore false.

Answer: False.

16 Topic 12: Diffusion and Effusion

16.1 Key Principles

- **Diffusion:** The mixing of gases as a result of their random motion and frequent collisions.
- **Effusion:** The escape of gas molecules through a tiny hole into a vacuum.
- **Graham's Law of Effusion:** The rate of effusion of a gas is inversely proportional to the square root of its molar mass (M).

$$\text{rate} \propto \frac{1}{\sqrt{M}}$$

- This means that lighter gases effuse (and diffuse) more rapidly than heavier gases.
- Because rate is distance/time, the time it takes for a gas to effuse is directly proportional to the square root of its molar mass ($\text{time} \propto \sqrt{M}$). A heavier gas will take longer to effuse.

16.2 Worked Examples

16.2.1 Example 1

Problem: Which gas effuses most rapidly: A, B, C, or D from the graph in Topic 11, Example 3? **Source:** Kinetic molecular theory.pdf, Problem 3b. **Analysis:** The gas that effuses most rapidly is the one with the highest average speed, which corresponds to the one with the lowest molar mass. We look for the peak that is furthest to the right. **Step-by-Step Solution:**

1. Rate of effusion is fastest for the gas with the highest speed.
2. The graph shows that Gas D has the highest average speed (peak is furthest right).
3. Therefore, Gas D effuses most rapidly.

Answer: Gas D.

16.2.2 Example 2

Problem: Which curve represents the gas that diffuses more slowly, PH₃ or Kr? (Using the graph from Topic 11, Example 1). **Source:** Kinetic molecular theory.pdf, Problem 2c. **Analysis:** The gas that diffuses more slowly is the heavier gas (larger molar mass). **Step-by-Step Solution:**

1. We know from before that Kr (M = 83.8 g/mol) is heavier than PH₃ (M = 34.0 g/mol).
2. The heavier gas diffuses more slowly.
3. Curve A represents Kr. Therefore, Curve A represents the gas that diffuses more slowly.

Answer: Curve A.

16.2.3 Example 3

Problem: Arrange the following species in order of increasing rate of effusion: Cl₂, Ar, HI. **Source:** Kinetic molecular theory.pdf, Problem 6. **Analysis:** The rate of effusion is inversely proportional to the square root of the molar mass. To arrange them in order of increasing rate, we must first arrange them in order of decreasing molar mass (heaviest to lightest). **Step-by-Step Solution:**

1. Calculate the molar masses:

- M(HI) = 126.9 + 1.008 = 127.9 g/mol
- M(Cl₂) = 2 × 35.45 = 70.90 g/mol
- M(Ar) = 39.95 g/mol

2. Order by molar mass (heaviest to lightest): HI < Cl₂ < Ar.

3. The rate of effusion will be the reverse order (slowest to fastest): HI ↑ Cl₂ ↑ Ar.

Answer: The order of increasing rate of effusion is HI ↑ Cl₂ ↑ Ar.

16.2.4 Example 4

Problem: It takes 54 seconds for N₂O gas to effuse from a container. How much time would it take for the same amount of I₂ gas to effuse under the same conditions? **Source:** Kinetic molecular theory.pdf, Problem 7. **Analysis:** The time of effusion is directly proportional to the square root of the molar mass. We can set up a ratio: $\frac{\text{time}_A}{\text{time}_B} = \sqrt{\frac{M_A}{M_B}}$. **Step-by-Step Solution:**

1. Calculate molar masses:

- M(N₂O) = 2(14.01) + 16.00 = 44.02 g/mol
- M(I₂) = 2 × 126.9 = 253.8 g/mol

2. Set up the ratio equation. Let A = I₂ and B = N₂O.

$$\frac{\text{time}_{I_2}}{\text{time}_{N_2O}} = \sqrt{\frac{M_{I_2}}{M_{N_2O}}}$$

3. Rearrange and solve for time_{I₂}:

$$\text{time}_{I_2} = \text{time}_{N_2O} \times \sqrt{\frac{M_{I_2}}{M_{N_2O}}} = 54 \text{ s} \times \sqrt{\frac{253.8 \text{ g/mol}}{44.02 \text{ g/mol}}}$$

$$\text{time}_{I_2} = 54 \text{ s} \times \sqrt{5.765} = 54 \text{ s} \times 2.40 = 129.6 \text{ s}$$

Answer: 130 seconds.

17 Topic 13: Ratio of Molecular Speeds (Graham's Law)

17.1 Key Principles

- Graham's Law provides a quantitative relationship to compare the rates (or speeds) of effusion/diffusion of two different gases, A and B.

- **Memorization Required: The Formula**

$$\frac{\text{rate}_A}{\text{rate}_B} = \frac{u_{rms,A}}{u_{rms,B}} = \sqrt{\frac{M_B}{M_A}}$$

- Note the inverse relationship: Gas A is in the numerator on the left, but its molar mass M_A is in the denominator on the right.
- This equation can be used to find the rate ratio if molar masses are known, or to find an unknown molar mass if the rate ratio is known.

17.2 Worked Examples

17.2.1 Example 1

Problem: An unknown gas diffuses 1.12 times faster than argon (Ar) through a porous membrane. Calculate the molar mass of the unknown gas. **Source:** Kinetic molecular theory.pdf, Problem 4. **Analysis:** This is a direct application of Graham's Law to find an unknown molar mass. Let the unknown gas be X. The ratio of rates is given as $\frac{\text{rate}_X}{\text{rate}_{\text{Ar}}} = 1.12$. **Step-by-Step Solution:**

1. Look up the molar mass of Argon: $M(\text{Ar}) = 39.95 \text{ g/mol}$.

2. Set up Graham's Law:

$$\frac{\text{rate}_X}{\text{rate}_{\text{Ar}}} = \sqrt{\frac{M_{\text{Ar}}}{M_X}} \implies 1.12 = \sqrt{\frac{39.95 \text{ g/mol}}{M_X}}$$

3. Square both sides to remove the square root:

$$(1.12)^2 = \frac{39.95}{M_X} \implies 1.2544 = \frac{39.95}{M_X}$$

4. Rearrange and solve for M_X :

$$M_X = \frac{39.95 \text{ g/mol}}{1.2544} = 31.84 \text{ g/mol}$$

Answer: 31.8 g/mol.

17.2.2 Example 2

Problem: Calculate the diffusion rate of hydrobromic acid (HBr) relative to Xenon (Xe). **Source:** Kinetic molecular theory.pdf, Problem 5. **Analysis:** The question asks for the ratio of the rates, $\frac{\text{rate}_{\text{HBr}}}{\text{rate}_{\text{Xe}}}$. We can calculate this using their molar masses. **Step-by-Step Solution:**

1. Find the molar masses:

- $M(\text{HBr}) = 1.008 + 79.90 = 80.91 \text{ g/mol}$.
- $M(\text{Xe}) = 131.29 \text{ g/mol}$.

2. Set up the Graham's Law ratio:

$$\frac{\text{rate}_{\text{HBr}}}{\text{rate}_{\text{Xe}}} = \sqrt{\frac{M_{\text{Xe}}}{M_{\text{HBr}}}} = \sqrt{\frac{131.29 \text{ g/mol}}{80.91 \text{ g/mol}}}$$

3. Calculate the result:

$$\frac{\text{rate}_{\text{HBr}}}{\text{rate}_{\text{Xe}}} = \sqrt{1.6227} = 1.27$$

Answer: HBr diffuses 1.27 times faster than Xe.

17.2.3 Example 3

Problem: Determine the molar mass and identity of a diatomic gas that moves 4.67 times as fast as CO₂. **Source:** lecture_slides.pdf, slide 54. **Analysis:** Similar to Example 1. Let the unknown diatomic gas be X. The rate ratio is $\frac{\text{rate}_X}{\text{rate}_{\text{CO}_2}} = 4.67$. We need to solve for M_X . **Step-by-Step Solution:**

1. Molar mass of CO₂ is 44.01 g/mol.

2. Set up Graham's Law:

$$4.67 = \sqrt{\frac{M_{CO_2}}{M_X}} = \sqrt{\frac{44.01 \text{ g/mol}}{M_X}}$$

3. Square both sides:

$$(4.67)^2 = \frac{44.01}{M_X} \implies 21.81 = \frac{44.01}{M_X}$$

4. Solve for M_X :

$$M_X = \frac{44.01 \text{ g/mol}}{21.81} = 2.018 \text{ g/mol}$$

5. Identify the diatomic gas: A diatomic gas with a molar mass of 2 g/mol is hydrogen, H_2 .

Answer: Molar Mass = **2.02 g/mol**; Identity = **H_2** .

17.2.4 Example 4

Problem: How much faster does $^{35}\text{UF}_6$ effuse than $^{38}\text{UF}_6$? Molar mass of F is 19.00 g/mol. **Source:**

Questions and Problems (textbook).pdf, Chapter 11, Problem 11.12. **Analysis:** We need to find the ratio of the rates, $\frac{\text{rate}_{235}}{\text{rate}_{238}}$. This requires calculating the molar mass of each isotopic compound. **Step-by-Step Solution:**

1. Calculate the molar masses:

- $M(^{35}\text{UF}_6) = 235.04 + 6(19.00) = 349.04 \text{ g/mol}$
- $M(^{38}\text{UF}_6) = 238.05 + 6(19.00) = 352.05 \text{ g/mol}$

2. Set up the Graham's Law ratio:

$$\frac{\text{rate}_{235}}{\text{rate}_{238}} = \sqrt{\frac{M_{238}}{M_{235}}} = \sqrt{\frac{352.05 \text{ g/mol}}{349.04 \text{ g/mol}}}$$

3. Calculate the result:

$$\frac{\text{rate}_{235}}{\text{rate}_{238}} = \sqrt{1.0086} = 1.0043$$

Answer: $^{35}\text{UF}_6$ effuses **1.0043 times faster** than $^{38}\text{UF}_6$.