

States of Matter

SOL 1.

$$PV = m/MRT \Rightarrow P * m/V * RT/M = dRT/M$$

$$\therefore d = MP/RT$$

Substituting the value, we get

$$D = 17 * 5/0.082 * 303 = \mathbf{3.42 \text{ g/litre}}$$

SOL 2.

Given, moles = mass/mol.wt

Mass of gas = 3.7 g, mass of hydrogen = 0.184g

$$T_1 = 298\text{K}, T_2 = 17^\circ\text{C} = 273 + 17 = 290\text{K}$$

$$\text{Moles of } H_2 = n_1 = \text{Mass}/M. \text{ wt.} = 0.184/2 = 0.092$$

$$\text{Moles of gas} = n_2 = \text{Mass}/M. \text{ wt.} = 3.7/M$$

$$\text{For hydrogen } P_1V_1 = n_1RT_1 \quad \dots\dots\dots(i)$$

$$\text{For gas } P_2V_2 = n_2RT_2 \quad \dots\dots\dots(ii)$$

(\because Pressure and volume of gas are same)

\therefore From equation (i) and equation (ii)

$$P_1V_1/P_2V_2 = n_1RT_1/n_2RT_2 \text{ or } 1 = 0.092 * 298/n_2 * 290$$

$$\text{or } n_2 = 0.092 * 298/290 \text{ or } 3.7/M = 0.092 * 298/290$$

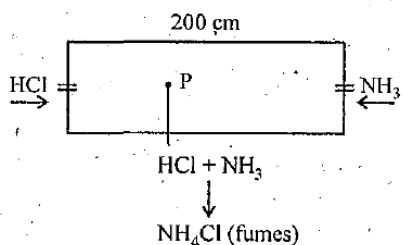
$$\text{or } 3.7/M = 0.0945$$

$$\therefore M = 3.7/0.0945 = \mathbf{39.15}$$

SOL 3.

Let NH_3 diffuse through = x cm

HCl diffuses through = y cm



According to Graham's law of diffusion

$$x/y = \sqrt{\text{Mol. wt HCl}/\text{Mol. wt of } NH_3} = \sqrt{36.5/17} = 1.465$$

$$x = 1.465 y \quad \dots(1)$$

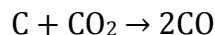
$$x + y = 200 \text{ cm} \quad \dots(2)$$

From these equations; $y = 85.2$ cm

Distance between P and X = $y = \mathbf{85.2 \text{ cm}}$.

SOL 4.

Following reaction takes place in tube



Volume of mixture of CO and CO₂ = 1L

Let volume of CO₂ in mixture = x

∴ Volume of CO in mixture = 2x

∴ original volume of CO in mixture = 1 - x

Total volume of CO after reaction = (1 - x) + 2x = 1 + x

1 + x = 1.6 (∵ It is given total volume after reaction = 1.6L)

∴ x = 0.6 L

∴ volume of CO₂ = 0.6 L

Volume of CO = 0.4 L

CO₂ : CO = 3 : 2

SOL 5.

NOTE THIS STEP: Since the evacuated bulb contains a mixture of oxygen and another gas in the molar ratio of 1 : 1 at a total pressure of 4000 mm, the partial pressure of each gas is 2000 mm.

The drop in the pressure of oxygen after 74 minutes

$$= 500/47 = 787.2 \text{ mm of Hg}$$

∴ After 74 minutes, the pressure of oxygen

$$= 2000 - 787.2 = 1212.8 \text{ mm of Hg}$$

Let the rate of diffusion of other gas be r_n , then

$$\frac{r_n}{r_{O_2}} = \sqrt{\frac{32}{79}}$$

$$\begin{aligned} \therefore \text{Drop in pressure for the other gas} &= 787.2 * 32/79 \\ &= 501.01 \text{ mm of Hg} \end{aligned}$$

∴ pressure of the other gas after 74 minutes

$$= 2000 - 501.01 \text{ mm} = 1498.99 \text{ mm of Hg}$$

Molar ratio = Moles of unknown gas/Moles of O₂

$$= 1498.99/1212.8 = 1.24/1 = \mathbf{1.24 : 1.}$$

[Partial pressure ∝ mole fraction]

SOL 6.

Since the pressures of gases are different, and the temperature is constant, the rate at which molecules of the two gases diffuse is directly proportional to the pressure. This rate of diffusion is also directly proportional to the distance travelled by the gas. Hence

$$r_1 \text{ (of HCl gas) at pressure } P = 60 = kP/\sqrt{36.5} \quad \dots\dots(i)$$

$$\text{and } r_2 \text{ (of NH}_3\text{) at 1 atm. Pressure } P = 40 = k * 1/\sqrt{17} \quad \dots(ii)$$

From (i) and (ii)

$$r_1/r_2 = 60/40 = kP/\sqrt{36.5} * \sqrt{17}/k * 1, P = 60/40 * \sqrt{36.5}/\sqrt{17} = \mathbf{2.197 \text{ atm}}$$

SOL 7.

TIPS/Formulae:

Total kinetic energy = $n (3/2 RT)$

where n = Number of moles of the gas

R = Gas constant

T = Absolute temperature

Molecular weight of methane,

$$\text{CH}_4 = 12 + 4 * 1 = 16$$

∴ Number of moles of methane in 8.0 gm of methane

$$= 8.0/16.0 = 0.5$$

$R = 8.314 \text{ joules/K/mole}$

$$T = 27 + 273 = 300 \text{ K}$$

∴ Total kinetic energy of the molecules in 8.0 gm of methane at 27°C = $n * 3/2 RT$

$$= 0.5 * 3/2 * 8.314 * 300 = 1870.65 \text{ joules}$$

∴ Average kinetic energy = $1870.65 / 6.023 * 10^{23} * 0.5$

$$= 6.21 * 10^{-21} \text{ joules/molecule}$$

SOL 8.

NOTE THIS STEP : First we should calculate the number of moles of the gas under the given condition by the relation $PV = nRT$

Here $P = 7.6 * 10^{-10} \text{ mm Hg}$

$$= 7.6 * 10^{-10} / 760 \text{ atm.} = 1 * 10^{-12} \text{ atm.}$$

$V = 1 \text{ litre}$

$$T = 273 + 0 = 273 \text{ K}$$

$R = 0.082 \text{ litre atm./K/mol}$

Putting the values in equation

$$n = PV/RT = 1 * 10^{-12} * 1 / 0.082 * 273 \text{ moles}$$

now since 1 mole = $6.023 * 10^{23}$ molecules

$$10^{-12} / 0.082 * 273 \text{ moles} = 6.023 * 10^{23} * 10^{-12} / 0.082 * 273 \text{ molecules}$$

$$= 2.7 * 10^{10} \text{ molecules}$$

SOL 9.

From ideal gas equation,

$$PV = nRT$$

$$PV = (m/M) RT \text{ or } M = m RT / PV$$

Let the molecular wt. of A and B be M_A and M_B respectively.

$$\text{Then } M_A = 2 RT / 1 * V ; M_B = 3 * RT / 0.5 * V$$

$$\therefore M_A / M_B = 2RT / V * 0.5V / 3 RT = 2 * 0.5 / 3 = 1/3$$

Therefore, the ratio $M_A : M_B = 1 : 3$

SOL 10.

If temperature and pressure are the same. [From Avogadro's law]

SOL 11.

$$C\sqrt{3PV/M}, P_1V_1/T_1 = P_2V_2/T_2$$

Where, $P_1 = 82 \text{ cm} = 82 * 13.61 * 981 \text{ dynes}$

$V_1 =$ Volume of gas at 82 cm and 20°C

P_2, V_2 and T_2 are the conditions at NTP,

$$\Rightarrow V_1 = P_2 V_2 / T_2 * T_1 / P_1 = 76.0 * 22400 / 273 * 293 / 82 = 22282 \text{ cc}$$

$M =$ molecular weight of ozone (O_3) = 48

$$\therefore C = \sqrt{3 * (82 * 13.6 * 981.) * 22282 / 48}$$

$$= 3.9 * 10^4 \text{ cm sec}^{-1}$$

ALTERNATESOLUTION

$$U_{\text{rms}} = \sqrt{3RT/M}$$

Given $T = 20^\circ\text{C} = 20 + 273 = 293\text{K}$

$R = 8.314 * 10^7 \text{ erg per degree per mol}$

$M(\text{of } O_3) = 48$

$$\therefore U_{\text{rms}} = \sqrt{3 * 8.314 * 10^7 * 293 / 48} = 3.9 * 10^4 \text{ cm sec}^{-1}$$

SOL 12.**TIPS/Formulae:**

No. of balloons that can be filled = V of H_2 available / V of one balloon

Calculation of total volume of hydrogen in the cylinder at N.T.P

$$P_1V_1/T_1 = PV_2/T_2$$

$$P_1 = 1 \text{ atm} \quad P_2 = 20 \text{ atm}$$

$$V_1 = ? \quad V_2 = 2.82 \text{ l}$$

$$T_1 = 23 \text{ K} \quad T_2 = 273 + 27 = 300 \text{ K}$$

$$\therefore V_1 = 20 * 2.82 * 273 / 300 * 1 = 51.324 \text{ l} = 51324 \text{ ml}$$

Actual volume to be transferred into balloons

$$= 51324 - 2820 \text{ ml} = 48504 \text{ ml}$$

[\therefore 2820 ml of H_2 will remain in cylinder]

$$\text{No. of balloons that can be filled up} = 48504 / 4851 = 9.999 = 10$$

$$\text{Volume of one balloon} = 4/3 \pi r^3 = 4/3 * 22/7 * (21/2)^3$$

[$\therefore r = \text{diameter}/2$]

$$= 4851 \text{ ml} = 4.851\text{L}$$

ALTERNATESOLUTION

Volume of balloon = 4.851 L (as calculate above)

Let no. of balloon to be filled n

$$\therefore \text{Total volume occupied by } n \text{ balloons} = 4.851 * n$$

Volume of H_2 present in cylinder = 2.82 L (given)

$$\therefore \text{Total volume of } H_2 \text{ at NTP} = (4.851n + .82)\text{L}$$

$$\begin{aligned}
 P_1 &= 1 \text{ atm} & P_2 &= 20 \text{ atm} \\
 V_1 &= 4.85 * n + 2.82 \text{ L} & V_2 &= 2.82 \text{ L} \\
 T_1 &= 273 \text{ K} \\
 P_1 V_1 / T_1 &= P_2 V_2 / T_2 \\
 \text{or } 1 * (4.85 * n + 2.82) / 273 &= 20 * 2.82 / 300 \\
 \therefore n &= 48.504 / 4.851 = 10
 \end{aligned}$$

SOL 13.

TIPS/Formulae:

$$\text{Average velocity} = \sqrt{8RT/\pi M}$$

$$\text{and Most probable velocity} = \sqrt{2RT/M}$$

Given -For CO₂

Average velocity at T₁ Most probable velocity at T₂

$$= 9 * 10^4 \text{ cm/sec} = 9 * 10^4 / 100 \text{ m/sec.}$$

$$= 9 * 10^2 = \text{m/sec.}$$

$$\therefore 9 * 10^2 = \sqrt{8 * 8.314 * T_1 / 3.14 * 44 * 10^{-3}} \dots (A)$$

[Average velocity at T₁K]

$$\text{and } 9 * 10^2 = \sqrt{2 * 8.314 * T_2 / 44 * 10^{-3}} \dots (B)$$

[Most probable velocity at T₂]

On solving, T₁ = **1682.5 K**, T₂ = **2143.4 K**

SOL 14.

Applying the general gas equation

$$PV = nRT = m/M RT$$

Here, Mol. wt. of acetylene i.e., C₂H₂ (M) = 26, P = 740/760 atm,

$$T = 50^\circ\text{C} = 50 + 273 = 323 \text{ K}$$

$$\therefore V = mRT/MP \text{ or } V = 5 * 0.082 * 323 * 760 / 26 * 740 = \mathbf{5.23 \text{ l}}$$

SOL 15.

Using gas equation; $PV = nRT$

Total no. of moles of gases in the mixture (n)

$$= PV/RT = 6 * 3 / 0.0821 * 300 = 0.7308 \text{ mol.}$$

Thus no. of moles of unknown gas = 0.7308 - 0.7

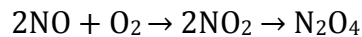
$$= 0.0308 \text{ mol.}$$

Now we know that

$$r_1/r_2 = \text{moles of hydrogen gas} / \text{moles of unknown gas} = 0.7 / 0.0308$$

Also we know that $r_1/r_2 = \sqrt{M_2/M_1}$

$$\therefore M_2 = (r_1/r_2)^2 M_1 \text{ or } M_2 = (0.7/0.0308)^2 * 2 = \mathbf{1033}$$

SOL 16.

Calculating the number of moles of NO and O₂ by applying the formula, $n = PV/RT$

Moles of NO in the larger flask = $1.053 * 0.250 / 0.082 * 300 = 0.0107$

[250 mL = 0.250 L]

Moles of O₂ in the smaller flask = $0.789 * 0.0100 * 0.082 * 300 = 0.0032$

[100 mL = 0.100 L]

The reaction takes place as follows.

	2NO	+	O ₂	→	N ₂ O ₄
Moles before	0.0107		0.0032		0
Reaction					
Mole after	90.0107-		0		0.0032
reaction	2 * 0032)				

Hence moles of NO reacting completely with 0.0032 moles of O₂ = $2 * 0.0032 = 0.0064$

Moles of NO left = $0.0107 - 0.0064 = \mathbf{0.0043}$

NOTE : Oxygen will be completely changed into NO₂ which in turn is completely converted into N₂O₄ which solidifies at 262 K. Hence at 220 K, the dimer is in the solid state and only NO present in excess will remain in the gaseous state occupying volume equal to 350 ml.

Hence pressure (P) of No gas left

$$= nR/V = 0.0043 * 0.082 * 220 / 0.350 = \mathbf{0.221 \text{ atm}}$$

[Total volume = 0.250 + 0.100 = 0.350 L]

SOL 17.

$$V_{\text{rms}} = \sqrt{3RT/M} = \sqrt{3P/d} \quad \dots(1)$$

Calculation of density (d) of gas molecules.

No. of gas molecules = $2 * 10^{21}$

Since mass of $6.023 * 10^{23}$ molecules of nitrogen = 28 g

Mass of $2 * 10^{21}$ molecules of N₂

$$= 28 * 2 * 10^{21} / 6.023 * 10^{23} = 56 / 602.3 = 0.093 \text{ g}$$

$$\therefore d = \text{Mass/Volume} = 0.093 / 1 \text{ g/l} \quad [\text{volume of flask} = 1 \text{ L}]$$

$$= 0.093 * 10^{-3} \text{ kg} / 10^{-3} \text{ m}^3 = 0.093 \text{ kg/m}^3$$

Substituting the value of d and P in equation (1)

$$V_{\text{rms}} = \sqrt{3 * 7.57 * 10^3 / 0.093} = \mathbf{494.16 \text{ m/sec}} \quad [V_{\text{rms}} = \sqrt{3P/d}]$$

$$(V_{\text{rms}})^2 = 3RT/M = 3P/d = 3 * 7.57 * 10^3 / 0.093 \quad [\text{squaring}]$$

$$RT/M = 7.57 * 10^3 / 0.093$$

$$T = 7.57 * 10^3 / 0.093 * 28 * 10^{-3} / 8.314 = \mathbf{274.13 \text{ K}}$$

Most probable velocity (α)/Root mean sq. velocity (V_{rms}) = 0.82 [given]

$$\Rightarrow \alpha = 0.82 * 494.16 = \mathbf{405.2 \text{ m/sec}}$$

ALTERNATESOLUTION

Given $V = 1\text{L} = 10^{-3}\text{m}^3$, $P = 7.57 * 10^{-3} \text{ Nm}^{-2}$, $R = 8.314$],

$N = 2 * 10^{21} / 6.023 * 10^{23}$ moles

$$PV = nRT$$

or $T = PV/nR$

$$= 7.57 \times 10^{-3} \times 6.023 \times 10^{23} / 2 \times 10^{21} \times 8.31 = 274.13 \text{ K}$$

$$U_{\text{rms}} = \sqrt{3RT/M} = \sqrt{3 \times 8.314 \times 274.13 / 28 \times 10^{-3}} \text{ m/s} = 494.15 \text{ m/s}$$

(Given U)

$$U_{\text{mp}}/U_{\text{rms}} = 0.82 \text{ (given)}$$

$$\therefore U_{\text{mp}} = 0.82 \times U_{\text{rms}} = 0.82 \times 494.15 = 405.2 \text{ m/sec}$$

SOL 18.

Partial pressure = Mole fraction * Total pressure

$$\therefore p_{\text{He}} = x_{\text{He}} \times P = 4/5 \times 20 = 16 \text{ bar}$$

[mole fraction of He = 4/5]

$$\therefore p_{\text{CH}_4} = 20 - 16 = 4 \text{ bar}$$

Now applying the formula

$$\frac{r_{\text{He}}}{r_{\text{CH}_4}} = \frac{P_{\text{He}}^0}{P_{\text{CH}_4}^0} \sqrt{\frac{M_{\text{CH}_4}}{M_{\text{He}}}} = \frac{16}{4} \sqrt{\frac{16}{4}} = \frac{16}{4} \sqrt{4}$$

$$\therefore r_{\text{He}} : r_{\text{CH}_4} = 8 : 1$$

\therefore Composition of the mixture (He : CH₄) effusing out = 8 : 1

SOL 19.

Calculation of volume of gas :

Weight of cylinder with gas = 29.0 kg

Weight of empty cylinder = 14.8 kg

\therefore Weight of gas in the cylinder = 14.2 kg

Pressure in cylinder = 2.5 atm

\therefore No. of moles (n) in 14.2 kg (14.2×10^3 g) of butane

$$N = \text{Wt. of butane} / \text{Mol. wt. of butane} = 14.2 \times 10^3 / 58 = 244.83 \text{ mol}$$

Applying gas equation,

$$V = nRT/P = 244.83 \times 0.0821 \times 300 / 2.5 = 2412 \text{ litres}$$

[27°C = 273 + 27 = 300]

Calculation of pressure in cylinder after use.

Weight of cylinder after use = 23.2 kg

Weight of empty cylinder = 14.8 kg

\therefore Wt. of unused gas = 8.4 kg = $8.4 \times 10^3 / 58$ moles of butane

$$\text{Thus } P = nRT/V = 8.4 \times 10^3 \times 0.0821 \times 300 / 58 \times 2412 = \mathbf{1.478 \text{ atm}}$$

[$V = 2412 \text{ L}$]

Calculation of volume of used gas at 2.5 atm and 27°C.

Weight of used gas = 14.2 - 8.4 = 5.8 kg

Pressure under normal usage conditions = 1 atm

$$V = nRT/P = 5.8 \times 10^3 / 58 \times 0.0821 \times 300 / 1 \quad [\because n = 5.8/58]$$

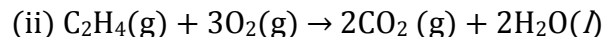
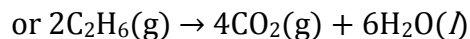
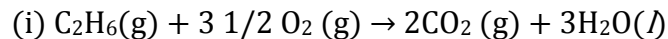
$$= 246 \text{ litres} = \mathbf{2.463 \text{ m}^3}$$

SOL 20.

Let the volume of ethane in mixture = x litre

∴ Volume of ethane = (40 - x) litre

Combustion reactions of ethane and ethane are :



Volume of O₂ required for complete combustion of ethane

$$= 7x/2 \text{ [For } x \text{ litres]}$$

Volume of O₂ required for complete combustion of ethane

$$= (40 - x) * 3 \text{ [For } (40 - x)L]$$

$$\therefore \text{ Total volume of } O_2 \text{ required} = 7x/2 + (40 - x)3 \text{ l}$$

Calculation of number of moles (n),

$$P = 1 \text{ atm, } V = 7x/2 + (40 - x)3 \text{ l; } R = 0.082 \text{ l atm K}^{-1} \text{ mol}^{-1};$$

$$T = 400 \text{ K}$$

$$\text{Since } n = PV/RT = 1 * [7x/2 + (40 - x)3] / 0.082 * 400 = 7x + (40 - x)6 / 2 * 0.082 * 400$$

$$\text{Mass of } n \text{ moles of } O_2 [7x + (40 - x)6 / 2 * 0.082 * 400] * 32 = 130$$

$$\text{or } 130 = [7x + 6 / 2 * 240 - 6x / 65.6] * 32$$

$$\Rightarrow 8528 = 32x + 240 * 32 \Rightarrow 32x = 848 \Rightarrow \text{or } x = 848 / 32 = 26.5$$

$$\text{Hence mole fraction (\%)} \text{ of ethane} = 26.5 / 40 * 100 = \mathbf{66.25\%}$$

$$\text{Mole fraction (\%)} \text{ of ethane} = \mathbf{33.75\%}$$

SOL 21.

Mixture

Krypton

$$r_{\text{mix}} = 1.16$$

$$r_{\text{Kr}} = 1$$

$$M_{\text{mix}} = ?$$

$$M_{\text{Kr}} = 84$$

We know that

$$r_{\text{mix}} / r_{\text{Kr}} = \sqrt{M_{\text{Kr}} / M_{\text{min}}} \text{ or } 1.16 / 1 = \sqrt{84 / M_{\text{mix}}}$$

$$\text{or } (1.16)^2 * 84 / M_{\text{mix}} \Rightarrow M_{\text{mix}} = 84 / (1.16)^2 = 62.426$$

Determination of the composition of the equilibrium mixture/Let the fraction of Cl₂

molecules dissociated at equilibrium = x

	Cl ₂	⇌	2Cl	Total
Initially	1		0	1
At equilibrium	1 - x		2x	1 - x + 2x = 1 + x

$$\therefore \text{ Total moles at equilibrium} = 1 - x + 2x = 1 + x$$

$$\therefore \text{ Normal molecular mass / Experimental molecular mass} = 1 + x$$

$$\therefore 71 / 64.426 = 1 + x$$

$$\therefore x = 0.137 = \mathbf{13.7\%}$$

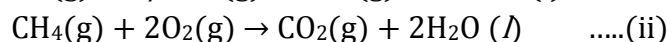
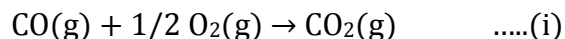
SOL 22.

TIPS/Formulae:

(i) He does not react with oxygen.

(ii) KOH absorbs only CO₂.

NOTE : When the mixture of CO, CH₄ and He gases (20 ml) are exploded by an electric discharge with excess of O₂, He gas remain as such and the other reactions involved are :



Let the volumes of CO and CH₄ to be 'a' ml and 'b' ml in the mixture then

Volume of He gas = [20 - (a + b)] ml

For the initial contraction of 13 ml

For the initial contraction of 13 ml,

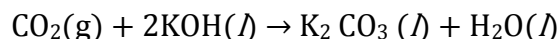
Volume of left hand side in the above reactions - 13 = volume
or right hand side.

$$\begin{aligned} \therefore [20 - (a + b)] + (a + 1/2 a) + (b + 2b) - 13 \\ = [20 - (a + 2b)] + a + b \text{ [neglect the volume of H}_2\text{O (l)}] \end{aligned}$$

(Since for gases, volume \propto no. of moles)

$$\therefore 1/2 a + 2b = 13 \text{ or } a + 4b = 26 \quad \dots\text{(iv)}$$

NOTE THIS STEP : The CO₂ produced above in reactions (ii) & (iii), (a + b) ml, reacts with KOH sol for a further contraction of 14 ml.



(a + b) ml

$$\therefore a + b = 14 \quad \dots\text{(v)}$$

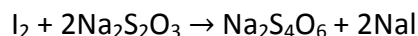
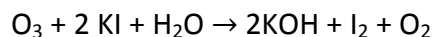
Solving (iv) & (v) we get, a = 10ml & b = 4 ml

$$\therefore \text{CH}_4 = 4/20 * 100 = 20\%, \text{CO} = 10/20 * 100 = 50\%$$

$$\& \text{He} = 100 - (20 + 50) = 30\%$$

SOL 23.

The concerned chemical reaction are :



Millimoles of ozone = Millimoles of I₂

$$\text{mM of O}_3 = \text{mM of I}_2 = 1/2 * \text{mM of Na}_2\text{S}_2\text{O}_3$$

$$= 1/2 * 40 * 1/10 = 2\text{mM} = 0.002 \text{ mole}$$

Calculation of total number of moles of O₂ and O₃

$$PV = nRT$$

$$1 * 1 = n * 0.0821 * 273 \text{ or } n = 1/0.0821 * 273$$

$$\text{Or } n = 0.044 \text{ mole}$$

$$\therefore \text{Moles of O}_2 = 0.044 - 0.002 = 0.042$$

$$\therefore \text{Wt. of O}_2 = \text{No. of moles} * \text{Mol. Wt.} = 0.042 * 32 = 1.344 \text{ g}$$

$$\text{Similarly, Wt. of O}_3 = 0.002 * 48 = 0.096 \text{ g}$$

$$\therefore \text{weight \% of } O_3 = 0.096/1.44 * 100 = \mathbf{6.6\%}$$

$$[\text{Total weight} = 1.344 + 0.096 = 1.44 \text{ g}]$$

$$\text{No. of photons or molecules of } O_3 = 0.096 * 6.023 * 10^{23} / 48 = \mathbf{1.2 * 10^{21}}$$

SOL 24.

Use vander Waal's eqn.

Real gas equation for one mole is given as:

$$(P + a/V^2)(V - b) = RT \text{ or } (P + a/V^2) = RT/V - b$$

$$\Rightarrow P = RT/V - b - a/V^2 = Rt/V(1 - b/V) - a/V^2$$

$$\Rightarrow PV = RT(1 - b/V)^{-1} - a/V$$

$$= RT(1 + b/V + b^2/V^2 + \dots) - a/V$$

(using binomial expansion)

$$\Rightarrow PV = RT(1 + b - a/RT / V + b^2/V^2 + b^3/V^3 + \dots) \quad \dots(i)$$

Give equation :

$$PV = RT(1 + B/V + \dots) \quad \dots(ii)$$

Comparing (1) and (2), we get, $B = \mathbf{b - a/RT / V}$

SOL 25.

$$\text{Weight of liquid} = 148 - 50 = 98 \text{ g}$$

Volume of liquid = $98/0.98 = 100 \text{ ml}$ = volume of vessel it means, vessel of 100 ml contain ideal gas at 760 mm Hg at 300 K

$$\text{Weight of gas} = 50.5 - 50 = 0.5 \text{ g}$$

Using, $PV = nRT = w/m RT$

$$760/760 * 100/1000 = 0.5/m * 0.082 * 300 [n = 0.5/m]$$

$$\therefore \text{Molecular weight of gas (m)} = \mathbf{123}$$

SOL 26.

	\rightleftharpoons		+	
PCl_5		PCl_3		Cl_2
Initial moles	1		0	0
Moles at eq	1 - 0.4	0.4		0.4

$$\therefore \text{Total moles at equilibrium} = 1 - 0.4 + 0.4 + 0.4 = 1.4$$

$$\text{Also Normal mol. Wt. of } PCl_5 / \text{Exp. Mol. Wt. of } PCl_5 = 1 + \alpha = 1.4$$

$$\text{Or } 208.5 / \text{Exp. mol. Wt. of } PCl_5 = 1.4$$

$$\therefore \text{Exp. mol. Wt. of } PCl_5 \text{ or m. wt. of mixture} = 208.5 / 1.4$$

Now using, $PV = w/m RT$ for mixture

$$d = w/V = Pm/RT = 1 * 208.5 / 1.4 * 0.082 * 400 = \mathbf{4.53 \text{ g/litre}}$$

SOL 27.

Van der waals equation for n moles of gas is

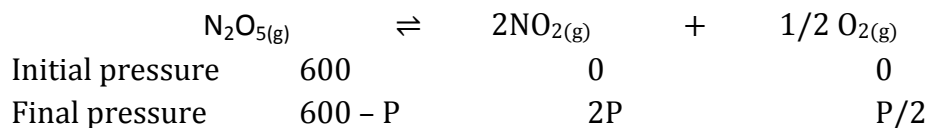
$$[P + n^2a/V^2] [V - nb] = nRT$$

Given $V = 4$ litre; $P = 11.0$ atm, $T = 300$ K;

$$b = 0.05 \text{ litre mol}^{-1}, n = 2;$$

$$\text{Thus, } [11 + 2^2a/4^2] [4 - 2 * 0.05] = 2 * 0.082 * 300$$

$$\therefore a = 6.46 \text{ atm litre}^2 \text{ mol}^{-2}$$

SOL 28.

$P \propto$ moles when V and T are constant

(where moles equivalent to pressure $-P$ are decomposed)

$$\text{Total pressure} = 600 - P + 2P + P/2 = 960 \text{ mm of Hg}$$

$$\therefore P = 240 \text{ mm Hg}$$

$$\text{Thus moles of } \text{N}_2\text{O}_5 \text{ decomposed} = 240/600 = 0.4$$

SOL 29.

We know that

$$r_1/r_2 = \sqrt{M_2/M_1} * P_1/P_2 \text{ or } n_1/t_1 * t_2/n_2 = \sqrt{M_2/M_1} * P_1/P_2$$

$$\text{or } 1/28 * 57/1 = \sqrt{M/28} * 0.8/1.6$$

$$\therefore M = 252$$

$$\left[\begin{array}{l} X_e + (F)_x = 252 \\ 131 + 19x = 252 \end{array} \right. \therefore x = 6$$

Thus compound of xenon with fluorine is **XeF₆**

SOL 30.

(I) Given $P = 1$ atm, $w = 12$ g; $T = (t + 273)$ K; $V = V$ litre

(II) If $T = t + 10 + 273 = t + 283$ K; $V = V$ litre,

$$P = 1 + 10/100 = 1.1 \text{ atm}$$

Using gas equation, $PV = w/m RT$

$$\text{Case I. } 1 * V = 12/m R (t + 273) \quad \dots(1)$$

$$\text{Case II. } 1.1 * V = 12/m R (t + 283) \quad \dots(2)$$

From (1) and (2), $t = -137^\circ\text{C}$ or $t = 100$ K

Also from (1), on substituting t and m (120), $V = 0.82$ litre

SOL 31.

van der Waal's equation for one mole of a gas is

$$[P + a/V^2] (V - b) = RT \quad \dots(1)$$

Given that volume occupied by CO_2 molecules, ' b ' = 0

Hence, (1) becomes $[P + a/V^2] V = RT$ or $P = RT/V - a/V^2$

Using $R = 0.082$, $T = 273\text{K}$, $V = 22.4\text{ l}$ for 1 mole of an ideal gas at 1 atm pressure.

$$\text{Case I. } 1 * V = 12/m R (t + 273) \quad \dots(1)$$

$$\text{Case I. } 1 * V = 12/m R (t + 273) \quad \dots(1)$$

$$\therefore P = 0.082 * 273 / 22.4 - 3.592 / (22.4)^4 = \mathbf{0.9922\text{ atm}}$$

SOL 32.

We know that, Compressibility factor, $Z = PV/RT$

$$0.5 = 100 * V / 0.082 * 273$$

$$\therefore V = 0.1117\text{L}$$

NOTE : Further when volume of a gas molecule is negligible, van der Waal's equation becomes

$$(P + a/V^2) (V - b) = RT$$

$$\text{Or } PV = RT - a/V \text{ or } a = RTV - PV^2$$

Substituting the values

$$A = (0.082 * 0.1119 * 273) - (100 * 0.1119 * 0.1119)$$

$$= \mathbf{1.253\text{ atm L}^2\text{ mol}^{-2}}$$

SOL 33.

$$(a) d = 0.36\text{ kg m}^{-3} = 0.36\text{ g/L}$$

(i) From Graham's Law of diffusion

$$\frac{r_v}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_v}}; 1.33 = \sqrt{\frac{32}{M_v}}$$

$$\therefore M_v = 32 / (1.33)^2 = \mathbf{18.09};$$

Where $M_v = \text{MW of the vapour}$

$$(ii) \text{ Thus, } 0.36\text{g} = 0.36 / 18.09\text{ mol}$$

0.36/18.09 mol occupies 1 L volume, so 1 mol occupies

$$18.09 / 0.36\text{ L} = 50.25\text{L}$$

Thus, molar volume of vapour = **50.25 L**

Assuming ideal behavior the volume of the vapour can be calculated by

$$V_1/T_1 = V_2/T_2 \Rightarrow 22.4 * 500 / 273 = \mathbf{41.025\text{L}}$$

(iii) Compressibility factor (Z)

$$= (PV)_{\text{obs}} / (PV)_{\text{ideal}} = 1 * 50.25 / 1 * 41.025 = \mathbf{1.224}$$

(iv) Z is greater than unity, hence it is the short range repulsive force that would dominate.

(\because actual density is less than given density)

$$(b) E = 3/2 KT = 3/2 * 8.31 / 6.02 * 10^{23} * 100$$

$$= \mathbf{2.07 * 10^{-20}\text{ J per molecule}}$$

(\because K, Boltzmann constant = R/N)

SOL 34.

$$C_{\text{rms}} = \sqrt{3} RT/M, C_{\text{av}} \sqrt{8} RT/\pi M$$

$$C_{\text{rms}}/C_{\text{av}} = \sqrt{3} RT/M * \sqrt{\pi M/8} RT = \sqrt{3\pi/8} = 1.085$$

$$C_{\text{rms}} = 1.085 * C_{\text{av}} = 1.085 * 400 = \mathbf{434 \text{ ms}^{-1}}$$

SOL 35.

The van der Waal equation (for one mole) of a real gas is

$$\left(P + \frac{a}{V_m^2}\right) (V_m - b) = RT$$

$$PV_m - Pb + a/V - ab/V^2 = RT$$

$$PV_m = RT + Pb - a/V_m + \frac{ab}{V_m^2} \quad \dots(i)$$

NOTE THIS STEP:

To calculate the intercept $P \rightarrow 0$, hence $V_m \rightarrow \infty$ due to which the last two terms on the right side of the equation (i) can be neglected.

$$\therefore PV_m = RT + Pb$$

When $P = 0$, intercept = **RT**