Unit 8 Gas Laws

Volume - space matter occupies; for gases it is measured in liters.

Pressure - force which acts on an object due to a gas; SI unit for pressure is pascal (N/m^2).

Boyle's Law - pressure and volume of a gas are inversely proportional; as the pressure rises, then the volume decreases and vice-versa.

$$P_1 \times V_1 = P_2 \times V_2$$

Charles' Law - temperature and volume of a gas are directly proportional; as the temperature increases, the volume increases.

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

Assumptions of Kinetic Molecular Theory and Gases
1. Gas particles are very small and the space between them is huge, so the volume of gas particles is insignificant.
2. Since the gas particles are so small relative to their volume, they do not attract or repel each other. So, gases expand to fill all space available.
3. The movement of gas particles is constant and random. Collisions are rare and elastic, i.e. energy is transferred from one particle to another.

Compressibility- how much the volume of matter decreases under pressure. Gases have large amounts of space between them, but liquids and solids do not.

Units of Gases

Pressure (**P**)- kilopascals (kPa)   Volume (**V**)- liters (L)

Temperature (**T**)- kelvins (K)   number of moles (**n**)

Example 1: A high-altitude balloon contains 30.0 liters of helium gas at 103 kPa. What is the volume when the balloon rises to an altitude where the pressure is only 25.0 kPa?

$$P_1 = 103 \text{ kPa} \qquad P_2 = 25.0 \text{ kPa}$$
$$V_1 = 30.0 \text{ L} \qquad V_2 = ?$$

$$P_1 \times V_1 = P_2 \times V_2 \quad \text{becomes} \quad V_2 = \frac{P_1 \times V_1}{P_2}$$

$$V_2 = \frac{103.0 \, kPa \times 30.0 \, L}{25.0 \, kPa} = 1.24 \times 10^2 \, L$$

Example 2: A balloon inflated in a room at 24°C has a volume of 4.00 L. The balloon is then heated to a temperature of 58°C. What is the new volume if the pressure remains constant?

$$V_1 = 4.00 \text{ L} \qquad V_2 = ?$$
$$T_1 = 24°C \text{ (or 297K)} \qquad T_2 = 58°C \text{ (or 331K)}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{becomes} \quad V_2 = \frac{V_1 x T_2}{T_1}$$

$$V_2 = \frac{V_1 \times T_2}{T_1} = \frac{4.00 \, L \times 331 \, K}{297 \, K} = 4.46 \, L$$

Gay-Lussac's Law- the pressure of a gas is directly proportional to the kelvin temperature if the volume remains constant.

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

Example 3: The gas left in a used aerosol can is at a pressure of 103 kPa at 25°C. If this can is thrown in a fire, what is the pressure of the gas when its temperature reaches 928°C ?

$$P_1 = 103 \text{ kPa} \qquad P_2 = ?$$
$$T_1 = 25°C \ (298K) \qquad T_2 = 928°C \ (1201K)$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad \text{becomes} \qquad P_2 = \frac{P_1 \ x \ T_2}{T_1}$$

$$P_2 = \frac{103 \, kPa \ x \ 1201 \, K}{298 \, K} = 415 \, kPa$$

Combined Gas law- If anything is kept constant, you will have one of the previous laws.

Avogadro's law

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

$$1 \, mole \ of \ any \ gas \ @ \ STP = 22.4 \, L \quad and \quad \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Finding the Ideal Gas law- Real gases differ from ideal gases because intermolecular forces tend to reduce the distance between real gas particles and because real gas particles have volume.

$$\frac{P_1 \ x \ V_1}{n_1 \ x \ T_1} = \frac{P_2 \ x \ V_2}{n_2 \ x \ T_2}$$

An ideal gas has an answer which is always a constant, **R**

$$R = \frac{P \ x \ V}{T \ x \ mol} = \frac{101.3 \, kPa \ x \ 22.4 \, L}{273 \, K \ x \, 1 \, mol} = 8.31 \frac{kPa - L}{K - mol}$$

note: R = 8.31 kPa-L/mol-K = 0.0821 atm-L/mol-K = 62.4 mm Hg-L/mol-K

If P x V / n x T = R, then the **Ideal Gas Law** becomes **PV=nRT**

Example 4: You fill a rigid steel cylinder with a volume of 20.0 L with N_2 gas to a final pressure of 2.00×10^4 kPa at 28°C. How many moles of $N_2(g)$ does the cylinder contain?

P = 2.00×10^4 kPa   V = 20.0 L   T = 28°C (301K)   n = ?

$$PV = nRT \quad \therefore \quad n = \frac{P \times V}{R \times T} = \frac{2.00 \times 10^4 \; kPa \; x \; 2.0 \times 10^1 L}{\frac{8.31 \, L-kPa}{K-mol} \; x \; 301 \, K} = 1.60 \times 10^2 \, mol$$

Dalton's Law of Partial Pressures- Total pressure is the sum of the partial pressure of each gas present.

Example 5: Air contains O_2, N_2, CO_2, and trace amounts of other gases. What is the partial pressure of O_2 at 101.3 kPa of total pressure if the partial pressures of N_2, CO_2 and other gases are 79.10 kPa, 0.040 kPa, and 0.94 kPa respectively?

$$P_{total} = P_{N_2} + P_{O_2} + P_{CO_2} + P_{others} \quad \therefore \quad P_{O_2} = P_{total} - (P_{N_2} + P_{CO_2} + P_{others})$$

$$P_{O_2} = 101.3 kPa - (79.10 \, kPa + 0.040 \, kPa + 0.94 \, kPa) = 21.22 \, kPa$$