States of Matter

SOL 1.

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

∴$d = MP/RT$

Substituting the value, we get

D = 17 *5/0.082 *303 = **3.42 g/litre**

SOL 2.

Given, moles = mass/mol.wt

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

T_1 = 298K, T_2 = 17°C = 273 + 17 = 290K

Moles of H_2 = n_1 = Mass/M. wt. = 0.184/2 = 0.092

Moles of gas = n_2 = Mass/M. wt. = 3.7/M

For hydrogen $P_1V_1 = n_1RT_1$ ....(i)

For gas $P_2V_2 = n_2RT_2$ ....(ii)

(∵ Pressure and volume of gas │are same)

∴ From equation (i) and equation (ii)

$P_1V_1/P_1V_1 = n_1RT_1/n_2RT_2$ or 1 = 0.092 *298/n_2 *290

or n_2 = 0.092 *298/290 or 3.7/M = 0.092 *298/290

or 3.7/M = 0.0945

∴ M = 3.7/0.0945 = **39.15**

SOL 3.

Let NH_3 diffuse through = x cm

HCI diffuses through = y cm
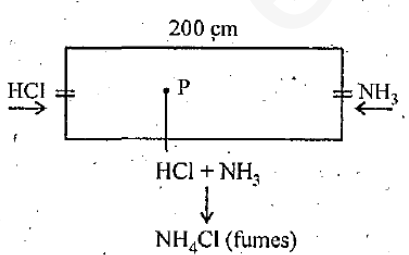
According to Graham's law of diffusion

x/y = √Mol. wt HCI/Mol. wt of NH_3 = √36.5/17 = 1.465

x = 1.465 y   ...(1)
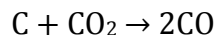
x + y = 200 cm   ...(2)

From these equations; y = 85.2 cm

Distance between P and X = y = **85.2 cm.**

SOL 4.

Following reaction takes places in tube

$C + CO_2 \rightarrow 2CO$

Volume of mixture of CO and CO_2 = 1L

Let volume of CO_2 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 CO_2 = 0.6 L

Volume of CO = 0.4 L

CO_2 : 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 mm of Hg

∴ After 74 minutes, the pressure of oxygen

= 2000 – 787.2 = 1212.8 mm of Hg

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

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

∴ Drop in pressure for the other gas = 787.2 * 32/79

= 501.01 mm of Hg

∴ pressure of the other gas after 74 minutes

= 2000 – 501.01 mm = 1498.99 mm of Hg

Molar ratio = Moles of unknown gas/Moles of O_2

= 1498.99/1212.8 = 1.24/1 = **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 (of HCI gas) at pressure $P = 60 = kP/\sqrt{36.5}$ ......(i)

and r_2 (of NH_3) at 1 atm. Pressure $P = 40 = k * 1/\sqrt{17}$ ....(ii)

From (i) and (ii)

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

SOL 7.

TIPS/Formulae:

Total kinetic energy $= n\,(3/2\,RT0$

where n = Number of moles of the gas

R = Gas constant

T = Absolute temperature

Molecular weight of methane,

$CH_4 = 12 + 4 * 1 = 16$

∴ Number of moles of methane in 8.0 gm of methane

$= 8.0/16.0 = 0.5$

$R = 8.314$ joules/K/mole

$T = 27 + 273 = 300$ K

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

$= 0.5 * 3/2 * 8.314 * 300 = 1870.65$ joules

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

$= 6.21 * 10^{-21}$ 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 $P = 7.6 * 10^{-10}$ mm Hg

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

$V = 1$ litre

$T = 273 + 0 = 273$K

$R = 0.082$ litre atm./K/mol

Putting the values in equation

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

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

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

$= 2.7 * 10^{10}$ molecules

SOL 9.

From ideal gas equation,

$PV = nRT$

$PV = 9m/M)\,RT$ or $M = m\,RT/PV$

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

Then $M_A = 2\,RT/1 * V$; $M_B = 3 * RT/0.5 * V$

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

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 cm = 82 * 13.61 * 981 dynes

V_1 = Volume of gas at 82 cm and 20°C

P_2, V_2 and T_2 are the conditions at NTP,

=> $V_1 = P_2 V_2/T_2 * T_1/P_1$ = 76.0 *22400/273 * 293/82 = 22282 cc

M = molecular weight of ozone (O_3) = 48

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

= 3.9 *10⁴ cm sec⁻¹

ALTERNATESOLUTION

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

Given T = 20°C = 20 + 273 = 293K

R = 8.314 * 10⁷ erg per degree per mol

M (of O_3) = 48

∴ $U_{rms} = \sqrt{3}$ * 8.314 * 10⁷ *293/48 = 3.9 * 10⁴ 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 atm P_2 = 20 atm

V_1 = ? V_2 = 2.82 l

T_1 = 23 K T_2 = 273 + 27 = 300 K

∴ V_1 = 20 * 2.82 * 273/300 * 1 = 51.324 l = 51324 ml

Actual volume to be transferred into balloons

= 51324 – 2820 ml = 48504 ml

[∵ 2820 ml of H_2 will remain in cylinder]

No. of balloons that can be filled up -= 48504/4851 = 9.999 = 10

Volume of one balloon = 4/3 πr³ = 4/3 * 22/7 * (21/2)³

[∵ r = diameter/2]

= 4851 ml = 4.851L

ALTERNATESOLUTION

Volume of balloon = 4.851 L (as calculate above)

Let no. of balloon to be filled n

∴ Total volume occupied by n balloons = 4.851 * n

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

∴ Total volume of H_2 at NTP = (4.851n + .82)L

$P_1 = 1$ atm $\qquad\qquad P_2 = 20$ atm

$V_1 = 4.85 * n + 2.82$ L $\qquad V_2 = 2.82$ L

$T_1 = 273$ K

$P_1 V_1 / T_1 = P_2 V_2 / T_2$

or $1 * (4.85 * n + 2.82)/273 = 20 * 2.82/300$

$\therefore n = 48.504/4.851 = $ **10**

SOL 13.

TIPS/Formulae:

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

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

Given –For CO_2

Average velocity at T_1 Most probable velocity at T_2

$= 9 * 10^4$cm/sec $= 9 * 10^4/100$ m/sec.

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

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

[Average velocity at T_1K]

and $9 * 10_2 = \sqrt{2 * 8.314 * T_2/44 * 10^{-3}}$ .....(B)

[Most probable velocity at T_2]

On solving, $T_1 = $ **1682.5 K**, $T_2 = $ **2143.4 K**

SOL 14.

Applying the general gas equation

$PV = nRT = m/M\ RT$

Here, Mol. wt. of acetylene i.e., C_2H_2 (M) $= 26$, P $= 740/760$ atm,

$T = 50^oC = 50 + 273 = 323$ K

$\therefore V = mRT/MP$ or $V5 * 0.082 *323 *760/26 *740 = $ **5.23 l**

SOL 15.

Using gas equation; $PV = nRT$

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

$= PV/RT = 6 *3/0.0821 *300 = 0.7308$ mol.
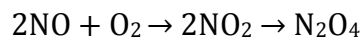
Thus no. of moles of unknown gas $= 0.7308 – 0.7$

$= 0.0308$ mol.

Now we know that

$r_1/r_2 = $ moles of hydrogen gas/moles of unknown gas $= 0.7/0.0308$

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

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

SOL 16.

$2NO + O_2 \rightarrow 2NO_2 \rightarrow N_2O_4$

Calculating the number of moles of NO and O_2 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 O_2 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	$\rightarrow$	N_2O_4
Moles before Reaction	0.0107		0.0032		0
Mole after reaction	90.0107- 2 * 0032)		0		0.0032

Hence moles of NO reacting completely with 0.0032 moles of O_2 = 2 * 0.0032 = 0.0064

Moles of NO left = 0.0107 – 0.0064 = **0.0043**

NOTE : Oxygen will be completely changed into NO_2 which in turn is completely converted into N_2O_4 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 = **0.221 atm**

[Total volume = 0.250 + 0.100 = 0.350 L]

SOL 17.

$V_{rms} = \sqrt{3RT/M} = \sqrt{3P/d}$ ...(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 N_2

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

$\therefore d$ = Mass/Volume = 0.093/1 g/1   [volume of flask = 1 L]

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

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

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

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

RT/M 7.57 * 10^3/0.093

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

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

$=> \propto = 0.82 * 494.16 =$ **405.2 m/sec**

ALTERNATESOLUTION

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

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

$PV = nRT$

or $T = PV/nR$
$= 7.57 *10^{-3} * 6.023 *10^{23}/2 *10^{21} * 8.31 = 274.13$ K
$U_{rms} = \sqrt{3RT/M} = \sqrt{3 * 8.314 * 274.13/28 *10^{-3}}$ m/s = 494.15 m/s
(Given U)
$U_{mp}/U_{rms} = 0.82$ (given)
$\therefore U_{mp} = 0.82 * U_{rms} = 0.82 * 494.15 = 405.2$ m/sec

SOL 18.

Partial pressure = Mole fraction * Total pressure
$\therefore p_{He} = x_{He} * P = 4/5 * 20 = 16$ bar
[mole fraction of He = 4/5]
$\therefore p_{CH_4} = 20 – 16 = 4$ bar

Now applying the formula

$$\frac{r_{He}}{r_{CH_4}} = \frac{P^0_{He}}{P^0_{CH_4}} \sqrt{\frac{M_{CH_4}}{M_{He}}} = \frac{16}{4} \sqrt{\frac{16}{4}} = \frac{16}{4} \sqrt{4}$$

$\therefore r_{He} : r_{CH_3} = 8 : 1$

$\therefore$ Composition of the mixture (He : CH$_4$) 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 = Wt. of butane/Mol. wt. of butane = 14.2 * 10^3/58 = 244.83 mol
Applying gas equation,
$V = nRT/P = 244.83 * 0.0821 * 300/2.5 = 2412$ 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 *10$_3$/58 moles of butane
Thus $P = nRT/V = 8.4 * 10^3 * 0.0821 * 300/58 * 2412 = $ **1.478 atm**
[$V = 2412$ L]
Calculation of volume of used gas at 2.5 atm and 27°C.
Weight f used gas = 14.2 – 8.4 = 5.8 kg
Pressure under normal usage conditions = 1 atm
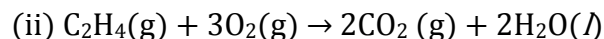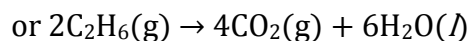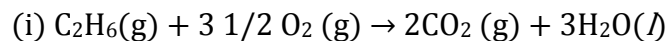$V = nRT/P = 5.8 *10^3/58 * 0.0821 *300/1$ [$\because n = 5.8/58$]
$= 246$ litres = **2.463 m³**

SOL 20.

Let the volume of ethane in mixture = x litre

∴ Volume of ethane = (40 - x) litre

Combustion reactions f ethane and ethane are :

(i) $C_2H_6(g) + 3\frac{1}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$

or $2C_2H_6(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

(ii) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$

Volume of O_2 required for complete combustion of ethane

= 7x/2 [For x litres]

Volume of O_2 required for complete combustion of ethane

= (40 - x) * 3 [For (40 - x)L]

∴ Total volume of O_2 required = 7x/2 + (40 - x)3 l

Calculation of number of moles (n),

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

T = 400 K

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

Mass of n moles of O_2 [7x +(40 -x)6 2 * 0.082 * 400] * 32 = 130

or 130 = [7x +6 240 – 6x/65.6] * 32

=> 8528 = 32x + 240 * 32 => 32x = 848 => or x = 848/32 = 26.5

Hence mole fraction (%) of ethane = 26.5/40 * 100 = **66.25%**

Mole fraction (%) of ethane = **33.75%**

SOL 21.

Mixture   Krypton

r_{mix} = 1.16 r_{Kr} = 1

M_{mix} = ? M_{Kr} = 84

We know that

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

or $(1.16)^2$ 84/M_{mix} => M_{mix} = 84/$(1.16)^2$ = 62.426

Determination of the composition of the equilibrium mixture/Let the fraction of CI_2 molecules dissociated at equilibrium = x

	CI_2	⇌	2CI	Total
Initially	1		0	1
At equilibrium	1 – x		2x	1 – x + 2x = 1 + x

∴ Total moles at equilibrium = 1 – x + 2x = 1 + 1

∵ Normal molecular mass/Experimental molecular mass = 1+x

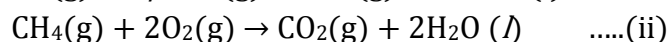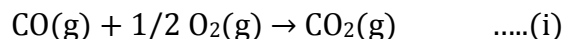∴ 71/64.426 = 1+ ∝

∴ ∝ = 0.137 = **13.7%**

SOL 22.

TIPS/Formulae:

(i) He does not react with oxygen.

(ii) KOH absorbs only CO_2.

NOTE : When the mixture of CO, CH_4 and He gases (20 ml) are exploded by an electric discharge with excess of O_2, He gas remain as such and the other reactions involved are :

$CO(g) + 1/2\ O_2(g) \rightarrow CO_2(g)$ .....(i)

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O\ (l)$ .....(ii)

Let the volumes of CO and CH_4 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

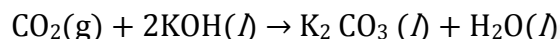For the initial contraction of 13 ml,

Volume of left hand side in the above reactions – 13 = volume

or right hand side.

$\therefore [20 - (a+b)] + (a + 1/2\ a) + (b + 2b) - 13$

$= [20 - (a + 2b)] + a + b$ [neglect the volume of $H_2O\ (l)$]

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

$\therefore 1/2\ a + 2b = 13$ or $a + 4b = 26$ ....(iv)

NOTE THIS STEP : The CO_2 produced above in reactions (ii) & (iii), $(a + b)$ ml, reacts with KOH sol for a further contraction of 14 ml.

$CO_2(g) + 2KOH(l) \rightarrow K_2CO_3\ (l) + H_2O(l)$

$(a + b)$ ml

$\therefore a + b = 14$ ....(v)

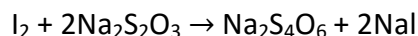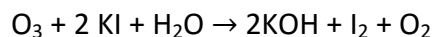Solving (iv) & (v) we get, $a = 10$ml & $b = 4$ ml

$\therefore CH_4 = 4/20 * 100 = 20\%$, $CO = 10/20 * 100 =$ **50%**

& He $= 100 - (20 + 50) =$ **30%**

SOL 23.

The concerned chemical reaction are :

$O_3 + 2\ KI + H_2O \rightarrow 2KOH + I_2 + O_2$

$I_2 + 2Na_2S_2O_3 \rightarrow Na_2S_4O_6 + 2NaI$

Millimoles of ozone = Millimoles of I_2

mM of O_3 = mM of I_2 = 1/2 * mM of $Na_2S_2O_3$

= 1/2 * 40 * 1/10 = 2mM = 0.002 mole

Calculation of total number of moles of O_2 and O_3

$PV = n R T$

$1 * 1 = n * 0.0821 * 273$ or $n = 1/0.0821 * 273$

Or $n = 0.044$ mole

$\therefore$ Moles of $O_2 = 0.044 - 0.002 = 0.042$

$\therefore$ Wt. of O_2 = No. of moles * Mol. Wt. = 0.042 * 32 = 1.344 g

Similarly, Wt. of O_3 = 0.002 * 48 = 0.096 g

∴ weight % of O_3 = 0.096/1.44 * 100 = **6.6%**

[Total weight = 1.344 + 0.096 = 1.44 g]

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

SOL 24.

Use vander Waal's eqn.

Real gas equation for one mole is given as:

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

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

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

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

(using binomial expansion)

$=> PV = RT (1 + b - a/RT /V + b^2/V^2 + b^3/V^3 + \ldots\ldots\ldots)$ .....(i)

Give equation :

$PV = RT (1 + B/V + \ldots\ldots.)$ ...(ii)

Comparing (1) and (2), we get, B = **b –a/RT /V**

SOL 25.

Weight of liquid = 148 – 50 = 98 g

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

Weight of gas = 50.5 - 50 = 0.5g

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

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

∴ Molecular weight of gas (m) = **123**

SOL 26.

	PCI_5	⇌	PCI_3	+	CI_2
Initial moles	1		0		0
Moles at eq	1 -0.4		0.4		0.4

∴ Total moles at equilibrium = 1 – 0.4 + 0.4 + 0.4 = 1.4

Also Normal mol. Wt. of PCI_5/Exp. Mol. Wt. of PCI_5 = 1 + ∝ = 1.4

Or 208.5/Exp. mol. Wt. of PCI_5 = 1.4

∴ Exp. mol. Wt. of PCI_5 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 = **4.53 g/litre**

SOL 27.

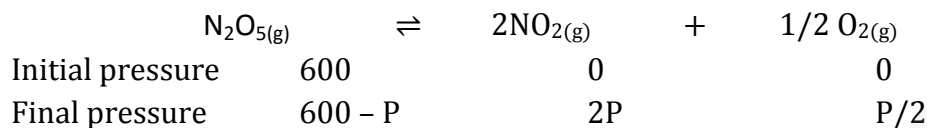Van der waals equation for n moles of gas is

$[P + n^2a/V^2] [V – nb] = nRT$

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

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

Thus, $[11 + 2^2a/4^2] [4 – 2 *0.05] = 2 * 0.082 *300$

∴ **a = 6.46 atm litre² mol⁻²**

SOL 28.

$$N_2O_{5(g)} \rightleftharpoons 2NO_{2(g)} + 1/2 \, O_{2(g)}$$

	$N_2O_{5(g)}$	$2NO_{2(g)}$	$1/2 \, O_{2(g)}$
Initial pressure	600	0	0
Final pressure	600 – P	2P	P/2

P ∝ moles when V and T are constant

(where moles equivalent to pressure –P are decomposed)

Total pressure = 600 – P + 2 P + P/2 = 960 mm of Hg

∴ P = 240 mm Hg

Thus moles of N_2O_5 decomposed = 240/600 = **0.4**

SOL 29.

We know that

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

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

∴ M = 252

$\begin{bmatrix} X_e + (F)_x = 252 \\ 131 + 19x = 252 \ ∴ x = 6 \end{bmatrix}$

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 atm

Using gas equation, PV = w/m RT

Case I. 1 * V = 12/m R (t + 273)   ….(1)

Case II. 1.1 * V = 12/m R (t + 283) ….(2)

From (1) and (2), t = - 137°C or t = **100 K**

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

SOL 31.

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

$[P + a/V^2] (V - b) = RT$ ….(1)

Give 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 = 273K, V = 22.4 l for 1 mole of an ideal gas at 1 atm pressure.

Case I. 1 * V = 12/m R (t + 273)   ….(1)

Case I. 1 * V = 12/m R (t + 273)   ….(1)

$\therefore$ P = 0.082 *273/22.4 – 3.592/(22.4)⁴ = **0.9922 atm**

SOL 32.

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

0.5 = 100 *V/0.082 *273

$\therefore$ V = 0.1117L

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

$(P + a/V^2) (V - 0) = RT$

Or PV = RT – a/V or a = RTV – PV²

Substituting the values

A = (0.082 * 0.1119 *273) – (100 * 0.1119 * 0.1119)

= **1.253 atm L² mol⁻²**

SOL 33.

(a) d = 0.36 kg m⁻³ = 0.36 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 = $ **18.09;**

Where M_v = MW of the vapour

(ii) Thus, 0.36g = 0.36/18.09 mol

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

18.09/0.36 L = 50.25L

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

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

$V_1/T_1 = V_2/T_2$ => 22.4 * 500/273 = **41.025L**

(iii) Compressibility factor (Z)

= (PV)$_{obs}$/(PV)$_{ideal}$ = 1 *50.25/1 *41.025 = **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 *10²³ *100

= **2.07 *10⁻²⁰ J per molecule**

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

SOL 34.

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

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

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

SOL 35.

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

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

$PV_m – Pb + a/V – ab/V^2 = RT$

$PV_m = RT + Pb – a/V_m + \dfrac{ab}{V_m^2}$ .....(i)

NOTE THIS STEP:

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

∴ $PV_m = RT + Pb$

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