

Chapter 5 - Gasses

Section 5.1 – Pressure

Barometer -A device which can measure atmospheric pressure.

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr} = 101,325 \text{ Pa} = 14.7 \text{ lb/in}^2$$

$$\text{Pressure} = \text{Force/Area}$$

Section 5.2 - The Gas Laws

1. **Boyles Law** - The inverse relationship which exists between pressure and volume at a constant temperature. Volume is directly proportional to the inverse of pressure.

$$P_1V_1 = k = P_2V_2$$

Ideal Gas - If a gas behaves like Boyles Law

2. **Charles Law** - The volume of a gas at constant pressure increases linearly with the temperature of the gas. V is directly proportional to T(expressed in K) at constant pressure.

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

Kelvin (K) - is a temperature measurement based on **absolute zero**.

$$K = 273 + \text{Celsius temperature}$$

3. **Avogadros Law** -For a gas at Constant temperature & Pressure, the volume (V) is directly proportional to the number of moles (n) of gas.

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

Section 5.3 - The Ideal Gas Law

This equation is a combination of the above laws into one law. This works for any rule.

$$PV = nRT$$

Where "**R**" is the Universal Gas Constant:

$$R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$$

Section 5.4 - Gas Stoichiometry

Molar Volume -The volume of an ideal gas at **STP(Standard Temperature and Pressure)**.

1 Mol of any gas at STP will have a Molar Volume of **22.42 Liters**.

Molar Mass - can be derived by manipulating the ideal gas law by using the measured density of the gas.

$$\text{Molar Mass} = \frac{dRT}{P}$$

Where "**d**" is the density of the gas in grams/liter.

Section 5.5 - Dalton's Law of Partial Pressure

Partial Pressures - For a mixture of gasses in a container, the total pressure exerted is the sum of the pressures that each gas would exert if it were alone.

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots$$

Mole Fraction (χ) - The ratio of the number of moles of a given component in a mixture to the total number of moles in the mixture. For ideal gasses, they are directly related to it's partial pressure.

$$c_1 = \frac{n_1}{n_{\text{Total}}} = \frac{P_1}{P_{\text{Total}}}$$

The partial pressure of a particular component of a gaseous mixture is the mole fraction of that component times the total pressure.

$$P_1 = c_1 \times P_{\text{Total}}$$

Section 5.6 - The Kinetic Molecular Theory of Gasses

Kinetic Molecular Theory (KMT) - A simple model that attempts to explain the properties of an ideal gas.

1. The particles are so small compared with the distances between them that the volume of the individual particles can be assumed to be negligible (zero).
2. The particles are in constant motion, the collisions of the particles with the walls of the container are the cause of the pressure exerted by the gas.
3. The particles are assumed to exert no force on each other; they are assumed neither to attract nor repel each other.
4. The average kinetic energy of a collection of gas particles is assumed to be directly proportional to the Kelvin temperature of the gas.

The Meaning of Temperature - The Kelvin temperature indicates the **average kinetic energy (KE)** of the gas particles.

$$(\text{KE})_{\text{avg}} = \frac{3}{2} RT$$

Root Mean Square Velocity - the square root of the average of the squares of the individual velocities of gas particles.

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

"**M**" is a mole of gas particles in a kilogram.

"**R**" is converted into **joules**, so it has the value of 8.3145 J/Kmol

Section 5.7 – Effusion and Diffusion

Diffusion - The term used to describe the mixing of gasses.

Effusion - The term used to describe the passage of a gas through a tiny orifice into an evacuated chamber. The rate of effusion measures the speed at which the gas is transferred into the chamber.

Graham's Law of Effusion - The rate of effusion of a gas is inversely proportional to the square root of the mass of its particles.

$$\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Where M_1 and M_2 represent the molar masses of the gases.

Rate of effusion - Measures the speed at which the gas is transferred into the chamber and is equal to the square root of the molar mass of the gas. Lighter molecules move faster than the larger molecules.

Section 5.8 – Real Gasses

Ideal Gas Behavior - The behavior approached by the real gasses under certain conditions. This a hypothetical concept. No gas exactly follows the ideal gas law although many gases come very close at low pressures and/or high temperatures.

Real Gases - We can approximate the behavior of real gases using **Van Der Waals Equation**.

$$\left[P_{\text{obs}} + a \left(\frac{n}{V} \right)^2 \right] \times (V - nb) = nRT$$

Where "a" is a correction factor for pressure and "b" is a correction factor for volume of a given gas.
(See table 5.3 on page 223)

****Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.**