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 kPa P_2 = 25.0 kPa
V_1 = 30.0 L V_2 = ?

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

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

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 L V_2 = ?
T_1 = 24°C (or 297K) T_2 = 58°C (or 331K)

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

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

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 25ºC. If this can is thrown in a fire, what is the pressure of the gas when its temperature reaches 928ºC ?

P_1= 103 kPa P_2= ?
T_1=25ºC (298K) T_2= 928ºC (1201K)

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

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

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

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

Avogadro's law

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

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 28ºC. How many moles of $N_2(g)$ does the cylinder contain?

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

$$PV = nRT \;\; \therefore \;\; n = \frac{P\,x\,V}{R\,x\,T} = \frac{2.00\,x\,10^4\ kPa\ x\ 2.0x10^1 L}{\frac{8.31\,L-kPa}{K-mol}\ x\ 301\,K} = 1.60\,x\,10^2\,mol$$

Dalton's Law of Partial Pressures- Total pressure in a mixture of gases is equal to the sum of the partial pressures 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?

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

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

Motion of Gas Particles

Effusion- process of a confined gas escaping through a tiny hole in its container.

Graham's Law- the rate of effusion or diffusion of a gas is inversely proportional to the square root of the molar mass of the gas.

$$KE = \frac{1}{2}mv^2 \qquad \therefore \qquad \frac{1}{2}m_a v_a^{\,2} = \frac{1}{2}m_b v_b^{\,2}$$

to solve for V

$$\frac{v_a^{\,2}}{v_b^{\,2}} = \frac{m_b}{m_a} \qquad \therefore \qquad \frac{v_a}{v_b} = \sqrt{\frac{m_b}{m_a}}$$

Example 6: Calculate the ratio of diffusion rates for ammonia gas (NH_3) and hydrogen chloride (HCl) at the same temperature and pressure.

molar mass of NH_3 = 17.04 g/mol molar mass of HCl = 36.46 g/mol

velocity ratio is $$\frac{V_{NH_3}}{V_{HCl}} = \sqrt{\frac{m_{HCl}}{m_{NH_3}}}$$

$$\frac{V_{NH_3}}{V_{HCl}} = \sqrt{\frac{36.46\,\frac{g}{mol}}{17.04\,\frac{g}{mol}}} = 1.46$$

The rate of diffusion of ammonia is 1.46 times faster than the rate of diffusion of hydrogen chloride.