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*Combination of gas laws:

By combining the laws stated by Boyle, Charles and Avogadro, we get the Ide gas equation which is known as the combination of gas laws.

If, for a gass of fixed mass, volume be "V", pressure be "P", ~~and~~ temperature be "T" and mole numb be "n". Then according to —

Boyle's law: $\boxed{V \propto \frac{1}{T}}$ ——— (i)

[n & T are constant]

Charles' law: $\boxed{V \propto T}$ ——— (ii)

[n & P are constant]

Avogadro's law: $\boxed{V \propto n}$ ——— (iii)

[P & T are constant]

Thus, $V \propto \frac{nT}{P}$ [when n, T, P are not - constant]

$\therefore V = k\frac{nT}{P}$ [k is proportionality constant]

For 1 mole gas, we represent 'k' with R. "R" is called molar gas constant. Replacing 'k' with 'R':

$$V = R\frac{nT}{P}$$

$$\therefore PV = nRT$$

Using this equation we can determine the mass, molecular mass, density and mole no. of ~~an~~ any ideal gas

* The volume of the gass particles are negligible compared to the volume of the container but in reality the concept of ideal gases are not justified.

Explanation:

We know, the volume of ideal gas particles are negligible. Compared to the total volume of the gas that can be neglected. But at high and low temperature this is not correct. At low and high temperatures gas maybe liquid or even solid. They have a volume which can not be neglected. So if we subtact the volume of the particles from the volume of ideal gas we get volume of real gas.

* In between the particles of ideal gas, the attraction force is considered absent but real gas has inter molecular attraction

Explanation:

~~Accor~~ The concept that ideal gases have no inter molecular attraction force is not justified in reality. If pressure is applied then the temperature of every molecule is reduced and the gas is converted to liquid. So, in real gas particles inter molecular attraction exists.

$$\text{Pressure of real gas} = \left(P + \frac{n^2 a}{V^2}\right)$$

Let in a container of 'V' volume Three gases A, B, C are present of which mole no. are n_A, n_B, & n_C and partial pressure P_A, P_B, P_C respectively. Total pressure of the gas, $= P$

$$\therefore P = P_A + P_B + P_C \quad \text{———} \quad (i)$$

$$n = n_A + n_B + n_C \quad \text{———} \quad (ii)$$

For "A" gas according to ideal gas equation,

$$P_A V = n_A RT$$

$$\text{or, } P_A = n_A \frac{RT}{V} \quad \text{———} \quad (iii)$$

Similarly ~~for~~ $P_B = n_B \frac{RT}{V}$ ——— (iv)

$$P_C = n_C \frac{RT}{V} \quad \text{———} \quad (v)$$

From (i)

$$P = P_A + P_B + P_C$$

$$= n_A \frac{RT}{V} + n_B \frac{RT}{V} + n_C \frac{RT}{V}$$

$$= \frac{RT}{V}(n_A + n_B + n_C)$$

$$= n\frac{RT}{V} \quad \text{———} \quad (vi)$$

(ii) ÷ (vi)

$$\frac{p_A}{p} = \frac{n_A RT \times V}{V \times nRT} = \frac{n_A}{n}$$

or, $p_A = \frac{n_A}{n} p$

Let, $\frac{n_A}{n} = x_A$

Partial pressure of A, $p_A = x_A p$

simillarly,

$p_B = x_B \times p$

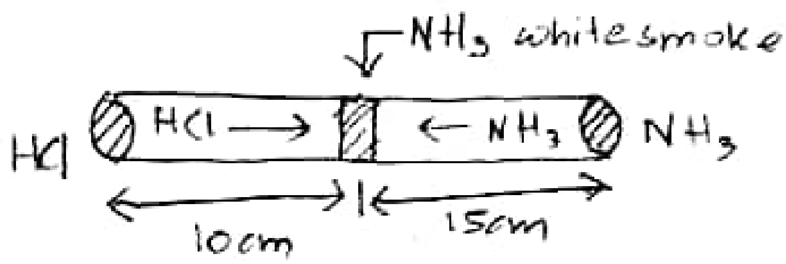
$p_c = x_c \times p$

(8) At constant pressure and temperature the rate of diffusion of any gas is inversely proportionaly to the square root of the density of the gas.

$\therefore$ rate of diffusion = r

density = d

$$r \propto \frac{1}{\sqrt{d}}$$



~~NH_3~~

Molecular mass of NH_3 = 17

" " " HCl = 36.5

G. vapour density of NH_3 = 8.5

" " " HCl = 18.25

$$\therefore \frac{r_1}{r_2} = \frac{15}{10} = 1.5 \text{ ——— (i)}$$

$$\therefore \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{18.25}{8.5}} = 1.5 \text{ ——— (ii)}$$

From (i) & (ii) $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$ ——— (iii)

And,

$$d_1 = \frac{M_1}{V} \qquad d_2 = \frac{M_2}{V}$$

[M = molecular mass.
V = molar volume]

From 3 (iii)

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

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

$$\therefore \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{d_2}{d_1}}$$

~~In~~ Molar mass is more important for Graham's gas law. Thus due to the difference of mass of gas molecules, rate of diffusion of different gases are ~~fit~~ different.