Unit 4 - Properties of Gases  
- Chapter 10

Physical Properties of Gases

Generate all of the variables that can be measured for the gas in a balloon.

Pressure
Temperature
Volume
Amount
mass, # of molecules, density, etc.

10.1 Describe measuring gas pressures using barometers and manometers. Relate pressure units.
[Readings 10.1-10.2; Problems 2. 28, 29, 30, 32, 34, &38]
Pressure

Defined as the weight (force) exerted by a gas on an area.

\[ F = mg; \text{units: } kg \, m/s^2 = F = 1 \, N \text{ (Newton)} \]

\[ 1 \, N/m^2 = 1 \, Pa \text{ (Pascal)} \]

Consider the Experiment:

Why doesn’t the water spill out?

Consider the Experiment:

Atmospheric pressure can support 34 ft. of water.

Average force exerted by the atmosphere at sea level is \( 1.01325 \times 10^5 \) Pa.

Standard Atm of Pressure
Torricelli Barometer
Use mercury (Hg) in place of water.
Hg is 13.6 times as dense.
1 atm will support 760 mmHg or 760 torr.
\[ \text{=29.92 in = 14.7 lbs/in}^2 \]
\[ \text{= 1.01325 x 10}^5 \text{ Pa} \]

Manometer
Inconvenient to use a barometer to measure the pressure of a gas sample.

\[ P_A + h = P_B \]
If the pressure in B is 735 torr and h is 20 mmHg, what is the pressure in A?

A barometer reads 745 mmHg. The left column of the manometer is 68 mm high and the right column is 52 mm high. What is the pressure of the gas sample A?

10.2 Apply the ideal gas law to relate and calculate values for pressure, volume, temperature, and amount of a gas. [Readings 10.3-10.4; Even Problems 42-70]
Pressure vs Volume

Boyle’s Law (1622)
P vs V @ constant T and amount

\[ P \propto \frac{1}{V} \text{ or } PV = \text{constant} \text{ @ constant T and amount} \]
Volume vs Temperature

Charles’ Law

\[ V \propto T \text{ (in K) } @ \text{ constant P and amount} \]

\[ V = T \cdot \text{constant} @ \text{constant P and amount} \]

Volume vs Amount

Avagadro’s Hypothesis

Equal volumes of gases @ same conditions of T & P contain same # of molecules.

\[ N_x = N_y @ \text{constant V, P, T} \]

\[ \therefore V \propto N @ \text{constant P, T} \]

Volume vs Amount

1 mole samples 
0°C and 1 atm

Average 22.4 L
Summary

Boyle’s Law
\[ V \propto \frac{1}{P} \]

Charles’ Law
\[ V \propto T \]

Avagadro’s Hypothesis
\[ V \propto n \]

Combined Gas Law Expression:
\[ V \propto \frac{nT}{P} \]
\[ \frac{PV}{nT} = \text{constant} = R \]
\[ PV = nRT \]

Ideal Gas Law

\[ \frac{PV}{nT} = \text{constant} = R \]
\[ = \frac{(1 \text{ atm})(22.4 \text{ L})}{(1 \text{ mole})(273 \text{ K})} \]
\[ = 0.0821 \text{ L-atm/mole-K} \]

\[ PV = nRT \quad R = 0.0821 \text{ L-atm/mole-K} \]

All gases behave similarly

Ideal Gas Law

Three Applications:

\[ PV = nRT \]
\[ (\text{MW}) = \frac{\text{DRT}}{P} \]
\[ (\frac{PV}{nT})_1 = (\frac{PV}{nT})_2 \]
Sample Problem

8.00 g of methane gas is burned in oxygen. What volume of carbon dioxide gas is produced at 1.00 atm and 25.0°C?

Sample Problem

The density of an unknown gas is 0.1784g/L at STP. What is the gas?

Sample Problem

A sample of gas is contained in a balloon which has a volume of 2.0 L, a pressure of 900 torr, and a temperature of 27°C. The balloon is placed in a sealed cubic box 100 cm on a side. The system is heated to 127°C and the balloon bursts. What is the pressure of the gas after the balloon bursts?
10.3 Use the ideal gas law and stoichiometric equivalencies to calculate the amount of product formed from a given amount of reactant. [Readings 10.5; Problems 72, 74, 76, 78, & 80]

Avogadro’s Principle

- Equal volumes of two ideal gases, contain the same number of particles
- \( \text{N}_2 (g) + 3 \text{H}_2 (g) \rightarrow 2 \text{NH}_3 (g) \)
- 5 \( \text{N}_2 \) molecules produce 10 \( \text{NH}_3 \) molecules
- 5.0 ml of \( \text{N}_2 \) will produce 10 ml of \( \text{NH}_3 \)
- How much hydrogen would be required to completely react with 0.158 ml of \( \text{N}_2 \)

Example Problem

- Calculate the volume of \( \text{CO}_2 \), at STP, produced from the decomposition of 152 g \( \text{CaCO}_3 \)
- \( \text{CaCO}_3 (s) \rightarrow \text{CaO} (s) + \text{CO}_2 (g) \)
10.4 Apply Dalton’s Law of partial pressure to calculate the pressure of combined gases and to calculate the partial pressures of gases in mixtures. [Readings 10.6; Problems 82, 84, 86, 88, & 90]

Combined Samples of Gases

\[ N_T = N_A + N_B \]
\[ N = PV / RT \]
\[ P_T V_T / RT = P_A V_A / RT_A + P_B V_B / RT_B \]
\[ P_T V_T / T = P_A V_A / T_A + P_B V_B / T_B \]
\[ P_T V_T = P_A V_A + P_B V_B \]
\[ P_T = P_A + P_B \]

Dalton’s Law of Partial Pressure

\[ P_T = 220 \text{ torr} = P_A + P_B \]

What is \( P_A \) and \( P_B \)?

Partial Pressure

\[ P_T = 220 \text{ torr} = P_A + P_B \]

\[ P_A = \chi_A P_T = (N_A / N_T) P_T \]
\[ P_A = \chi_A P_T = (14 / 22) 220 \text{ torr} = 140 \text{ torr} \]
\[ P_B = P_T - P_A = 220 \text{ torr} - 140 \text{ torr} = 80 \text{ torr} \]
Partial Pressures

- Calculate the pressure of a 1.0L flask which contains 0.315 g O₂ and 0.052 mol of argon.

\[
P = \frac{(0.00984 \cdot 0.315 \text{ g} + 0.052 \text{ mol}) \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \cdot 298 \text{ K}}{1.00 \text{ L}}
\]

\[= 1.513 \text{ atm}\]

Vapor Pressure

Liquids evaporate and produce a partial pressure dependent on temperature and the liquid’s nature. For example: \(P_{\text{H}_2\text{O}} \text{ @25°C} = 24 \text{ mm Hg}\)

Problem: \(\text{O}_2\) is generated into a 250 mL beaker by water displacement at a pressure of 730 torr and a temperature of 25°C. What is the volume \(\text{O}_2\) at STP?
10.5 Relate mole fractions to partial pressures [Readings 10.6; Problems practice exercise 11, and Review problems 90 & 91]

Mole Fractions

- \( P_A = \chi_A P_T = (N_A/N_T) P_T \)
- \( \chi_A = (N_A/N_T) \)
- The mole fraction
- \( \chi_A = P_A / P_T \)

Example Problem

- A 4.00 L divers tank contains 9.54 atm partial pressure of oxygen and a total pressure of 42.3 atm. What is the mole fraction of oxygen in the tank?
10.6 Relate MW and speeds of molecules using Graham’s law. [Readings 10.7; Problems 80 & 82] 

Effusion 
Definition: Passage of a gas through a tiny orifice into an evacuated chamber. The rate of effusion is the speed at which a gas does this.

Graham’s Law 
For all gases: \( KE = \frac{1}{2} m u^2 = cT \) 
∴ if two gases are at the same temperature: 
\[
\frac{1}{2} m_A u_A^2 = \frac{1}{2} m_B u_B^2 = cT \\
\frac{u_A^2}{u_B^2} = \frac{m_B}{m_A}
\]
If one mole of each gas is in sample: 
\[
\frac{u_A^2}{u_B^2} = \frac{MW_B}{MW_A}
\] or
Graham’s Law

\[ \frac{\text{effusion rate}_A}{\text{effusion rate}_B} = \frac{M_B}{M_A} \]

Graham’s Law

\[ \frac{\text{effusion rate}_A}{\text{effusion rate}_B} = \frac{\sqrt{M_B}}{M_A} \]

Mₐ and Mₐ are the molecular masses of gases A and B.

Problem

Find the ratio of effusion rate for H₂ and air.
Example Problem

• HCl effuses at a rate which is 1.5 times that of another hydro-halogen gas, HX. Identify which gas HX would be.

Diffusion

Definition: The mixing of gases. The rate of diffusion is the speed of the mixing process.
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10.7 Describe gases in terms of KMT. [Readings 9.6; Problems 8, 73, 75, 77]

Kinetic Molecular Theory (KMT)

Explain ideal gas law regularity using the atomic/molecular model. Model: KMT (Kinetic Energy)
KMT Assumptions
attraction & repelling forces between molecules are negligible
molecules have mass but size is negligible

Postulates of the KMT
1. Gases consist of molecules in continuous random motion.
2. The volume of the particles is negligible compared to the total volume
3. Molecular collisions are elastic.
4. Gas particles are independent of each other

Kinetic Molecular Theory
- Pressure – caused by the collisions of gas atoms, or molecules with the walls of the container
- Pressure increases with the number and or force of the collisions
Kinetic Theory - Pressure-Volume

Explain why pressure is higher in right cylinder.

Kinetic Theory - Volume-Pressure

Click picture to view animation

1. Before viewing each animation predict what you expect to see.
2. Explain how changing the volume causes the pressure to change.
3. How is the rate of collisions with the walls related to pressure?

Kinetic Theory - Temperature-Pressure

Explain why pressure in higher in right bulb.
Kinetic Theory - Temperature-Pressure

1. Before viewing each animation predict what you expect to see.
2. Explain how changing the temperature causes the pressure to change.
3. How is the rate of collisions with the walls related to pressure?

Kinetic Theory - Temperature-Volume

Why does heating a gas expand its volume at constant P?

Increase the Amount

\[ P = nC_v \text{ @ constant } V \text{ and } T \]
Increase the Amount

\[ P = nC_3 \text{ @ constant } V \text{ and } T \]

10.8 Distinguish between ideal and real gases.
[Readings 10.9; Problems 24-27]

Real Gases

Real gases do not obey the ideal gas law, \( PV = nRT \), due to their attractive forces and the finite volumes of their molecules.

1. How do real gas pressures and volumes differ from ideal gas pressures and volumes?
2. Which of these properties of real gases dominates below and above 350 atm for oxygen?
Real Gases

Assumptions

1. Attraction & repelling forces between molecules are negligible.
2. Molecules have mass but size is negligible.

∴ All ideal gases behave the same.

What conditions favor the ideal behavior of gases?
When are the assumptions not good assumptions?

Assumption 1 - Minimize Interaction

What conditions favor the ideal behavior of gases?

When are the assumptions not good assumptions?

Assumption 1 - Minimize Interaction

What conditions favor the ideal behavior of gases?

When are the assumptions not good assumptions?
Assumption 1 - Minimize Interaction

What conditions favor the ideal behavior of gases?
High Temperature

When are the assumptions not good assumptions?
at Low Temperatures

Assumption 2 - Negligible Size

Area of screen - 60,000 cm$^2$

Area of O$_2$ if screen contained 2 molecules
@STP - 12 cm$^2$

Area of H$_2$ if screen contained 2 molecules
@STP - 6 cm$^2$

Assumption 2 - Negligible Size

H$_2$ gas @STP
0.01% of space is H$_2$ - 99.99% empty space
Assumption 2 - Negligible Size

O₂ gas @STP
0.02% of space is O₂ - 99.98% empty space
It makes little difference which gas is used.

Assumption 2 - Negligible Size

H₂ gas @High P
9.4% of space is H₂ - 90.6% empty space

Assumption 2 - Negligible Size

O₂ gas @High P
19% of space is O₂ - 81% empty space
Assumption 2 - Negligible Size

What conditions favor the ideal behavior of gases?

When are the assumptions not good assumptions?

van der Waals Equation

\[(P + \frac{a{n^2}}{V^2})(V-nb) = nRT\]