Chem 103

CHAPTER 10
Gases

د. جميل ملكاوي

Chapter 10
Gases

Properties of Gases:

1. Low molar mass and simple molecular formulas.
   e.g., O_2, N_2, F_2, Cl_2, CH_4, NH_3, …
2. Nonmetals with covalent bonds.
3. Has no shape or volume
4. Form homogeneous mixtures.
5. Molecules are far apart.

Pressure:

Pressure: The force, F, that act on a given area, A.

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

Units of Pressure:

$$P = \frac{N}{m^2}$$

$$= N / m^2$$

$$= Nm^2$$

$$= Pascal(Pa)$$

1kPa = 1000 Pa

1 bar =$1x10^5$ Pa = 1000 kPa

Atmospheric Pressure:

Atmospheric pressure: is the pressure exerted by earth's atmosphere.

Barometer: A device used to measure atmospheric pressure.

The pressure of a column of liquid is given by:

$$P = \rho g h$$

ρ = the density of the liquid.
g = Gravitational acceleration.
h = height of the column.

Pressure of Enclosed Gases:

Manometer:

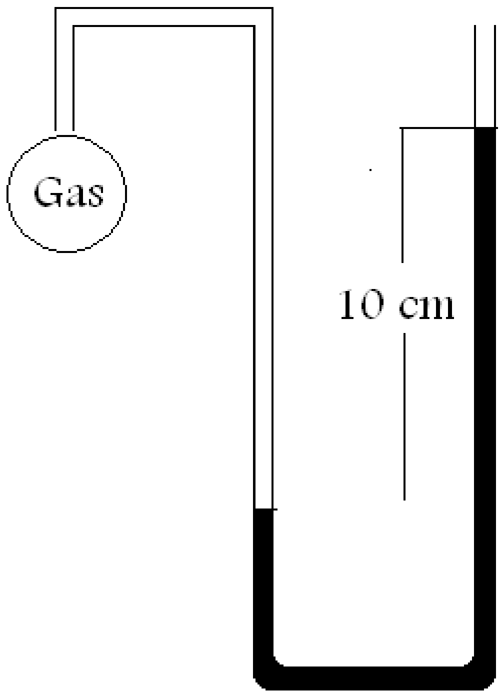
1. Open End Manometer.
2. Closed End Manometer.

Example: The pressure of a gas was measured at 750 torr with an open-end manometer filled with mercury. The result was as shown in the figure bellow. Calculate the gas pressure.

$P_g > P_{atm}$

$P_g = P_{atm} + P_{Hg}$

$P_{atm} = 750$ torr



$P_{Hg} = 10$ cm Hg = 100 mm Hg = 100 torr

$P_g = 750 + 100 = 850$ torr

The Gas Laws:

The properties of Gases depends on the Variables:

1. Pressure
2. Volume
3. Temperature
4. Amount of Gas

1. The Pressure-volume Relationship: Boyle's Law

$$y = ax$$

$$V = \text{constant} \times \frac{1}{P}$$

$$or$$

$$PV = \text{constant}$$

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

$$PV$$

Boyle's Law: The volume of a fixed quantity of gas maintained at constant temperature is inversely proportional to the pressure of the gas.

2. The Temperature-Volume Relationship: Charle’s and Gay-Lussac’s Law

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

$$V \; \alpha \; T$$

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

3. The Quantity-Volume Relationship: Avogadros Law

<u>Gay-Lussac Law of Combining Volumes</u>: At a given pressure and temperature, the volumes of gases that react with one another are in the ratios of small whole numbers.

<u>Avogadro’s Hypothesis:</u> Equal volumes of gases at the same temperature and pressure contain equal number of molecules.

22.4 L of any gas at STPcontain Avogadro’s number of molecule 22.4 L of any gas at STPcontains imole (6.022×10^{23} molecule)

$$V \alpha\ n$$
$$V = \text{constant} \times n$$
$$or$$
$$\frac{V}{n} = \text{constant}$$

<u>The Ideal Gas Law</u>:

$$V \alpha \frac{1}{P}$$
$$V \alpha T$$
$$V \alpha n$$

$$V \alpha \frac{nT}{P}$$

$$V = R\left(\frac{nT}{P}\right)$$

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

$$PV = nRT$$

Ideal Gas Law: PV = nRT

Ideal Gas: a hypothetical gas whose pressure, volume and temperature are completely described by the ideal gas law.

The Gas Constant R:

R = 0.08206 L atm/K mol
= 8.314 Pa m^3/K mol
= 8.314 J/K mol

Standard Temperature and Pressure (STP)

Standard Temperature = 0 ^oC = 273 K
Standard Pressure = 1 atm = 760 torr

Molar Volume of a gas: is the volume of 1 mol of a gas.

At STP the molar volume of a gas

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

$$V = \frac{1\, mol \times 0.0821\, L\ atm / K\ mol \times 273K}{1\, atm}$$

$$V = 22.41\, L$$

Example: The pressure of a gas in a can 1.5 atm at 25 ^{o}C. What would be the pressure if the can is heated to 400 ^{o}C?

$$\frac{P}{T} = \frac{nR}{V} = cons\tan t$$

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

$$\frac{1.5}{298} = \frac{P_2}{673}$$

$$P_2 = 3.4 \quad atm$$

Example: A balloon has a volume of 6 L at 1 atm and 22 ^{o}C. What is the volume of the balloon at 0.45 atm at –21 ^{o}C?

$$\frac{PV}{T} = nR = cons\tan t$$

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

$$\frac{1\times 6}{295} = \frac{0.45\times V_2}{252}$$

$$V_2 = 11.4 \quad L$$

Molar Mass of Gases:

$$PV = nRT$$

$$n = \frac{m}{M}$$

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

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

Density of Gases:

The density of a gas with molar mass = M

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

$$M = \left(\frac{m}{V}\right)\frac{RT}{P}$$

$$M = \frac{dRT}{P}$$

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

Example: What is the density of CCl_4 (M=154 g/mol) vapor at 714 torr and 125 ^{o}C?

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

$$d = \frac{(\frac{714}{760})\ atm \times 154\ \frac{g}{mol}}{0.0821\ \frac{atm\ L}{K\ mol} \times 398\ K} = 4.43\ g/L$$

<u>Gas Stoichiometry: Volume of Gases In Chemical Reactions</u>

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

Example: Car airbags are filled N_2 by the reaction:

$$2NaN_{3(s)} \rightarrow 2Na_{(s)} + 3N_2$$

If an air bag has a volume of 36 L is to be filled N_2 at 1.15 atm and 26 ^{o}C. How many grams of NaN_3 must be decomposed?

Mole s of N_2 needed to fill the page:

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{1.15 \times 36}{0.0821 \times 299} = 1.7\ mol$$

$$\text{mole of NaN}_3 \text{ needed} = 1.7 \text{ mol N}_2 \times \frac{2\ mol\ NaN_3}{3\ mol\ N_{23}} = 1.13\ mol$$

$$\text{grams of NaN}_3 \text{ needed} = 1.13\ mol\ NaN_3 \times \frac{65\ g\ \text{NaN}_3}{1\ mol\ \text{NaN}_3} = 73\ g$$

Gas Mixtures: Dalton's Law of Partial Pressures

John Dalton:

P_t= Total pressure
P_1= Partial pressure of gas 1

$$P_t = P_1 + P_2 + P_3 + ...$$

$$P_1 = \frac{n_1 RT}{V}$$

$$P_2 = \frac{n_2 RT}{V}$$

$$P_t = (n_1 + n_2 + n_3 + ...)\frac{RT}{V}$$

$$P_t = \frac{n_t RT}{V}$$

Partial Pressure and Mole Fraction:

$$\frac{P_1}{P_t} = \frac{n_1 \frac{RT}{V}}{n_t \frac{RT}{V}}$$

$$\frac{P_1}{P_t} = \frac{n_1}{n_t}$$

$$\frac{P_1}{P_t} = X_1$$

$$P_1 = X_1 P_t$$

Where X_1 = mole fraction of gas 1

In general

$$P_i = X_i P_i$$

Collecting Gas Over Liquid:

Example: A sample of $KClO_3$ was decomposed and the evolved O_2 was collected over water according to the reaction:

$$2KClO_{3(s)} \longrightarrow 2KCl_{(s)} + 3O_{2(g)}$$

The volume of gas collected was 250 mL at 765 torr and 26 ^oC. Calculate:

A. The number of moles of O_2 collected if the vapor pressure of water at 26 ^oC is 25 torr.
B. The grams of $KClO_3$ decomposed.

V = 250 mL = 0.25 L
T = 26 ^oC = 299 K
P = 765 –25 = 740 torr = 0.974 atm

$$n_{O2} = \frac{PV}{RT}$$

$$n_{O2} = \frac{0.974 \times 0.25}{0.0821 \times 299} = 9.9 \times 10^{-3} mol$$

$$g\ KClO_3 = 9.9 \times 10^{-3}\ mol\ O_2 \times \frac{2\ mol\ KClO_3}{3\ mol\ O_3} \times \frac{122.6\ g\ KClO_3}{1\ mol\ KClO_3}$$

$$g\ KClO_3 = 0.811$$

Molecular Diffusion and Effusion:

Diffusion: The spread of a substance through space.

The average speed of molecules (u)

$$u \alpha \frac{1}{\sqrt{M}}$$

Lighter gaseous molecules diffuses faster than heavy molecules.

Effusion: The escape of gas molecules through tiny holes.

Rate of effusion (r)

$$r \alpha \frac{1}{\sqrt{M}}$$

ROOT MEAN SQUARE SPEED (u_{rms}):

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

To compare the rate of effusion of different

$$\frac{r_2}{r_1} = \sqrt{\frac{M_1}{M_2}}$$

Example: An unknown gas composed of homonuclear diatomic molecules effuses at a rate that is only 0.355 times that of oxygen at the same temperature. What is the formula of the unknown gas?

$$\frac{r_x}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_x}}$$

$$\frac{r_x}{r_{O_2}} = 0.355$$

$$0.355 = \sqrt{\frac{32}{M_x}}$$

$$(0.355)^2 = \frac{32}{M_x}$$

$$M_x = \frac{32}{(0.355)^2} = 254\ g/mol$$