Gases (Chapter 5)

Important Properties of a Gas

Quantity: n = moles

Volume: V = container size (usually L or mL)

Temperature: $T \approx$ average kinetic energy of molecules
(must be in K for all "gas laws")

Pressure: P = force/area

Units of Pressure: SI unit is the ***pascal*** (Pa)

1 atm = 101,325 Pa (not commonly used)

More important:

1 mm Hg = 1 torr

1 atm = 760 torr = 760 mm Hg

Pressure - Volume - Temperature Relationships

1. Boyle's Law (at constant T and n)

   $V \propto 1/P$ or PV = constant

2. Charles' Law (at constant P and n)

   $V \propto T$ or V / T = constant

3. Gay-Lussac's Law (at constant V and n)

   $P \propto T$ or P / T = constant

4. Combined Gas Law (for constant n)

PV / T = constant or

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

(remember that T must be in units of K -- practice problems in book!)

Ideal Gas Law

1. **Avogadro's Principle**

at constant P and T, $V \propto n$

i.e., at constant T and P, equal volumes of gases contain equal numbers of moles

2. Standard Molar Volume

at Standard Temperature and Pressure (0℃ and 1 atm),

1 mole of any gas occupies 22.4 L (i.e., **22.4 L / mole**)

3. The **Ideal Gas Equation**

$$\boxed{\mathbf{PV = nRT}}$$

where R = "universal gas constant"

= 0.0821 L·atm / mole·K **memorize !**

{ Useful in many different kinds of calculations involving gases!!! }

Problem:

At STP, the density of a certain gas is 4.29 g/L. What is the molecular mass of the gas?

$$(4.29 \text{ g / L}) \times (22.4 \text{ L / mole}) = 96.0 \text{ g / mole}$$

Problem:

Acetylene (welding gas), C_2H_2, is produced by hydrolysis of calcium carbide.

$$CaC_{2\,(s)} + 2\,H_2O \longrightarrow Ca(OH)_{2\,(s)} + C_2H_{2\,(g)}$$

Starting with 50.0 g of CaC_2, what is the theoretical yield of acetylene in liters, collected at 24°C and a pressure of 745 torr?

1st find yield in **moles**:

$$50.0 \text{ g } CaC_2 \times \frac{1 \text{ mole } CaC_2}{64.10 \text{ g } CaC_2} \times \frac{1 \text{ mole } C_2H_2}{1 \text{ mole } CaC_2} = 0.780 \text{ mole } C_2H_2$$

now use ideal gas law to find **volume** of C_2H_2:

$$PV = nRT \quad \Longrightarrow \quad V = \frac{nRT}{P}$$

$$V = \frac{(0.780 \text{ mole}) \times (0.0821 \text{ L atm / mole K}) \times (297 \text{ K})}{(745 \text{ torr}) \times (1 \text{ atm / } 760 \text{ torr})} = \boxed{19.4 \text{ L}}$$

Dalton's Law of Partial Pressures

For a mixture of gases: $P_{total} = P_a + P_b + P_c + \ldots\ldots$

Gases are often prepared and **collected over water**:

$$P_{total} = P_{gas} + P_{water}$$

where P_{water} = ***vapor pressure*** of water (depends on temperature)

e.g., at 25°C, $P_{water} = 23.8$ torr

at 50°C, $P_{water} = 92.5$ torr

Problem:

A sample of N_2 gas was prepared and collected over water at 15℃. The total pressure of the gas was 745 torr in a volume of 310 mL. Calculate the mass of N_2 in grams.

$$P_{total} = P_{gas} + P_{water}$$

$$745 \text{ torr} = P_{gas} + 12.8 \text{ torr}$$

$$P_{gas} = 732.2 \text{ torr}$$

$$PV = nRT \quad \Longrightarrow \quad n = \frac{PV}{RT}$$

$$n = \frac{(732.2 \text{ torr}) \times (1 \text{ atm} / 760 \text{ torr}) \times (0.310 \text{ L})}{(0.0821 \text{ L atm} / \text{mole K}) \times (288 \text{ K})} = 0.0126 \text{ mole } N_2$$

$$\text{mass } N_2 = (0.0126 \text{ mole } N_2) \times (28.0 \text{ g } N_2 / \text{mole } N_2) = \boxed{0.354 \text{ g } N_2}$$

Graham's Law of Effusion

diffusion: "mixing" of gases throughout a given volume

effusion: "leaking" of a gas through a small opening

Graham's Law: effusion rate $\propto 1/\sqrt{d}$ where d = density

but, $d \propto FM$ (formula mass)

so, effusion rates of two gases can be compared as a proportion:

$$\frac{ER_a}{ER_b} = \sqrt{\frac{FM_b}{FM_a}}$$

e.g., He (FM = 4.0 g/mole) effuses 2 times faster than CH_4 (FM = 16.0)

Kinetic Theory of Gases -- READ BOOK

Basic Postulate: An gas consists of a very large number of very small particles, in constant random motion, which undergo perfectly elastic collisions with each other and the container walls.

There is a distribution of kinetic energies of the particles (Figure 5.19)

$$\text{Temp} \propto \text{average KE}$$

The kinetic theory "explains" the gas laws, pressure, etc. based on motion and kinetic energy of gas molecules.

e.g., **Boyle's Law** ($P \propto 1 / V$) -- at constant Temp (same average KE)

If volume of container is reduced, there are more gas particles per unit volume, thus, more collisions with the container walls per unit area.
$\therefore$ higher pressure

Real Gases -- Deviations from Ideal Gas Law

For real gases, small corrections can be made to account for:

Actual volume of the gas particles themselves, and
Intermolecular attractive forces

One common approach is to use the ***Van der Waals' Equation***:

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

Don't memorize !

Where, a and b are empirical parameters that are dependent on the specific gas (e.g., Table 5.5).

a $\approx$ intermolecular attractive forces

b $\approx$ molecular size