

The Gas State

- Gases are everywhere – atmosphere, environmental processes, industrial processes, bodily functions
- Gases have unique properties from liquids and solids
- Gases are compressible (very important for storage)
- Gas particles are widely separated and move at very fast speeds
- Most gases have relatively low densities
- Gas have relatively low viscosity (resistance to movement) allowing them move freely through pipes and even small orifices

The Gas State

- Chemical behavior of gases depends on composition
- Physical behavior of all gases is similar
- Gases are miscible mixing together in any proportion
- Behavior of gases described by **ideal gas law** and **kinetic-molecular theory**, the cornerstone of this chapter
 - Gas volume changes greatly with pressure
 - Gas volume changes greatly with temperature
 - Gas volume is a function of the amount of gas

Introduction to Pressure

- To understand gases, we must first define *force* and *pressure* (primary metric for gases)
- *Force* (F) is a function of the mass of an object under acceleration, such as the force of gravity on masses

$$\mathbf{F = mass \times acceleration\ of\ gravity = m \times a}$$

- *Pressure* (P) is the Force (F) exerted per unit area of surface by molecules in motion

$$\mathbf{P = Force/unit\ area = F / A}$$

Introduction to Pressure

■ Pressure Units

$$P = \frac{\text{Force}}{\text{Area}} = \frac{F}{A} = \frac{\text{mass}(m) \times \text{acceleration} \left(\frac{d}{t^2} \right)}{\text{area}(d^2)} = \frac{\text{kg} \times \frac{m}{s^2}}{m^2} = \frac{\text{kg}}{m \cdot s^2}$$

$$1 \text{ Pascal (Pa)} = 1 \text{ kg/m} \cdot \text{s}^2$$

$$1 \text{ Atmosphere (atm)} = 101,325 \text{ Pascals} = 101.325 \text{ kPa}$$

$$1 \text{ Bar} = 100 \text{ kPa}$$

$$1 \text{ Atmosphere (atm)} = 1.013 \text{ bar}$$

$$1 \text{ Atmosphere (atm)} = 14.7 \text{ psi}$$

$$1 \text{ Atmosphere (atm)} = 760 \text{ mm Hg} = 760 \text{ torr}$$

$$1 \text{ torr} = 1 \text{ mm Hg} = 133.3 \text{ Pa}$$

Ways to Measure Pressure

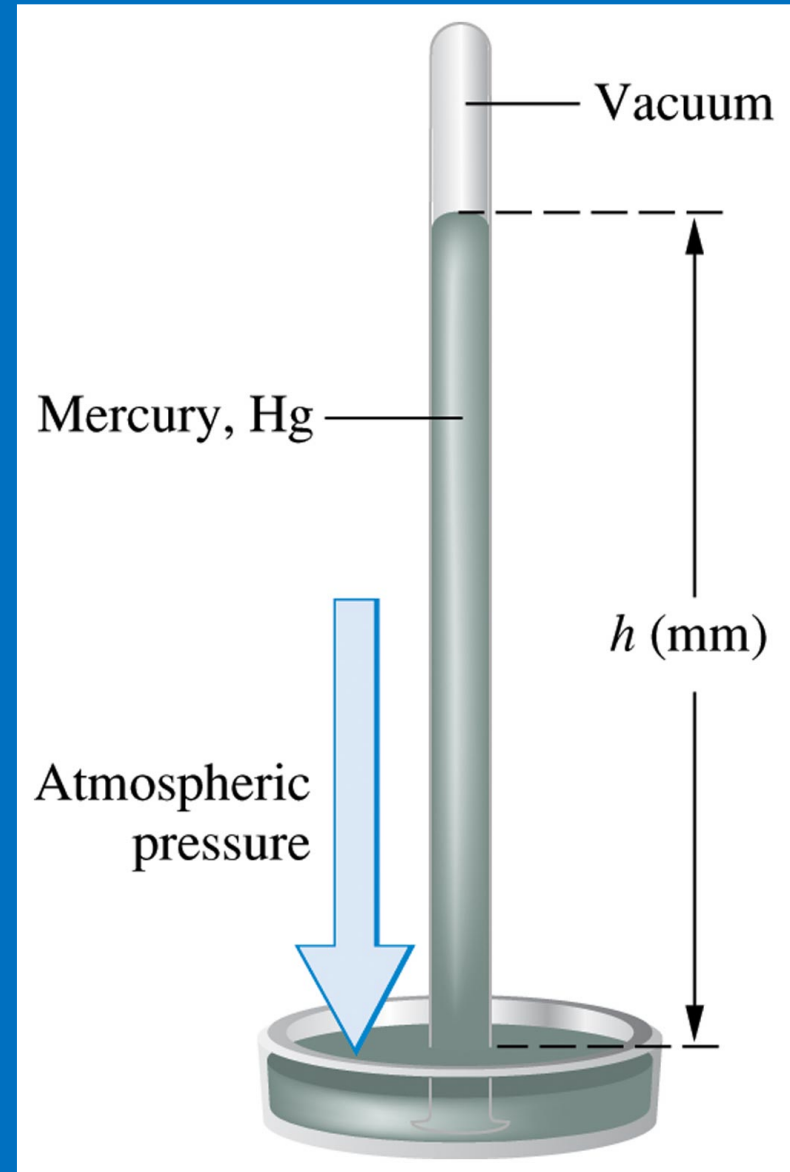
- **Barometer** – device for measuring pressure of the atmosphere
- **Manometer** – device for measuring the pressure of a gas or liquid in a vessel
- **Barometric Pressure** involves the acceleration due to gravity (**g**), the density (**d**) of a liquid in the manometer column and the height (**h**) of the liquid in the column

$$P = g d h$$

Ways to Measure Pressure

Mercury Barometer

At sea level,
 $h = 760 \text{ mm}$
 $= 1 \text{ atm}$
 $= 101.325 \text{ kPa}$



The Empirical Gas Laws

- Gas behavior can be described by pressure (P), temperature (T), volume (V), and molar amount (n)
- Holding any two of the variables constant allows relations between the other two
- **Boyle's Law**

The volume of a sample of gas at a given temperature varies **inversely** with the applied pressure

$$V \propto 1/P \quad (\mathbf{T \ \& \ n \ are \ constant})$$

$$\mathbf{PV = Const}$$

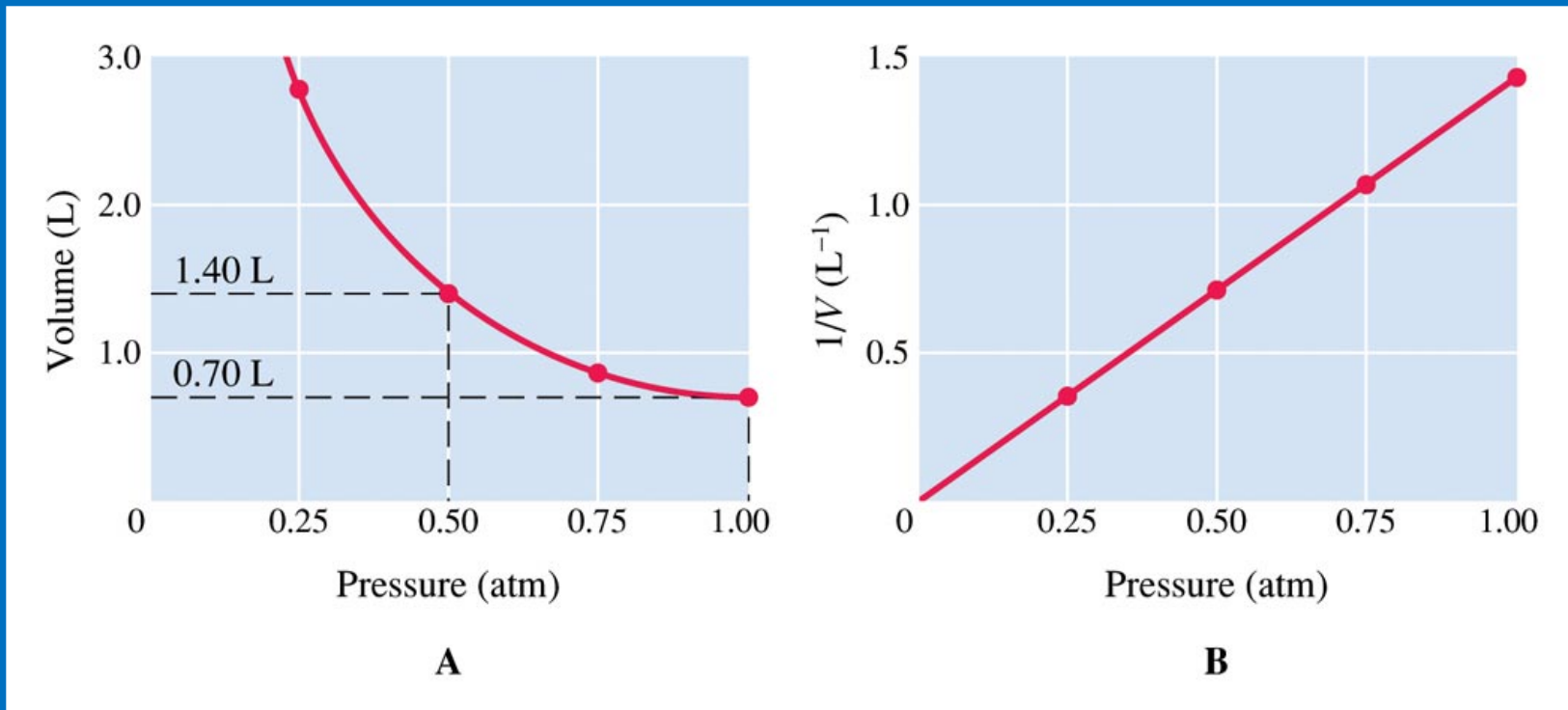
$$\mathbf{P_1V_1 = P_2V_2}$$

$$\mathbf{1/V = a(P) + b \quad (\mathbf{Linear \ Plot})}$$

The Empirical Gas Laws

Boyle's Law

Two Graphs Illustrating Gas Pressure-Volume Relationship



Practice Problem

Boyle's Law

A sample of Chlorine gas has a volume of 1.8 L at 1.0 atm. If the pressure increases to 4.0 atm (at constant temperature), what would be the new volume?

using $P_1 V_1 = P_2 V_2$

$$V_2 = \frac{P_1 V_1}{P_2} = \frac{(1.0 \text{ atm}) \times (1.8 \text{ L})}{(4.0 \text{ atm})}$$

$$V_f = 0.45 \text{ L}$$

The Empirical Gas Laws

Charles's Law

- The volume occupied by any sample of gas at constant pressure is directly proportional to its absolute temperature

$$V \propto T_{\text{abs}} \quad T_{\text{abs}} (\text{K}) = ^\circ\text{C} + 273.15$$

P & n are constant

- Assumes constant moles and pressure
- Temperature on absolute scale ($^\circ\text{C} + 273.15$)

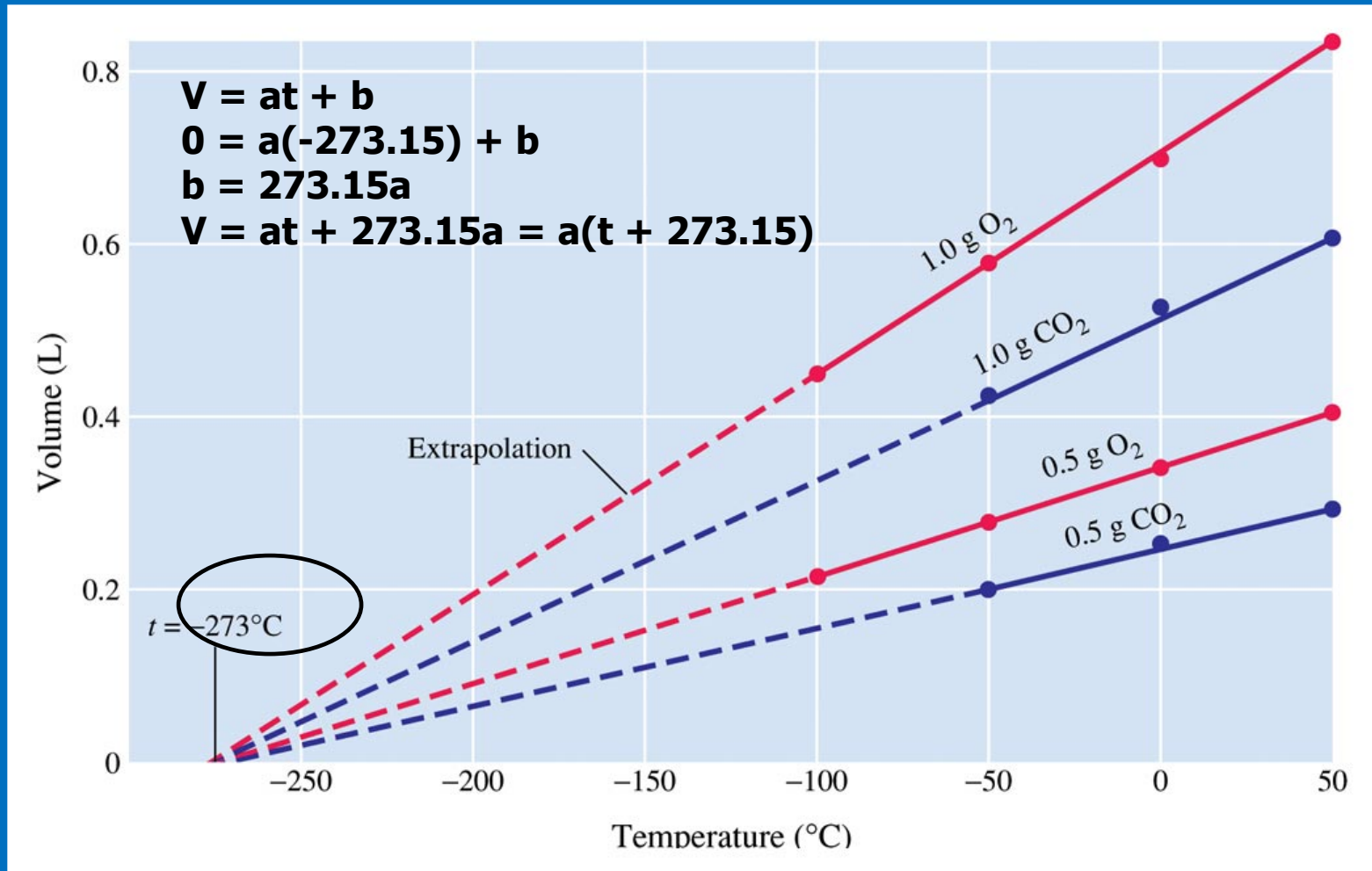
$$\frac{V}{T} = \text{constant} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V = a(T) + b \quad (\text{Linear Plot})$$

The Empirical Gas Laws

Charles's Law

Linear Relationship of Gas Volume and Temperature at Constant Pressure



Practice Problem

Charles's Law

A sample of Methane gas that has a volume of 3.8 L at 5.0°C is heated to 86.0°C at constant pressure. Calculate its new volume.

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

Convert temperature (°C) to Absolute (Kelvin) K

$$5.0^{\circ}\text{C} = 5.0 + 273.15 = 278.15 \text{ K}$$

$$86.0^{\circ}\text{C} = 86.0 + 273.15 = 359.15 \text{ K}$$

$$V_2 = \left(\frac{T_2}{T_1}\right) V_1 = \left(\frac{359.15 \text{ K}}{278.15 \text{ K}}\right) \times 3.8 \text{ L}$$

$$V_2 = 4.9 \text{ L}$$

The Empirical Gas Laws

- Amontons's Law - Pressure / Temperature relationship
 - Charles's Law is expressed as the effect of temperature change on gas volume

$$\frac{V}{T} = \text{constant}$$

- Volume and Pressure are interdependent; therefore the effect of temperature on volume is closely related to its effect on pressure

$$P \propto T_{\text{abs}}$$

P & n fixed

$$\frac{P}{T} = \text{constant}$$

The Empirical Gas Laws

- **Combined Gas Law:** In the event that all three parameters, P , V , and T , are changing, their combined relationship is defined as follows assuming the mass of the gas (number of moles) is constant

Recall:

$$V \propto 1/P$$

$$PV = \text{constant}$$

$$V \propto T$$

$$V/T = \text{constant}$$

$$P \propto T$$

$$P/T = \text{constant}$$

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

$$(PV)/T = \text{constant}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Practice Problem

The pressure of 4.0 L of Nitrogen in a flexible container is decreased to one-half its original pressure, and its absolute temperature is increased to double the original temperature. The new volume is now

- a. 2.0 L b. 4.0 L c. 8.0 L d. 16.0 L e. 32.0 L

Ans: d

$$\frac{P_1 V_1 \text{ (new)}}{T_1} = \frac{P_2 V_2 \text{ (old)}}{T_2} \quad V_1 = \frac{T_1 \cdot P_2 \cdot V_2}{P_1 \cdot T_2}$$

New Pressure (P_1) is equal to 1 / 2 original pressure $\therefore P_1 = \frac{P_2}{2}$

New Temperature (T_1) is equal to twice original temperature $\therefore T_1 = 2T_2$

$$V_1 = \frac{T_1 P_2 V_2}{P_1 T_2} = \frac{2T_2 P_2 V_2}{\frac{P_2}{2} T_2} = \frac{2T_2 P_2 V_2}{\frac{P_2 T_2}{2}} = \frac{4P_2 V_2 T_2}{\frac{P_2 T_2}{2}} = 4 \times V_2 = 4 \times 4.0 = 16 \text{ L}$$

Avagadro's Law

Avogadro's Law

- The volume of a sample of gas is directly proportional to the number of moles of gas, n

$$V \propto n$$

$$\frac{V}{n} = \text{constant}$$

- Equal volumes of any two gases at the same temperature and pressure contain the same number of molecules (mol)

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

Avagadro's Law

■ Avogadro's Law

- The volume of one mole of gas is called the:

molar gas volume, V_m

- Volumes of gases are often compared at standard temperature and pressure (STP), chosen to be

0 °C (273.15 °K) and 1 atm pressure

- At STP, the molar volume, V_m , that is, the volume occupied by one mole of any gas, is **22.4 L/mol**

$$V_{\text{stp}} / n_{\text{stp}} = V_m = 22.4 \text{ L (at STP)}$$

Practice Problem

A sample of Fluorine gas has a volume of 5.80 L at 150.0 °C and 10.5 atm of pressure. How many moles of Fluorine gas are present?

First, use the combined Empirical Gas Law to determine the volume at STP. Then, use Avagadro's law to determine the number of moles

$$\frac{V_{\text{stp}} P_{\text{stp}}}{T_{\text{stp}}} (\text{new}) = \frac{P_1 V_1}{T_1} (\text{old})$$

$$V_{\text{STP}} = \frac{P_1 V_1 T_{\text{stp}}}{P_{\text{stp}} T_1} = \frac{(10.5 \text{ atm}) (5.80 \text{ L}) (273 \text{ K})}{(1.0 \text{ atm}) (423 \text{ K})} = 39.3 \text{ L}$$

$$\frac{V_{\text{stp}}}{n_{\text{stp}}} = \frac{V_2}{n_2} \quad \frac{22.4 \text{ L}}{1 \text{ mol}} = \frac{39.3 \text{ L}}{n} \quad n = 1.75 \text{ mol}$$

The Ideal Gas Law

- From the empirical gas laws, we see that volume varies in proportion to pressure, absolute temperature, and the mass of gas (moles) present

Boyle's Law $V \propto 1/P$ $V = \text{Const} \times \frac{1}{P}$

Charles' Law $V \propto T_{\text{abs}}$ $V = \text{Const} \times T$

Avogadro's Law $V \propto n$ $V = \text{Const} \times n$

Combined Gas Law $(PV) / T = \text{constant}$

The Ideal Gas Law

- This implies that there must exist a proportionality constant governing these relationships
- From the **Combined Gas Law**, we can obtain the following relationship

$$V_m = "R" \times n \times \left(\frac{T_{\text{abs}}}{P} \right)$$

where "R" is the proportionality constant referred to as the **Ideal Gas Constant**, which relates Molar Volume (V) to the ratio of Temperature to Pressure T/P

The Ideal Gas Law

- The numerical value of “R” can be derived using Avogadro’s law, which states that *one mole* of any gas at STP will occupy 22.4 liters

$$R = \frac{VP}{nT}$$

$$R = \frac{(22.4 \text{ L}) (1.00 \text{ atm})}{(1.00 \text{ mol}) (273 \text{ K})}$$

$$= 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

The Ideal Gas Law

- The **ideal gas equation** is usually expressed in the following form:

$$PV = nRT$$

P – Pressure (in atm)

V – Volume (in liters)

n – Number of atoms (in moles)

R – Universal gas constant 0.0821 L•atm/mol•K

T – Temperature (in °Kelvin = °C + 273.15)

Practice Problem

A steel tank has a volume of 438 L and is filled with 0.885 kg of O₂.

Calculate the pressure of Oxygen in the tank at 21°C

$$\text{Use } PV = nRT \quad \therefore P = \frac{nRT}{V}$$

$$V = 438 \text{ L} \quad T = 21^\circ\text{C} = 21 + 273.15 = 294.15^\circ\text{K}$$

$$n = 0.885 \text{ kg O}_2 \text{ (convert to moles)} \quad R = 0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}}$$

$$n = \text{mol O}_2 = 0.885 \text{ kg O}_2 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 27.7 \text{ mol O}_2$$

$$P = \frac{nRT}{V} = \frac{27.7 \text{ mol O}_2 \times 0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}} \times 294.15^\circ\text{K}}{438 \text{ L}} = 1.53 \text{ atm}$$

Practice Problem

A 50-L cylinder of Nitrogen, N_2 , has a pressure of 17.1 atm at 23 °C. What is the mass (g) of Nitrogen in the cylinder?

$$PV = nRT \quad n = \frac{PV}{RT} \quad n = \frac{m}{M_m} \quad R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

$$V = 50 \text{ L} \quad P = 17.1 \text{ atm} \quad T = 23 \text{ }^\circ\text{C} + 273 = 296 \text{ }^\circ\text{K}$$

$$M_m(N_2) = 2 \frac{\text{mol N}}{\text{mol } N_2} \times 14.01 \frac{\text{g N}}{\text{mol N}} = 28.02 \frac{\text{g}}{\text{mol } N_2}$$

$$m = \frac{M_m \times P \times V}{R \times T} = \frac{28.02 \frac{\text{g}}{\text{mol}} \times 17.1 \text{ atm} \times 50 \text{ L}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 296 \text{ }^\circ\text{K}}$$

$$m = 985.8 \text{ g } N_2$$

Ideal Gas Law

- Molecular Weight Determination
- If we substitute $(\text{mass}(m)/\text{Mol Wgt}(M_m) = \text{moles } (n))$ in the ideal gas equation, we obtain a direct mathematical expression for the molar mass (molecular weight) of an ideal gas

$$\text{moles } (n) = \frac{m}{M_m} = \frac{\text{mass (g)}}{\text{molar mass} \left(\frac{\text{g}}{\text{mol}} \right)}$$

$$PV = nRT = \left(\frac{m}{M_m} \right) RT$$

$$M_m = \frac{mRT}{PV}$$

Ideal Gas Law

■ Application of Ideal Gas Law

➤ Density Determination

- Consider the derivation of the molar mass equation for gases

$$PV = \left(\frac{m}{M_m}\right)RT$$

- Solve for m/V (g/L), i.e. Density (d)

$$\frac{m}{V} = d = \frac{PM_m}{RT}$$

- Note that gas density is a function of its molar mass

Practice Problem

What is the density of Methane gas (natural gas), CH_4 , at 125°C and 3.50 atm ?

$$\frac{m}{V} = d = \frac{PM_m}{RT}$$

$$T = (125^\circ\text{C} + 273.15) = 398.15^\circ\text{K}$$

$$P = 3.50\text{ atm}$$

$$M_m(\text{CH}_4) = 12.01 \frac{\text{g C}}{\text{mol C}} + 4 \times 1.01 \frac{\text{g H}}{\text{mol H}} = 16.05 \frac{\text{g CH}_4}{\text{mol CH}_4}$$

$$R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

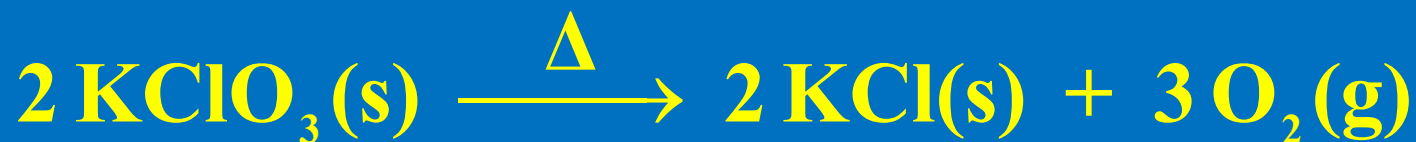
$$d = \frac{PM_m}{RT} = \frac{3.50\text{ atm} \times 16.05 \frac{\text{g CH}_4}{\text{mol CH}_4}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{1000\text{ mL}}{1\text{ L}} \times 398.15^\circ\text{K}} = 0.00172 \frac{\text{g}}{\text{ml}}$$

Practice Problem

Stoichiometry Problem Involving Gas Volumes

Suppose you heat 0.0100 mol of Potassium Chlorate, KClO_3 , in a test tube. How many liters of Oxygen can you produce at 298 K and 1.02 atm?

Consider the following reaction, which is often used to generate small quantities of Oxygen



Con't on next slide

Practice Problem

First we must determine the number of moles of Oxygen produced by the reaction.

$$0.0100 \text{ mol KClO}_3 \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3}$$

$$= 0.0150 \text{ mol O}_2$$

Con't on next slide

Practice Problem

Now we can use the ideal gas equation to calculate the volume of Oxygen under the conditions given.

$$V = \frac{nRT}{P}$$

$$V = \frac{(0.0150 \text{ mol O}_2)(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})}{1.02 \text{ atm}}$$

$$V = 0.360 \text{ L}$$

Practice Problem

When a 2.0 Liter bottle of concentrated HCL was spilled, 1.2 kg of CaCO₃ (FW – 100.09) was required to neutralize the spill. What volume of CO₂ was released by the reaction at 735 mm Hg and 20°C? The balanced equation is:



Compute moles CaCO₃

$$1.2 \text{ kg CaCO}_3 \left(\frac{1000 \text{ g CaCO}_3}{1 \text{ kg CaCO}_3} \right) \left(\frac{1 \text{ mol}}{100.09 \text{ g CaCO}_3} \right) = 11.9892 \text{ mol CaCO}_3$$

Molar Ratio CaCO₃ to CO₂ = 1:1 ∴ n = 11.9892 mol CO₂ released

Convert pressure to atmospheres $735 \text{ mmHg} \left(\frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.9671 \text{ atm}$

$$V = \frac{nRT}{P} \quad R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \quad T = 20^\circ \text{C} + 273.15 = 293.15^\circ \text{K}$$

$$V = \frac{(11.9892 \text{ mol CO}_2) (0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}) (293.15 \text{ K})}{0.9671 \text{ atm}} = 298. \text{L}$$

Partial Pressures of Gas Mixtures

- **Dalton's Law of Partial Pressures:** the sum of all the pressures of all the different gases (a, b, c, etc.) in a mixture equals the total pressure of the mixture.

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

- The individual pressures follow the ideal gas law

$$P_a V = n_a R T$$

Partial Pressures of Gas Mixtures

- The composition of a gas mixture is often described in terms of its mole fraction.
- The *mole fraction, X_a* , of a component gas is the fraction of moles of that component in the total moles of gas mixture.

$$\mathbf{n_{tot} = n_a + n_b + n_c}$$

$$X_a = \frac{n_a}{n_{tot}} = \frac{\frac{P_a V}{RT}}{\frac{P_t V}{RT}} = \frac{P_a}{P_T}$$

Partial Pressures of Gas Mixtures

The partial pressure of a component gas, "A", is then defined as

$$P_A = X_A \times P_{\text{tot}}$$

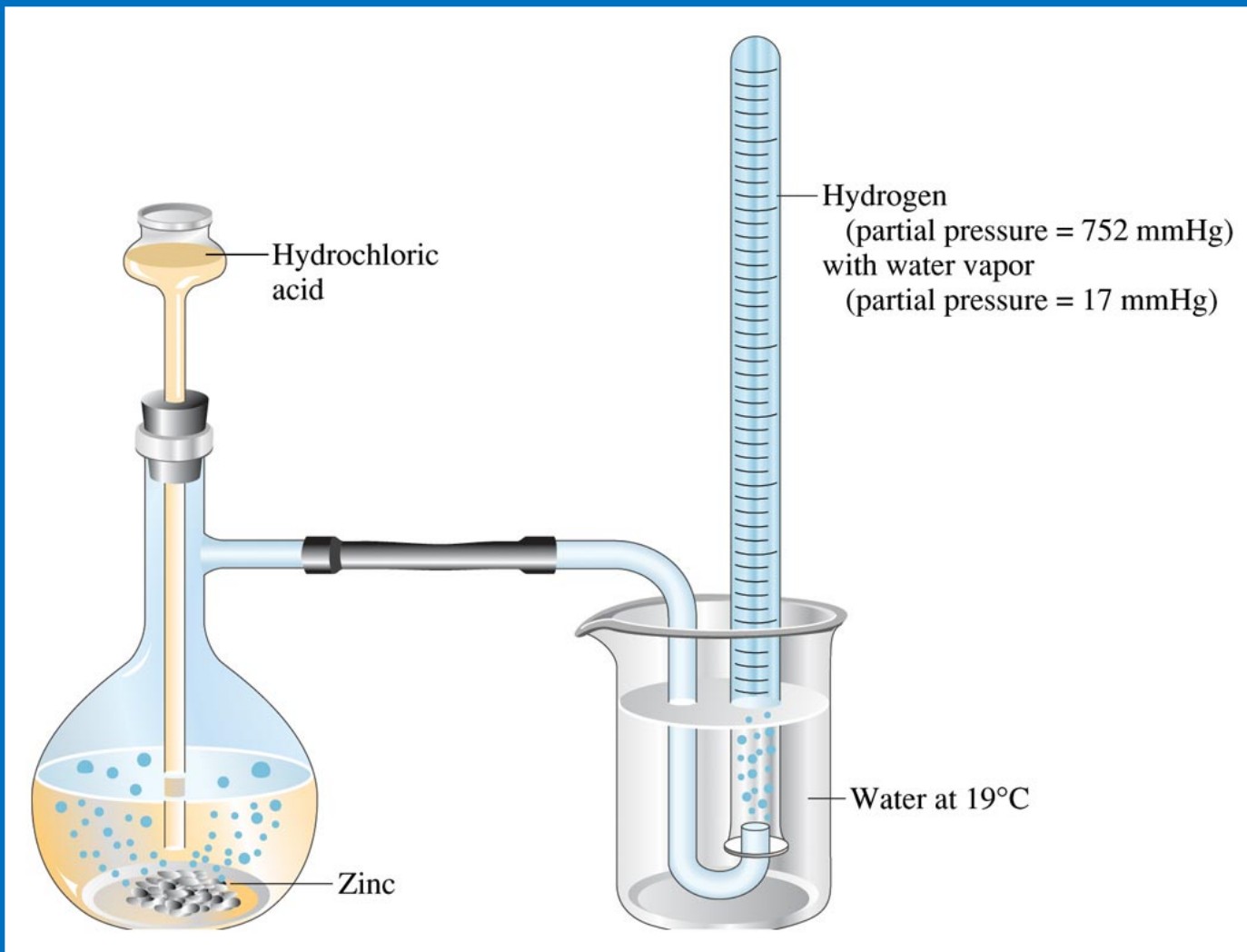
Applying this concept to the *ideal gas equation*, we find that each gas can be treated *independently*

$$P_A V = n_A RT$$

Collecting Gases “Over Water”

- A useful application of partial pressures arises when you collect gases over water
 - As gas bubbles through the water, the gas becomes saturated with water vapor
 - The partial pressure of the water in this “mixture” depends only on the temperature

Collection of Gas Over Water



Practice Problem

Suppose a 156 mL sample of H₂ gas was collected over Water at 19°C and 769 mm Hg.

What is the mass of H₂ collected?

First, find the partial pressure of the dry H₂

$$P_{\text{H}_2} = P_{\text{tot}} - P_{\text{H}_2\text{O}}$$

The vapor pressure of water at 19°C is 16.5 mm Hg

$$P_{\text{H}_2} = 769 \text{ mm Hg} - 16.5 \text{ mm Hg}$$

$$P_{\text{H}_2} = 752 \text{ mm Hg}$$

Con't on next Slide

Practice Problem (Con't)

- Now use the ideal gas equation, along with the partial pressure of the Hydrogen, to determine its mass

$$P_{\text{H}_2} = 752 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.989 \text{ atm}$$

$$V = 156 \text{ mL} = 0.156 \text{ L Hydrogen}$$

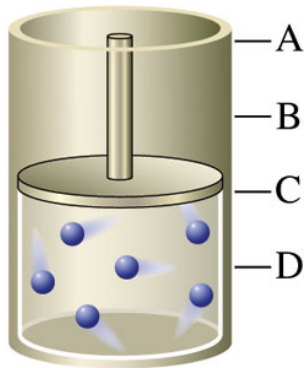
$$T = (19 + 273) = 292 \text{ }^\circ \text{K}$$

$$\begin{aligned} n &= PV / RT = [(0.989 \text{ atm})(0.156)] / [(0.0821)(292\text{K})] \\ &= 0.00644 \text{ mol H}_2 \end{aligned}$$

$$0.00644 \text{ mol H}_2 \times 2.016 \text{ g H}_2 / \text{mol H}_2 = 0.013 \text{ g Hydrogen}$$

Practice Problem

Conceptual Problem 5.25 Consider the following gas container equipped with a movable piston.



- By what factor (increase by 1, decrease by 1.5, etc.) would you change the pressure if you wanted the volume to change from volume C to volume D?
- If the piston were moved from volume C to volume A, by what factor would the pressure change?
- By what factor would you change the temperature in order to change from volume C to volume B?
- If you increased the number of moles of gas in the container by a factor of 2, by what factors would the pressure and the volume change?

$$PV = \text{const}$$

$$V/T = \text{const}$$

$$V/n = \text{const}$$

$$PV = nRT$$

- Pressure would increase by a factor of "2"
- Pressure would decrease by factor of "2"
- Increase temperature by a factor of "1.5"
- The volume would increase by a factor of "2"

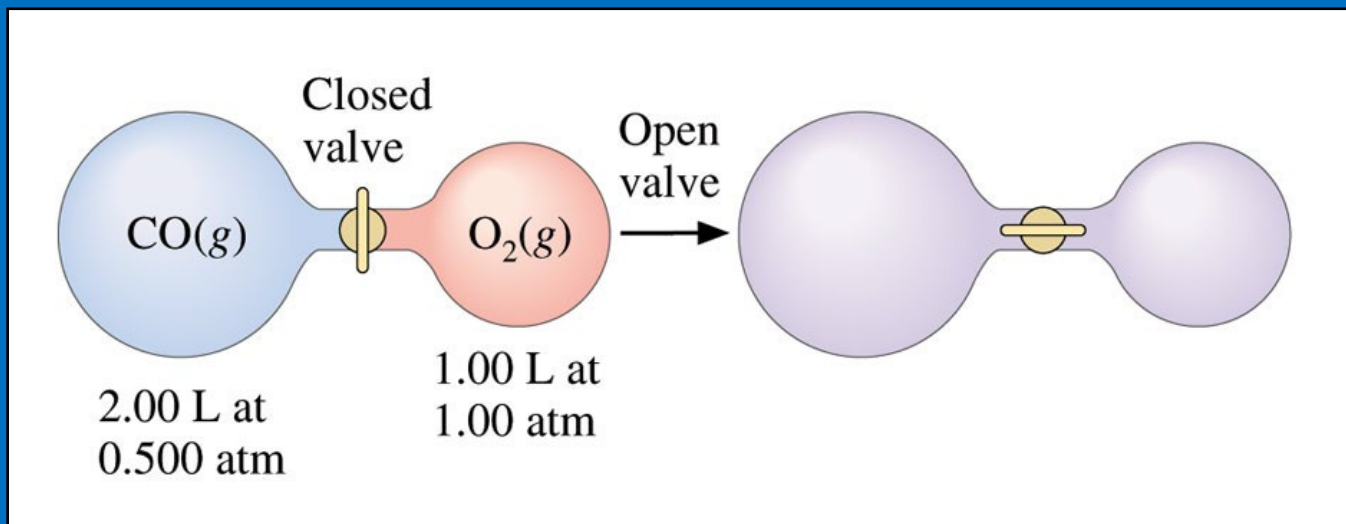
The pressure would remain the same

Practice Problem

Carbon Monoxide, CO , and Oxygen, O_2 , react according to:



Assuming that the reaction takes place and goes to completion, determine what substances remain, and what their partial pressures are after the valve is opened in the apparatus shown below. Assume that the temperature is fixed at 300 K.



Solution on next slide

Practice Problem (Con't)



Stoichiometric Molar Ratio: 2 : 1 → 2

Compute moles CO in first flask

$$PV = nRT \quad n = \frac{PV}{RT} = \frac{0.500 \text{ atm} \times 2.00 \text{ L}}{300 \text{ K} \times 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}} = 0.0406 \text{ mol}$$

Compute moles O₂ in second flask

$$PV = nRT \quad n = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 1.00 \text{ L}}{300 \text{ K} \times 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}} = 0.0406 \text{ mol}$$

$$\text{Actual Molar Ratio} \quad \frac{\text{mol CO}}{\text{mol O}_2} = \frac{0.0406}{0.0406} = 1.0$$

Actual ratio (1.0) < Stoichiometric ratio (2.0)

∴ CO is limiting

• Theoretical Yield of CO₂ is 0.0406 mol

Con't on next Slide

Practice Problem (Con't)

Amount of compounds remaining after reaction:

CO(g) - 0.0 mol (Limiting reagent completely consumed)

O₂(g) - 0.0203 mol (only half the O₂ was required to react with CO)

CO₂(g) - 0.0406 mol (based on 0.0405 mol of limiting reagent (CO))

Partial Pressure of O₂ in combined volume (3.0 L)

$$P_{\text{O}_2} = \frac{nRT}{V} = \frac{0.0203 \text{ mol} \times 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 300 \text{ K}}{3.0 \text{ L}} = 0.167 \text{ atm}$$

Partial Pressure of CO₂ in combined volume (3.0 L)

$$P_{\text{CO}_2} = \frac{nRT}{V} = \frac{0.0406 \text{ mol} \times 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 300 \text{ K}}{3.0 \text{ L}} = 0.333 \text{ atm}$$

Practice Problem

- A 100.0 mL sample of air exhaled from the lungs is analyzed and found to contain 0.0830 g of N₂, 0.0194 g O₂, 0.00640 g CO₂ and 0.00441 g water vapor at 35 °C.
- What is the partial pressure of each component and total pressure of the sample? What is the mole fraction of each component?

Calculate the moles of each gas

$$\text{mol N}_2 = 0.0830 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} = 0.00296 \text{ mol N}_2$$

$$\text{mol O}_2 = 0.0194 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.000606 \text{ mol O}_2$$

$$\text{mol CO}_2 = 0.00640 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} = 0.000145 \text{ mol CO}_2$$

$$\text{mol H}_2\text{O} = 0.00441 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.000244 \text{ mol H}_2\text{O}$$

Con't on next slide

Practice Problem

- A 100.0 mL sample of air exhaled from the lungs is analyzed and found to contain 0.0830 g of N₂, 0.0194 g O₂, 0.00640 g CO₂ and 0.00441 g water vapor at 35 °C.
- What is the partial pressure of each component and total pressure of the sample? What is the mole fraction of each component?

Calculate "Total" Moles

$$\begin{aligned}n_{\text{tot}} &= n_{\text{N}_2} + n_{\text{O}_2} + n_{\text{CO}_2} + n_{\text{H}_2\text{O}} \\&= 0.00296 \text{ mol N}_2 + 0.000606 \text{ mol O}_2 + \\&\quad 0.000145 \text{ mol CO}_2 + 0.000244 \text{ mol H}_2\text{O} \\&= 0.00396 \text{ mol}\end{aligned}$$

Calculate "Total" pressure

$$P_{\text{tot}} = \frac{n_{\text{tot}}RT}{V} = \frac{0.00396 \text{ mol} \times 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 308 \text{ }^\circ\text{K}}{100 \text{ mL} \times \frac{1\text{L}}{1000 \text{ mL}}} = 1.00 \text{ atm}$$

Con't on next slide

Practice Problem

- A 100.0 mL sample of air exhaled from the lungs is analyzed and found to contain 0.0830 g of N₂, 0.0194 g O₂, 0.00640 g CO₂ and 0.00441 g water vapor at 35 °C.
- What is the partial pressure of each component and total pressure of the sample? What is the mole fraction of each component?

Calculate mole fractions

$$X_{N_2} = \frac{n_{N_2}}{n_{tot}} = \frac{0.00296}{0.00396} = 0.747$$

$$X_{O_2} = \frac{n_{O_2}}{n_{tot}} = \frac{0.000606}{0.00396} = 0.153$$

$$X_{CO_2} = \frac{n_{CO_2}}{n_{tot}} = \frac{0.000145}{0.00396} = 0.037$$

$$X_{H_2O} = \frac{n_{H_2O}}{n_{tot}} = \frac{0.000244}{0.00396} = 0.063$$

Calculate the Partial Pressure of each gas

$$P_{N_2} = X_{N_2} * P_{tot} = 0.747 * 1.00 = 0.747 \text{ atm}$$

$$P_{O_2} = X_{O_2} * P_{tot} = 0.153 * 1.00 = 0.153 \text{ atm}$$

$$P_{CO_2} = X_{CO_2} * P_{tot} = 0.037 * 1.00 = 0.037 \text{ atm}$$

$$P_{H_2O} = X_{H_2O} * P_{tot} = 0.063 * 1.00 = 0.063 \text{ atm}$$

Practice Problem

You prepare nitrogen gas by heating Ammonium Nitrite:



If you collected the Nitrogen over water at 22 °C and 727 mm Hg, many liters of N₂ gas would be obtained from 5.68 g of NH₄NO₂?

Compute Moles of NH₄NO₂

$$\text{mol NH}_4\text{NO}_2 = n = 5.68\text{g} \times \frac{1 \text{ mol NH}_4\text{NO}_2}{64.06 \text{ g NH}_4\text{NO}_2} = 0.0887 \text{ mol NH}_4\text{NO}_2$$

Molar Ratio $\frac{\text{NH}_4\text{NO}_2}{\text{N}_2} = 1:1 \quad \therefore 0.0887 \text{ mol N}_2 \text{ produced}$

Compute Partial Pressure of Nitrogen Gas

$$P_{\text{N}_2} = P_{\text{tot}} - P_{\text{H}_2\text{O}} = 727 \text{ mm Hg} - 19.8 \text{ mm Hg} = 707.2 \text{ mm Hg}$$

$$V = \frac{nRT}{P} = \frac{0.0887 \text{ mol N}_2 \times 0.0821 \times \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 295^\circ \text{K}}{707.2 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}}} = 2.31 \text{ L}$$

Kinetic-Molecular Theory of Gases

- A simple model based on the actions of individual atoms
 - Gases consist of particles in constant motion
 - Pressure derived from bombardment with container
 - Kinetic energy described as $E_k = \frac{1}{2} mv^2$
- Postulates of Kinetic Theory
 - Volume of particles is negligible
 - Particles are in constant motion
 - No inherent attractive or repulsive forces
 - The average kinetic energy of a collection of particles is proportional to the temperature (K)

$$\overline{E}_k \propto T \quad \text{or} \quad \overline{E}_k = c \times T$$

(c is a constant that is the same for any gas)

Kinetic-theory Model of Gas Pressure

- Another approach to the Ideal-gas Equation
 - A gas consists of molecules in constant motion
 - Gas Pressure (P) is proportional to the frequency of molecular collisions and the average force exerted by a molecule in collision
 - Thus, $P \propto$ rate of collision \times avg force of molecule collision
 - “rate of collision” is proportional to the no. particles per unit volume (N/V) and their speed (u)
 - “avg force” depends on momentum ($m * u$)

$$\therefore P \propto (u \times N/V) \times mu$$
$$PV \propto Nmu^2$$

Since the average kinetic energy ($1/2 mu^2$) is proportional to Temperature (T); then $mu^2 \propto T$, i.e., energy related to Temp

The number of particles (N) can be represented as moles (n),

$$PV \propto nT$$

$$PV = nRT$$

Molecular Speeds: Diffusion and Effusion

Recall that the average kinetic energy of a collection of particles is proportional to the temperature (K)

$$\bar{E}_k \propto T \quad \text{or} \quad \bar{E}_k = c \times T$$

(c is a constant that is the same for any gas)

The constant "c" has been shown to be:

$$c = \frac{3}{2} \left(\frac{R}{N_A} \right)$$

**Where "R" is the Universal Gas Constant
"N_A" is Avogadro's Number**

$$\bar{E}_k = \frac{3}{2} \left(\frac{R}{N_A} \right) T$$

Temperature is related to the average (not "total") energy of molecular motion

Molecular Speeds: Diffusion and Effusion

From the general expression for kinetic energy of an object

$$E_k = \frac{1}{2} \text{mass} \times \text{velocity}^2 = \frac{1}{2} m v^2$$

The average kinetic energy of each molecule in a large population is:

$$\bar{E}_k = \frac{1}{2} m \bar{u}^2$$

Therefore:

$$\frac{1}{2} m \bar{u}^2 = \frac{3}{2} \left(\frac{R}{N_A} \right) T$$

$$N_A m \bar{u}^2 = 3RT$$

$N_A m = \text{Avogadro's Number} \times \text{Molecular Mass} = \text{Molar Mass (M)}$

$$\bar{u}^2 = \frac{3RT}{M}$$

Molecular Speeds: Diffusion and Effusion

The root-mean-square (rms) molecular speed, u , is a type of average molecular speed, equal to the speed of a molecule having the average molecular kinetic energy.

The square root of the previous equation gives:

$$u = \sqrt{\frac{3RT}{M_m}}$$

Unit relationships

$$R = 8.314 \text{ kg}\cdot\text{m}^2/\text{s}^2/\text{mol}\cdot\text{K}$$

$$T = \text{Temperature (K)}$$

$$M_m = \text{kg/mol (Molar Mass)}$$

$$1 \text{ Joule} = \text{kg}\cdot\text{m}^2/\text{s}^2$$

$$R = 8.314 \text{ J/mol}\cdot\text{K}$$

$$u \text{ (rms)} = \text{m/s}$$

Practice Problem

What is the ratio of the average speed of CH_4 (MW = 16 g/mol) molecules to that of SO_2 (MW = 64 g/mol) molecules at 298 K?

$$R = 8.314 \text{ J}/(\text{mol}\cdot\text{K}) = 8.314 \text{ kg}\cdot\text{m}^2/\text{s}^2/\text{mol}\cdot\text{K}$$

- a. 1:2 b. 1:1 c. 2:1 d. 3:2 e. 1:3

Ans: C $u = \sqrt{\frac{3RT}{M_m}}$ $R = 8.314 \frac{\text{kg}\cdot\text{m}^2}{\text{mol}\cdot\text{K}}$ Molar Mass (Mm) - kg / mol

$$\mu(\text{CH}_4) = \frac{\sqrt{3 \left(8.314 \frac{\text{kg}\cdot\text{m}^2}{\text{mol}\cdot\text{K}} \right) (298 \text{ }^\circ\text{K})}}{\frac{16 \text{ g CH}_4 \frac{1 \text{ kg}}{1000 \text{ g}}}{1 \text{ mol CH}_4}} = 681.6 \text{ m/s}$$

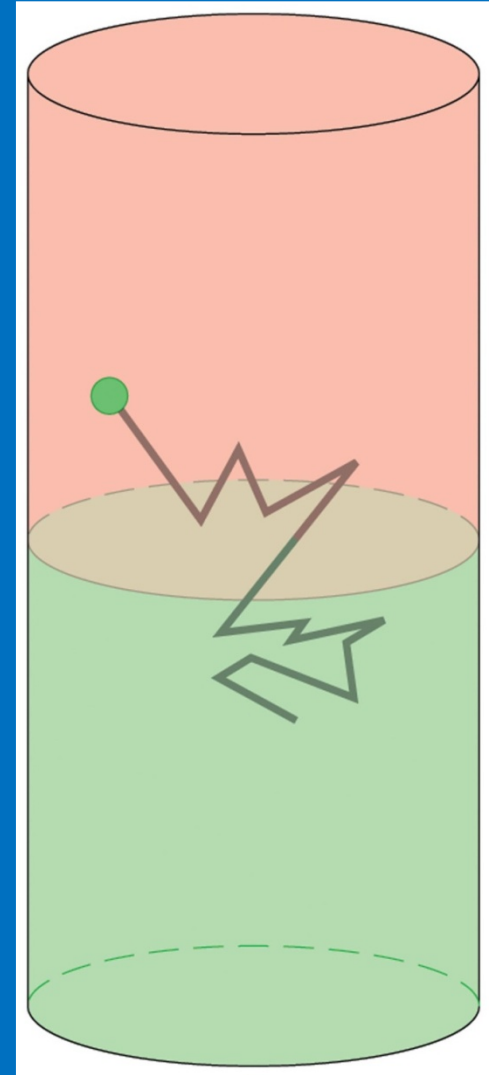
$$\frac{681.6}{340.8} = \frac{2}{1}$$

$$\mu(\text{SO}_2) = \frac{\sqrt{3 \left(8.314 \frac{\text{kg}\cdot\text{m}^2}{\text{mol}\cdot\text{K}} \right) (298 \text{ }^\circ\text{K})}}{\frac{64 \text{ g SO}_2 \frac{1 \text{ kg}}{1000 \text{ g}}}{1 \text{ mol SO}_2}} = 340.8 \text{ m/s}$$

Kinetic-theory Model of Gas Diffusion

Diffusion = process whereby a gas spreads out through another gas to occupy the space uniformly

Chemicals move from high concentration to low concentration



Molecular Speeds

Diffusion and Effusion

- **Diffusion** is the transfer of a gas through space or another gas over time
- **Effusion** is the transfer of a gas through a membrane or orifice
- The equation for the rms velocity of gases shows the following relationship between rate of effusion and molecular mass.

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{M_m}}$$

Molecular Speeds

Diffusion and Effusion

- **Effusion** is the process in which individual molecules flow through a hole without collisions between molecules.
- This occurs if the diameter of the hole is considerably smaller than the mean free path of the molecules.
- According to Graham's law, the rate at which gases effuse (i.e., how many molecules pass through the hole per second) is dependent on their molecular weight; gases with a lower molecular weight effuse more quickly than gases with a higher molecular weight.

Molecular Speeds

Diffusion and Effusion

- For two gases at the same temperature (and having the same specific heat), and thus having the same kinetic energy, the root mean square molecular speed, of each gas can be found using the equation

$$\frac{3}{2}k_b T = \frac{1}{2} \times Mm \times v_{rms}^2$$

where k_b = Boltzman Constant $k = 1.38066 \times 10^{-23} \text{ J / K}$

T = Temperature (K)

Mm = Molecular mass of particle

v = Root Mean Square Velocity (rms)

Practice Problem

Which of the following gases will have the *slowest rate of effusion*?

(assume they are all at the same temperature):

- a. CH₄ b. F₂ c. H₂ d. Ne e. SO₃

Ans: e **Rate of effusion** $\propto \frac{1}{\sqrt{M_m}}$

$$M_m \text{ CH}_4 = 12 + 4 \times 1.0 = 16 \text{ g/mol} \quad \frac{1}{\sqrt{M_m \text{ CH}_4}} = \frac{1}{\sqrt{16}} = 0.25$$

$$M_m \text{ F}_2 = 2 \times 19.00 = 38 \text{ g/mol} \quad \frac{1}{\sqrt{M_m \text{ F}}} = \frac{1}{\sqrt{38}} = 0.162$$

$$M_m \text{ H}_2 = 2 \times 1.0 = 2 \text{ g/mol} \quad \frac{1}{\sqrt{M_m \text{ H}_2}} = \frac{1}{\sqrt{2}} = 0.707$$

$$M_m \text{ Ne} = 1 \times 20.18 = 20 \text{ g/mol} \quad \frac{1}{\sqrt{M_m \text{ Ne}}} = \frac{1}{\sqrt{20}} = 0.223$$

$$M_m \text{ SO}_3 = 32 + 3 \times 16 = 80 \text{ g/mol} \quad \frac{1}{\sqrt{M_m \text{ SO}_3}} = \frac{1}{\sqrt{80}} = 0.112$$

Molecular Speeds

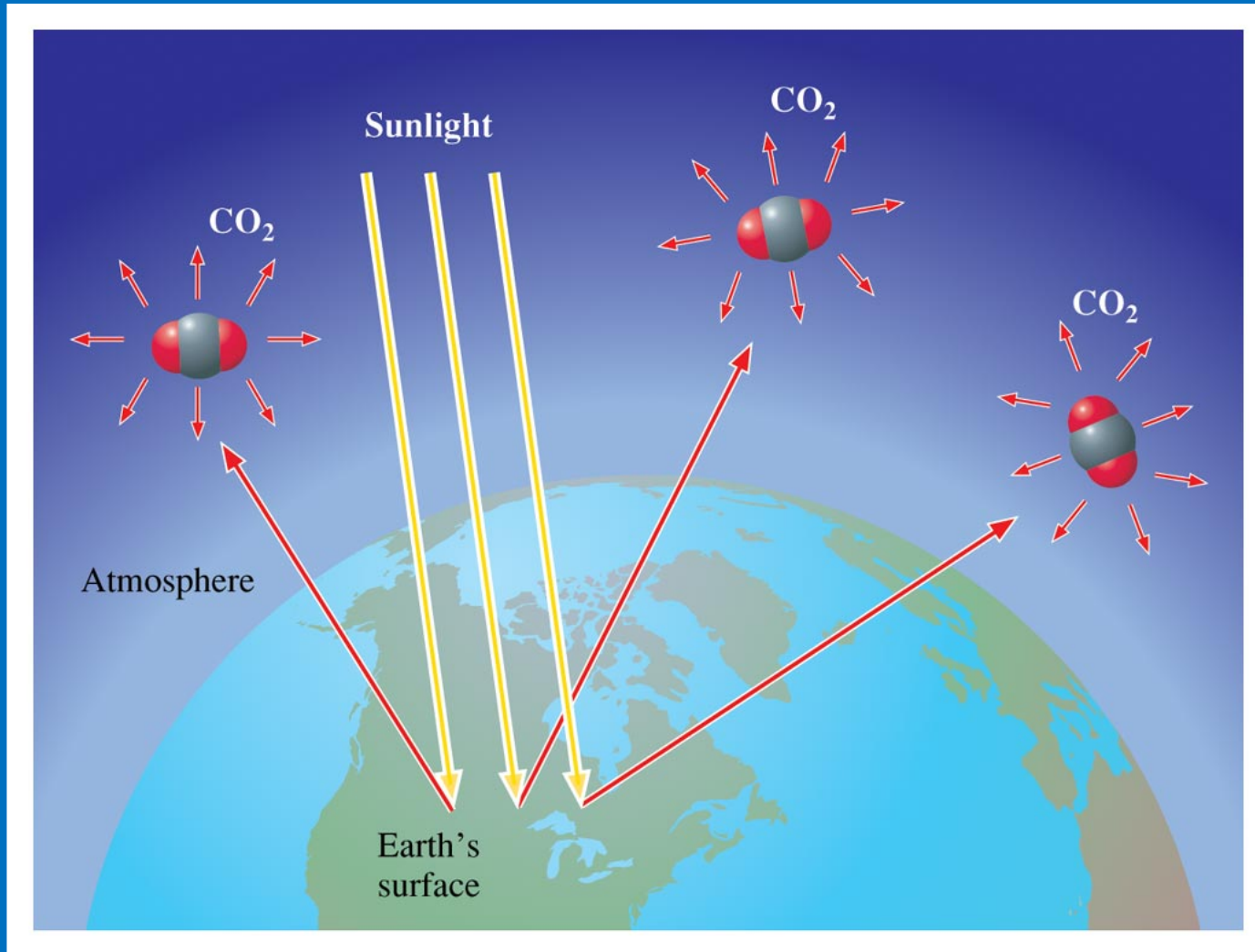
Diffusion and Effusion

Graham's law

The rate of effusion or diffusion is inversely proportional to the square root of its molecular mass

$$\frac{\text{Rate of effusion of gas "A"}}{\text{Rate of effusion of gas "B"}} = \sqrt{\frac{M_m \text{ of gas B}}{M_m \text{ of gas A}}}$$

Greenhouse Effect of Certain Gases in the Atmosphere



Real Gases

Deviations from Ideal Behavior

- Gases deviate from ideal behavior because the ideal gas law assumes that the molecules of the gas have no volume of their own, and that there are no interactions between the gas molecules
- Under **low pressure conditions**, the volume of space actually occupied by the molecules is minimized because there are fewer gas molecules present
- At **higher temperatures, i.e., above 273 °K (0°C)**, interactions between the gas molecules are minimized because the energy of the molecules is higher than the intermolecular forces of interactions between the molecules

Real Gases

Deviations from Ideal Behavior

- Real gases, as opposed to “Ideal Gases” do not act as “infinitesimal” points of mass
- Real gases have volumes determined by the sizes of their molecules, and the length and direction of their bonds
- Atoms in gases contain charged particles and many bonds are polar giving rise to attractive and repulsive forces.
- These real properties cause gases to:

Deviate from “Ideal Behavior”

Real Gases

Deviations from Ideal Behavior

- At ordinary conditions of relatively high temperature and low pressure, most gases exhibit “nearly” ideal behavior
- Even at **Standard Temperature & Pressure (STP)**, most gases deviate only slightly from ideal behavior

Gas	Molar Volume (L/mol)	Condensation Point (°C)
He	22.435	-268.9
H ₂	22.432	-252.8
Ne	22.422	-246.1
Ideal gas	22.414	---
Ar	22.397	-185.9
N ₂	22.396	-195.8
O ₂	22.390	-183.0
CO	22.388	-191.5
Cl ₂	22.184	-34.0
NH ₃	22.079	-33.4

Molar Volume of Some
Common Gases At STP

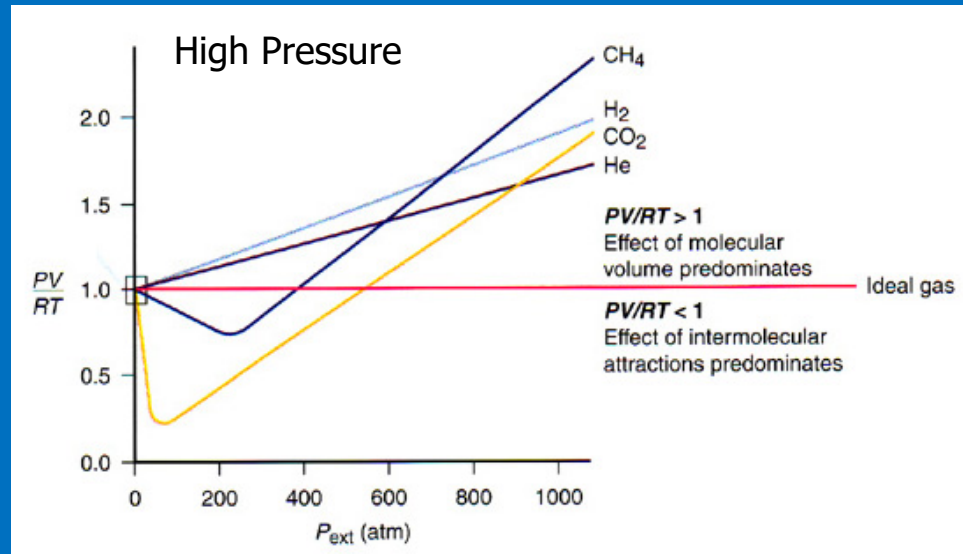
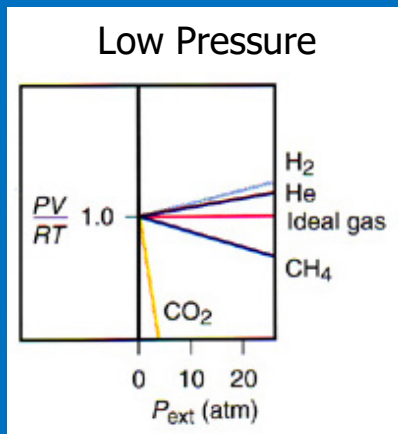
(0°C & 1 atm)

Real Gases

Deviations from Ideal Behavior

- At pressures greater than 10 atm, gases deviate significantly from ideal behavior
 - For 1 mole at moderately high pressure, values of PV/RT lower than ideal (less than "1") are due predominately to "Intermolecular Attractions"
 - At very high pressure, values of PV/RT greater than ideal (more than "1") are due primarily to

"Molecular Volume"



Real Gases

Deviations from Ideal Behavior

■ Intermolecular Attractions

- **Attractive Forces** between molecules are much **weaker** than the covalent bonding forces that hold a molecule together
- Most **Intermolecular Attractions** are caused by slight imbalances in electron distribution and are important only over relatively short distances
- At normal pressures the spaces between molecules of any real gas are **so large** that attractions are **negligible**
- As pressure rises, the intervening distances are much shorter and the attractions have a greater impact
- Increased attractions decrease force of collisions on container wall, lowering pressure, which results in a **smaller numerator in the PV/RT term ($PV/RT < 1$)**

Real Gases

Deviations from Ideal Behavior

■ Molecular Volume

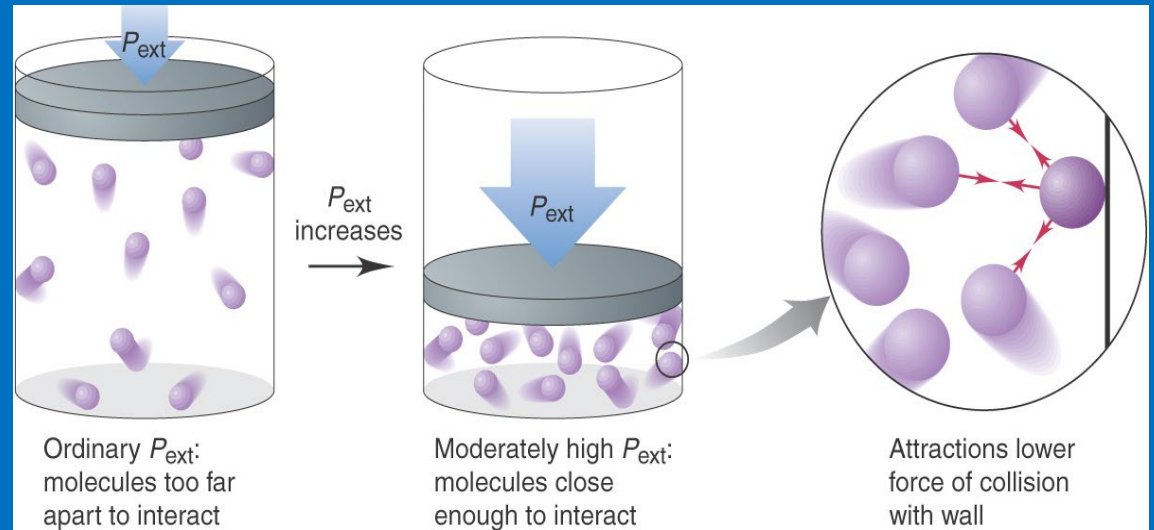
- At normal pressures, the space between molecules is very large compared to the volumes of molecules themselves (molecular volume)
- Free volume is essentially equal to the volume of the container
- At higher pressures, the free volume **decreases** as the molecular volume makes up a **greater proportion** of the container volume
- Since the volume of the container (V) does not change, the ratio of the PV/RT term is artificially high ($PV/RT > 1$)
- The molecular volume effect outweighs the intermolecular attraction effect as the pressure increases

Real Gases

Deviations from Ideal Behavior

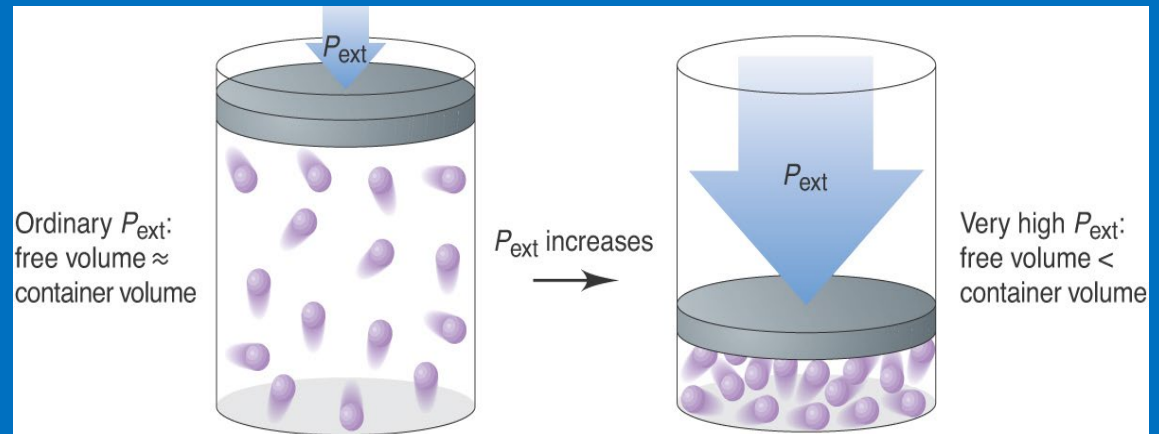
Effect of Intermolecular attractions on measured gas pressure

Intermolecular attractions decrease the force of collision with the wall, reducing actual pressure, thus the numerator in the PV/RT term is reduced ($PV/RT < 1$ for 1 mole)



At higher pressures, the free volume becomes significantly less than the container volume, thus the ratio of the PV/RT term is artificially high ($PV/RT > 1$ for 1 mole)

Effect of Molecular Volume on measured gas pressure



Van der Waals Equation

- The “Ideal Gas Law” equation must be redesigned to do two things
 - Adjust the measured ^{adjusts P up} pressure **up** by adding a factor that accounts for **intermolecular attraction**
 - Adjust the measure volume down by subtracting a factor from the entire container volume (V) that accounts for the **molecular volume**

Van der Waals Equation

■ Van der Waals Equation

$$\left(P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

adjusts
P up
adjusts
V down

- P – measured pressure
- V – measured Volume
- N – Amount (moles)
- R – Universal Gas Constant
- T – Temperature (°K)
- a & b – Van der Waals Constants

a & b = 0 for ideal gas

Van der Waals Constants (a & b)

adjusts P up

adjusts V down

Gas	$a \left(\frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2} \right)$	$b \left(\frac{\text{L}}{\text{mol}} \right)$
-----	---	--

He	0.034	0.0237
Ne	0.211	0.0171
Ar	1.35	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0511
H ₂	0.244	0.0266
N ₂	1.39	0.0391
O ₂	1.36	0.0318
Cl ₂	6.49	0.0562
CH ₄	2.25	0.0428
CO	1.45	0.0395
CO ₂	3.59	0.0427
NH ₃	4.17	0.0371
H ₂ O	5.46	0.0305

Practice Problem

A person exhales about 5.8×10^2 L of carbon dioxide (MW = 43.999) per day (at STP). The CO_2 exhaled by an astronaut is absorbed from the air of a space capsule by reaction with Lithium Hydroxide, LiOH .



How many grams of LiOH (FW = 23.943) are required per astronaut per day?

$$5.8 \times 10^2 \text{ L CO}_2 / \text{day} \left(\frac{1 \text{ mol}}{22.4 \text{ L}} \right) \left(\frac{2 \text{ mol Li(OH)}}{1 \text{ mol CO}_2} \right) \left(\frac{23.943 \text{ g Li(OH)}}{1 \text{ mol Li(OH)}} \right)$$

$$12. \text{ g Li(OH)} / \text{day}$$

Practice Problem

A sample contained a mixture of CaCO_3 and BaCO_3 . The amount of CaCO_3 was determined by reacting 7.80 g of the sample with an excess of HCl to release CO_2 from each compound:



If the sample reacted completely and produced 1,701.42 mL of CO_2 at 29.06 °C and 752.3 mm Hg, what was the mass percentage of CaCO_3 in the mixture?

Solution on next Slide

Practice Problem (Con't)

$$\text{Total Mass Sample} = m_{\text{BaCO}_3} + m_{\text{CaCO}_3} = 7.80 \text{ g}$$

$$\text{Volume CO}_2 \text{ Produced} = 1701.42 \text{ mL at } 29.06 \text{ }^\circ\text{C and } 752.3 \text{ mm Hg}$$

Compute moles CO₂ produced

$$n = \frac{PV}{RT} = \frac{752.3 \text{ mm Hg} \frac{1 \text{ atm}}{760 \text{ mm Hg}} * 1701.42 \text{ mL CO}_2 \frac{1 \text{ L}}{1000 \text{ mL}}}{0.0821 \left(\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) \times (29.06 \text{ }^\circ\text{C} + 273.15)} = 0.06788 \text{ mol CO}_2$$

$$n_{\text{CO}_2} = n_{\text{BaCO}_3} + n_{\text{CaCO}_3}$$

$$n_{\text{CO}_2} = \frac{m_{\text{BaCO}_3}}{M_{\text{BaCO}_3}} + \frac{m_{\text{CaCO}_3}}{M_{\text{CaCO}_3}}$$

$$m_{\text{BaCO}_3} = 7.80 - m_{\text{CaCO}_3}$$

$$\text{moles CO}_2 = 0.06788 \text{ mol} = \frac{7.80 \text{ g} - m_{\text{CaCO}_3}}{197.31 \text{ g}} + \frac{m_{\text{CaCO}_3}}{100.09 \text{ g}}$$
$$\frac{7.80 \text{ g} - m_{\text{CaCO}_3}}{1 \text{ mol BaCO}_3} + \frac{m_{\text{CaCO}_3}}{1 \text{ mol CaCO}_3}$$

$$m_{\text{CaCO}_3} = 5.77 \text{ g}$$

$$\% \text{ mass CaCO}_3 = (5.77 / 7.80) \times 100 = 73.8\%$$

Practice Problem

A gaseous mixture of (Propane) C_3H_8 and SO_2 has a density of 2.317 g/L at $-1.0\text{ }^\circ\text{C}$ and 746.6 mm Hg.

What is the mass percent of C_3H_8 ?

$$PV = nRT \quad PV = \frac{m}{M_m} RT \quad \frac{m}{V} = d = \frac{M_m * P}{RT}$$

$$d_{\text{mix}} = \frac{m_{C_3H_8}}{V_{\text{mix}}} + \frac{m_{SO_2}}{V_{\text{mix}}} = \frac{M_{m_{C_3H_8}} P_{C_3H_8}}{RT} + \frac{M_{m_{SO_2}} P_{SO_2}}{RT}$$

$$P_{\text{tot}} = P_{C_3H_8} + P_{SO_2}$$

$$\frac{d_{\text{mix}}}{P_{\text{tot}}} = \frac{1}{RT} \times \frac{\left(M_{m_{C_3H_8}} P_{C_3H_8} \right) + \left(M_{m_{SO_2}} P_{SO_2} \right)}{P_{\text{tot}}}$$

$$\text{mole fraction } C_3H_8 = X_{C_3H_8} = \frac{P_{C_3H_8}}{P_{\text{tot}}} \quad \text{mole fraction } SO_2 = X_{SO_2} = \frac{P_{SO_2}}{P_{\text{tot}}}$$

$$X_{C_3H_8} + X_{SO_2} = 1 \quad \therefore X_{C_3H_8} = 1 - X_{SO_2}$$

Con't on next slide

Practice Problem (Con't)

A gaseous mixture of C_3H_8 and SO_2 has a density of 2.317 g/L at $-1.0\text{ }^\circ\text{C}$ and 746.6 mm Hg.

What is the mass percent of C_3H_8 ?

$$\frac{d_{\text{mix}}}{P_{\text{tot}}} = \frac{1}{RT} \left[\frac{(M_{mC_3H_8} P_{C_3H_8}) + (M_{mSO_2} P_{SO_2})}{P_{\text{tot}}} \right]$$

$$\frac{d_{\text{mix}}}{P_{\text{tot}}} = \frac{1}{RT} \left[(M_{mC_3H_8} X_{C_3H_8}) + (M_{mSO_2} X_{SO_2}) \right]$$

$$\frac{d_{\text{mix}}}{P_{\text{tot}}} = \frac{1}{RT} \left[M_{mC_3H_8} (1 - X_{SO_2}) + M_{mSO_2} X_{SO_2} \right]$$

$M_{mC_3H_8}$ - Molar Mass C_3H_8

M_{mSO_2} - Molar Mass SO_2

$X_{C_3H_8}$ - Mole Fraction C_3H_8

X_{SO_2} - Mole Fraction SO_2

$$X_{SO_2} = \frac{\frac{d_{\text{mix}} RT}{P_{\text{tot}}} - M_{mC_3H_8}}{M_{mSO_2} - M_{mC_3H_8}}$$

$X_{C_3H_8} + X_{SO_2} = 1.0$

$X_{C_3H_8} = 1 - X_{SO_2}$

$$\frac{2.317 \text{ g/L} \times 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 272.15 \text{ }^\circ\text{K}}{746.6 \text{ mm Hg} \frac{1 \text{ atm}}{760 \text{ mm Hg}}} - 44.11 \text{ g/mol } C_3H_8$$

$$X_{SO_2} = \frac{\frac{2.317 \text{ g/L} \times 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 272.15 \text{ }^\circ\text{K}}{746.6 \text{ mm Hg} \frac{1 \text{ atm}}{760 \text{ mm Hg}}} - 44.11 \text{ g/mol } C_3H_8}{64.07 \text{ g/mol } SO_2 - 44.11 \text{ g/mol } C_3H_8}$$

$$X_{SO_2} = 0.4303$$

$$\therefore X_{C_3H_8} = 1 - 0.43 = 0.5697$$

Con't on next slide

Practice Problem (con't)

A gaseous mixture of C_3H_8 and SO_2 has a density of 2.317 g/L at $-1.0\text{ }^\circ\text{C}$ and 746.6 mm Hg.

What is the mass percent of C_3H_8 ?

Mole Fractions:

$$X_{SO_2} = 0.4303$$

$$X_{C_3H_8} = 1 - 0.4303 = 0.5697$$

Assume total moles in sample = 1

$$\therefore 0.5697 \text{ mol } C_3H_8 \quad 0.4303 \text{ mol } SO_2$$

$$\text{Mass } C_3H_8 = 0.5697 \text{ mol } C_3H_8 \times \frac{44.11 \text{ g } C_3H_8}{\text{mol } C_3H_8} = 25.1295 \text{ g } C_3H_8$$

$$\text{Mass } SO_2 = 0.4303 \text{ mol } SO_2 \times \frac{64.07 \text{ g } SO_2}{\text{mol } SO_2} = 27.5693 \text{ g } SO_2$$

$$\text{Total Mass} = 25.1295 + 27.55 = 52.5693 \text{ g}$$

$$\text{Mass \% } C_3H_8 = \frac{25.1295}{52.5693} = 47.80 \times 100 = 48\%$$

Equation Summary

Force: $F = \text{mass} \times \text{acceleration of gravity} = m \times a$

Pressure: $P = \text{force/unit area} = F / A$

Manometer liquids: $\frac{h_{\text{Decane}}}{h_{\text{Hg}}} = \frac{d_{\text{Hg}}}{d_{\text{Decane}}}$

Boyle's Law: $PV = \text{constant}$ $P_1V_1 = P_2V_2$

Charles's Law: $\frac{V}{T} = \text{constant}$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Combined Law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Avagadro's Law: $V \propto n$ $\frac{V}{n} = \text{constant}$

Ideal Gas Law: $PV = nRT$ $n = \frac{PV}{RT}$ $n = \frac{m}{M_m}$ $R = 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$

$PV = \left(\frac{m}{M_m}\right)RT$ $\frac{m}{V} = d = \frac{PM_m}{RT}$

Equation Summary

Partial Pressure: $P_{\text{tot}} = P_a + P_b + P_c + \dots$ $P_a V = n_a R T$

Root-Mean-Square Molecular Speed: $u = \sqrt{\frac{3RT}{M_m}}$

Rate of Effusion: **Rate of effusion** $\propto \frac{1}{\sqrt{M_m}}$

Grahams' Law: $\frac{\text{Rate of effusion of gas "A"}}{\text{Rate of effusion of gas "B"}} = \sqrt{\frac{M_m \text{ of Gas B}}{M_m \text{ of Gas A}}}$

Van der Waals $\left(P + \frac{n^2 a}{V^2} \right) (v - nb) = nRT$

adjusts P up **adjusts V down**