



**34.** (P) (V - nb) = nRT $P = \frac{nRT}{V nh}$ **38.**  $C_x H_y + \left(x + \frac{y}{4}\right) O_2 \longrightarrow XCO_2 + \frac{y}{2} H_2O$ 7.5 ml 36 ml  $36-7.5\left(x+\frac{y}{4}\right)+7.5x=28.5$  $\Rightarrow$  36-7.5  $\left(15+\frac{y}{4}\right)+7.5 x=28.5$ y = 4; x = 2So formula =  $C_2 H_4$ **40.**  $C_4H_{10} + \frac{13}{2}O_2 \longrightarrow 4CO_2 + 5H_2O_2$ xml n-butane y ml isobutane Volume of  $O_2 = x \times \frac{13}{2} + y \times \frac{13}{2}$ 42.  $C_{v}H_{v} + O_{2} \longrightarrow CO_{2} + H_{2}O$  $_{X \times} n_{Cx} H_{v} = n_{CO_2}$ (POAC on C)  $x \times 500 = 2500$ (x=5) $y \times nCx H_y = 2 \times n_{H_2O}$  (POAC on H)  $y \times 500 = 2 \times 3000$ y = 12 $Formula = C_5 H_{12}$ 43.  $C_x H_y + O_2 \longrightarrow X_{CO_2} + \frac{y}{2} + H_2O$  $15 \text{ ml} \quad \frac{357 \times 21}{100} \text{ ml}$ 75 ml  $\left(x + \frac{y}{4}\right) \times 15 = 75$   $x + \frac{y}{4} = \frac{75}{15}$  $x + \frac{y}{4} = 5 \qquad \qquad x + \frac{y}{4} = 5$  $3 + \frac{y}{4} = 5$ 15x + 15x + 282 = 327v = 8x = 3 $Formula = C_3 H_8$ 

44. Two flask initially at 27° and 0.5 atm, have same volume and 0.7 mole thus each flask has 0.35 mole Let n mole of gas are diffuse from II to I on heating the flask at 127°C Mole in I flask = 0.35 + n, Mole in II flask = 0.35 - nIf new pressure of flask is P then for I flask  $P \times V = (0.35 + n) \times R \times 300$ ; for II flask  $P \times V$  $=(0.35-n) \times R \times 400$ n = 0.5mole in I flask = 0.40mole in II flask = 0.30 $0.5 \times 2V = 0.7 \times 0.0821 \times 300$  (initially) V=17.24 Lt.  $P \times 17.24 = 0.30 \times 0.0821 \times 400$  (finally) P = 0.57 atm 46.  $H_2 + \frac{1}{2}O_2 \rightarrow H_2O_{(1)}$ a b 0 a–2b Reaction is studied at constant P &T. a+b=40a - 2b = 10b = 10mla = 30 mlmole fraction of  $H_2$  = volume fraction of  $H_2$  = 30/40 = 0.75. **48**. N<sub>2</sub> 2N at t = 0  $\frac{1.4}{28} = \frac{1}{20}$ 0 at  $t = t_f$   $\frac{1}{20} - x$ 2x but, x = 30% of  $\frac{1}{20} = \frac{3}{200}$ Final number of mole =  $\frac{1}{20} - x + 2x = \frac{1}{20} + x = \frac{1}{20} + \frac{3}{200}$  $=\frac{13}{200}$  $\therefore P = \frac{13}{200} \times \frac{0.0821 \times 1800}{5} = 1.92 \text{ atm.}$ **50.**  $\frac{t_{mix}}{t_{O_2}} = \frac{r_{O_2}}{r_{mix}} = \sqrt{\frac{M_{mix}}{32}}$  $\frac{234}{224} = \sqrt{\frac{M_{mix}}{32}}$  $M_{mix} = 34.92$  $\Rightarrow \frac{1}{\sqrt{M_{mix}}} = \frac{X_{gas}}{\sqrt{M_{gas}}} + \frac{X_{O_2}}{\sqrt{M_O}}$  $\Rightarrow \frac{1}{\sqrt{34.92}} = \frac{0.2}{\sqrt{M_{\text{max}}}} + \frac{0.8}{\sqrt{32}}$  $M_{gas} = 51.5$ 





**53.** Let both gases meet at  $n^{th}$  row



= 2

$$\frac{r_{NO_2}}{r_{C_6H_{10}Br}} = \frac{x}{12 - x} = \sqrt{\frac{179}{44}}$$

$$x = 24 - 2x$$

$$3x = 24$$

$$x = 8 = n - 1$$

$$n = 9^{\text{th}} \text{ Row}$$

56. I – Slope of isotherm below critical point < 0.</li>
 Slope of isotherm above critical point < 0.</li>
 Slope of isotherm at critical point = 0.



So slope of isotherm at critical point is maximum.

$$II - T_{c} = \frac{8a}{27 \, \text{Rb}}$$

 $T_{c} \propto a$ 

Larger value of  $T_c$  It means less decreases in temperature is required to liquifly the gas. Gas will liquify at higher temperature. So, easier'll be liquificatzion.

III- When gas is below critical temperature. It is 'liquid' so vander waal equation of state is not valid.

So, Answer (B).

59. If Z > 1 positive deviation Z < 1 negative deviation



- (a) at T = 500 K, P = 40 atm corresponds to 'a' substance gas (b) at T = 300 K, P = 50 atm corresponds to 'b' substance - liquid
- (c) at T < 300 K, P > 20 atm corresponds to 'c' substance liquid
- (d) at T < 500 K, P > 50 atm corresponds to 'd' substance liquid

So, Answer (D)

65. 
$$\sqrt{\frac{8RT}{\pi M}} = 2 \sqrt{\frac{8 \times R \times 300}{\pi M}} \implies T = 1200 \text{ K} = 927^{\circ}\text{C}$$
  
66.  $n_T = n_A + n_B$   
 $P_f V_f = P_A V_A + P_B V_B$   
 $P_f = \frac{2150P}{250}$ 



Decrease in P of A = 20 P 
$$-\frac{215}{25}$$
 P  
% drop  $= \frac{285}{25}$  P  $\times \frac{1}{20} \times 100 = 57\%$ .  
67. PV  $\times$  T  
68. Using  $p_1V_1 = P_2V_2$  1 × 2.5 = 0.5 × P\_2 = 5 har.  
 $\therefore$  % increase in pressure =  $\frac{(5-1)bar}{1bar} \times 100\% = 400\%$ .  
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 $\therefore$  % increase in pressure =  $\frac{(5-1)bar}{1bar} \times 100\% = 400\%$ .  
 $\therefore$  where for  $N_{ma} = H_2 \times N_2 < 0.3$  HBr.  
70. Net pressure of gas = P<sub>gu</sub>.  
 $P_{ua} = 650 \text{ rm.}$   
 $\frac{P_1V_1}{V_1} = \left(\frac{P_2V_2}{T_2}\right)_{STP}$   
6. N<sub>1</sub> + 3H<sub>1</sub>  $\longrightarrow$  2NH<sub>3</sub>  
 $t = 0$  limele 4 mole 0  
 $t = t_{imac} 0$  limele 2 mole 0  
 $t = \frac{1 \times 0.22}{RT}$  (B) Total moles =  $\frac{1 \times 0.2}{RT}$   
1. Suppose the cylinder will burst at  $T_1K$   
 $T_2 = \frac{P_1T_1}{P_1} (V_1 = V_2) = \frac{149 \times 200}{12} = 372.5 K$   
3. Given  $\frac{r_A}{r_B} = \frac{16}{3}; \frac{w_A}{w_B} = \frac{2}{3}$   
 $we have  $\frac{r_A}{r_B} = \frac{n_B}{m_A} \sqrt{\frac{M_B}{M_A}}$   
 $= \frac{16}{3} = \frac{2}{3} \left(\frac{M_B}{M_A}\right)^{3/2} \Rightarrow \left(\frac{M_B}{M_A}\right)^{3/2} = 8 \Rightarrow \frac{M_B}{M_A} = 4$   
 $\frac{P}{3} \times 2V = nRT$   
 $\therefore$  mole ratio  $= \frac{8}{3}$   $T = \frac{2}{3}$  T$ 



 $\frac{283}{273}V_0$ 

11. 
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$
$$\frac{V_0}{273} = \frac{V_2}{283} \implies V_2 =$$

12. (B) Frequency of collision will increase.

13. 
$$n_{Total} = \frac{PV}{RT} = \frac{1 \times 2}{0.0821 \times 299} = 0.081$$
 moles  
 $X_{H_2} = \frac{n_{H_2}}{n_{total}} = \frac{0.0788}{\frac{0.0821}{2} \times 299} = 0.967$ 

14. K.E.  $\infty$  Temperature

16. 
$$P_{c} = \frac{a}{27b^{2}}$$
  $T_{c}^{2} = \frac{64 a^{2}}{27 \times 27 R^{2} b^{2}}$   
 $V_{c} = 3b$   $\frac{T_{c}^{2}}{P_{c}} = \frac{64 a^{2}}{27 \times 27 R^{2} b^{2}} \times \frac{27 b^{2}}{a}$   
 $T_{c} = \frac{8a}{27 Rb}$   $a = \frac{27 R^{2} T_{c}^{2}}{64 P_{c}}$ 

**19.**  $yM_x \longrightarrow xM_y$  $\frac{y}{x} = \frac{12}{8} = \frac{3}{2}$  $M_{2} \& M_{3}$ 

**20.** No. of moles of  $N_2 = \frac{7}{28} = \frac{1}{4}$ 

Total moles =  $\frac{1}{4}$  +1 +  $\frac{1}{8}$ No. of moles of  $H_2 = 1$  Mole

No. of moles of SO<sub>2</sub> =  $\frac{1}{8}$  moles =  $\frac{1}{8}$  (2 + 8 + 1) =  $\frac{11}{8}$ 

$$P = \frac{nRT}{V} = \frac{11}{8} \times \frac{0.0821 \times 300}{6} = 5.64 \approx 5.7 \text{ atm.}$$

21. Let Temp (T) where  $V_{rms}$  of  $N_2 = V_{rms}$  of  $C_3H_8$  at STP



22. Pressure of Total mixture = 10 atm  

$$P_{A} + P_{B} + P_{C} = 10$$

$$3 + 1 + P_{C} = 10 \implies P_{C} = 6 \text{ atm}$$
Total moles of mixture = 10  

$$n_{A} + n_{B} + n_{C} = 10$$

$$\frac{P_{A}}{P_{B}} = \frac{n_{A}}{n_{B}} = \frac{3}{1} \implies \frac{P_{B}}{P_{C}} = \frac{n_{B}}{n_{C}} = \frac{1}{6}$$
Let  $n_{A} = K \implies n_{B} = \frac{K}{3} \qquad n_{C} = \frac{1}{6}$   

$$n_{B} = 2K$$

$$\Rightarrow K + \frac{K}{3} + 2K = 10 \implies \frac{K}{3} = \frac{n_{C}}{6} \implies n_{C} = 2K$$

$$\Rightarrow K \left(\frac{10}{3}\right) = 10 \text{ K} = 3, \implies n_{A} = 3$$

$$n_{B} = 1$$

$$n_{C} = 6$$

$$\therefore \text{ weight of 'C' in mixture = 2 \times 6 = 12$$

10

23. 
$$C_n H_{3n} O_m + yO_2 \longrightarrow nCo_2(g) + \frac{3n}{2} H_2O(1)$$

Contraction in volume = Contraction in moles of gas =

$$1 + \frac{3n}{4} - \frac{m}{2}$$
  
$$\Rightarrow \left(2n + \frac{3n}{2} - m\right) \times \frac{1}{2} = \mathbf{y} \implies \mathbf{n} + \frac{3n}{4} - \frac{m}{2} = \mathbf{y}$$

24. 
$$\frac{P_{N2}}{P_{CO}} = \frac{X_{N2}}{X_{CO}} = \frac{n_{N2}}{n_{co}} = \frac{x \times 28}{28 \times x} = 1 \quad P_{N_2} = P_{co}$$
 (A)

Where  $x_{n_2}$ ,  $x_{co}$  is mole fraction of  $N_2$  & CO and x is wt. of N<sub>2</sub> & CO taken.

- **25.** Average K.E. =  $\frac{3}{2}$  RT and T is constant 298 K
  - : K.E. is same for all gases at same Temperature.



 $\mathbf{P}_{\mathbf{A}}\mathbf{V}_{\mathbf{A}} \!=\! \mathbf{P}_{\mathbf{A}}^{'}\mathbf{V}^{'}$ Where  $V' = V_A + V_B = (V_A + V) ml$ 



=

$$P_{A} 100 = \frac{2}{5} P_{A} \times V' \qquad P_{A} = \frac{P_{A} \times 40}{100} = \frac{2}{5} P_{A}$$

$$250 = V' \qquad \Rightarrow V_{A} + V = 250 \text{ ml}$$

$$V = 150 \text{ ml}$$

27. No. of molecules in 22.4 L at STP

is  $6.02 \times 10^{23} = 1$  mole of gas.

**28.** 
$$P = \frac{nRT}{V} = \frac{2 \times 0.0821 \times 546}{44.8} = 2 \text{ atm}$$

**31.**  $3P_C V_C^2 = 3 \times \frac{a}{27b^2} \times (3b)^2 = a$ 

- New gas may have pressure equal to sum of both reacting gases or less or more depending on the reacting gases and product formed.
- 5. From Charle's law, volume is theoretically zero at 0 K and kinetic energy is directly proportional to absolute temperature. So, it is zero at 0 K.
- Pressure increases due to increase in number of moles in fixed volume. This is because number of collision per unit volume increases.

8. Rate of diffusion 
$$\propto \frac{1}{\sqrt{Molar mas}}$$

9. On collision, K.E. redistributes, one molecule acquires and other looses the K.E. so, average K.E. remains the same at the same temperature.

## EXERCISE - 3 Part # I : Matrix Match Type

- 2. (A) PV = K (Boyle's law)
  - $P_1V_1 = P_2V_2 = P_3V_3$
  - (B) From charle's law

$$V \propto T \implies \frac{V}{T} = K \implies \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$r \propto \frac{1}{\sqrt{M}}$$
 and  $d = \frac{PM}{RT} \implies d \propto M$ .

So, 
$$r \propto \frac{1}{\sqrt{d}}$$
.

(D) From Dalton's law of partial pressure at constant temperature.

$$P = P_1 + P_2 + \dots$$

(E) Vander Waal's equation (real gas equation)

$$\left(P + \frac{a}{V^2}\right)(V-b) = RT$$
 (For 1 mole)

(F)  $\frac{R}{N} = K$  (Boltzmann constant)

- (G) Molar volume = 22.4 L at STP
- (I) Constant temperature P V curve is called isotherm.
   (J) Graph between V and T at constant pressure called isobar.
- 3. (A) At low pressure, b is negligible in comparison to  $V_m$ .

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

$$\Rightarrow \frac{PV_m}{RT} = Z = 1 - \frac{a}{V_m RT} < 1$$

So, gas is more compressible than ideal gas.

(B) At high pressure,  $\frac{a}{V_m^2}$  is negligible in comparison to P

$$P(V_m - b) = RT$$

$$\Rightarrow \frac{PV_m}{RT} = Z = 1 + \frac{Rb}{RT} < 1$$

So, gas is less compressible than ideal gas.

(C) Low density of gas means pressure is low so, at

low pressure  $Z = 1 - \frac{a}{V_m RT} < 1$  and gas is more compressible than ideal gas.

(D) At 0°C H<sub>2</sub> and He have  $a \approx 0$ .

So,  $Z = 1 + \frac{Pb}{RT}$  and gas is less compressible than

ideal gas.

#### **Part # II : Comprehension**

#### Comprehension #1:

1. At critical point

$$\frac{\partial \mathbf{p}}{\partial V_{\mathrm{m}}} = 0 \Longrightarrow - \frac{\mathbf{R}\mathbf{T}_{\mathrm{C}}}{\mathbf{V}_{\mathrm{m}}^{2}} + \frac{2\mathbf{B}}{\mathbf{V}_{\mathrm{m}}^{3}} - \frac{3\mathbf{C}}{\mathbf{V}_{\mathrm{m}}^{4}} = 0$$

$$\Rightarrow -RT_{c} + \frac{2D}{V_{m}} - \frac{3C}{V_{m}^{2}} = 0 \Rightarrow RT_{c}V_{m}^{2} - 2BV_{m} + 3C = 0$$

as equation will have repeated root then D = 0

$$\Rightarrow T_{c} = \frac{B^{2}}{3RC}$$





At 100 K and pressure below 20 atm it may have liquid or gaseous state depending on the pressure.

#### **Comprehension #2:**

- **2.**  $P_1 = 75 \text{ cm of Hg}, V_1 = 24 \times A$  $P_2 = 75 + 10 + \frac{20.4 \times 10}{13.6} = 100 \text{ cm of Hg}$  $\Rightarrow$  75 × 24 = 100 × x x = 18 cm**P**<sub>2</sub>
- $\frac{\text{Case II}}{P_2 = (P_0 h)}$ 3. Case I  $P_1 = (P_0 + h)$  $P_1V_1 = P_2V_2$

Now in both the cases, the gas is the same and temperature is also constant, hence boyles law can be applied.

#### **Comprehension #3:**

 $2. \quad Z = \frac{PV_m}{RT} > 1$ 

 $\frac{PV_m}{RT} = \frac{1 \times 22.4}{R \times T}$ 

At same pressure = 1 atm.

$$1 \times V_m$$
  $1 \times 22$ 

 $\frac{1 \times V_{m}}{RT} > \frac{1 \times 22.4}{R \times T}$  $V_m > 22.4 L$  at STP for real gas.  $V_m = 22.4 L$  of real gas, we have to increase the ⇒ For, pressure.

# **EXERCISE - 4** Subjective Type The number of moles of the gas = $\frac{\frac{750 - 100}{750} \times \frac{50}{1000}}{\frac{1}{12} \times (20 + 273)}$

=0.0018 mol. Hence volume at STP =  $22400 \times .0018 = 40.3 \text{ mL}$ 

Since the vessel is open, the gas can escape out on 3. heating. In this case we can imagine a imaginary boundary to trace the final volume of the gas. (a) Now in the given situation, 3/5th of the gas (by amount) has escaped which means only 2/5th of the amount is occupying the complete volume of the open container.

And the Pressure is constant,

$$\frac{\mathbf{n}_2}{\mathbf{n}_1} = \frac{\mathbf{T}_1}{\mathbf{T}_2} \implies \frac{2\mathbf{n}/5}{\mathbf{n}} = \frac{300}{\mathbf{T}_2} \implies \mathbf{T}_2 = 750 \,\mathrm{K}.$$

**(b)** Similarly,  $\frac{n_2}{n_1} = \frac{300}{900} \implies 1/3$ rd of the gas remained in

the container. Hence, 2/3rd of the gas escaped.

(c) Half of the gas will escape out by doubling the temperature. Hence, the final temperature = 600 K.

Let the total number of molecules of the gas be n, of which n, are in the larger sphere and n, in the smaller sphere after the stopcock is opened

$$n = n_1 + n_2$$
 and  $pV = nRT$ 

$$\frac{\mathbf{pV}}{\mathbf{RT}_{1}} = \frac{\mathbf{p'V}}{\mathbf{RT}_{1}} + \frac{\mathbf{p'V}}{2T_{2}\mathbf{R}}$$

$$\mathbf{p'} = \frac{2\mathbf{pT}_{2}}{2T_{2} + T_{1}}$$

$$\mathbf{V}$$

$$\mathbf{V}$$

$$\mathbf{V}$$

Since P is constant, From Charles law,  $\frac{100}{50} = \frac{50+273}{t+273}$ 

$$\Rightarrow$$
 t = -111.5°C

5.

Let total moles of gas mixture be 100. 9.

$$P_{N_2} = \left(\frac{n_{N_2}}{n_T}\right) \times P_T = \frac{55}{100} \times 760 = 418 \text{ torr.}$$
$$P_{O_2} = \left(\frac{n_{O_2}}{n_T}\right) \times P_T = \frac{25}{100} \times 760 = 190 \text{ torr.}$$
$$P_{CO_2} = (760 - 418 - 190) = 152 \text{ torr.}$$

- 10. Molar mass of methane  $(CH_{4}) = 12 + 4 \times 1 = 16 \text{g mol}^{-1}$ 
  - Mass of carbon dioxide (CO<sub>2</sub>) =  $12 + 2 \times 16 = 44$ g mol<sup>-1</sup>

Moles of 
$$CH_4 = \frac{3.2}{16} = 0.2$$
  
Moles of  $CO_2 = \frac{4.4}{44} = 0.1$   
 $pCH_4 = \frac{nRT}{V} = \frac{0.2 \times R \times T}{V} = \frac{0.2 \times 8.314 \times 300}{9 \times 10^{-3}}$ 



$$= 0.554 \times 10^5 \,\mathrm{Nm^{-2}} = 5.54 \,10^4 \,\mathrm{Pa}$$

$$\begin{split} pCO_2 &= \frac{nRT}{V} = \frac{0.1 \, mol \times 8.314 \, Jmol^{-1}K^{-1} \times 300K}{9 \times 10^{-3} \, m^3} \\ &= 2.77 \times 10^4 \, Pa \\ \text{Total pressure of mixture} &= pCH_4 + pCO_2 \\ &= (5.54 + 2.77)10^4 + Pa \\ &= 8.218 \times 10^4 \, Pa. \end{split}$$

13. Based on Dalton law of partial pressure, partial pressure of  $H_2$  is determined hence, its number of moles. Partial pressure of  $H_2$  and water vapour are known, hence, mole fraction can be determined. Since number of moles of dry  $H_2$  is known, mass of dry  $H_2$  can be determined.

(a) 
$$n_{H_2O}$$
 (vapour) =  $\frac{P_{H_2O}V}{RT} = \frac{\frac{19}{760} \times 0.300}{0.0821 \times 294} = 0.31 \text{ m mole}$ 

**(b)** 
$$X_{H_2} = \frac{n_{H_2}}{n_{Total}} = \frac{P_{H_2}}{P_{Total}} = \frac{729}{748}$$
  
= 0.975

(c) 
$$P_{\text{Total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}} \implies 748 = P_{\text{H}_2} + 19$$

:. 
$$P_{H_2} = 748 - 19 = 729 \text{ Torr} = \frac{729}{760} \text{ atm}$$

$$n_{H_2} = \frac{P_{H_2}V}{RT} = \frac{\frac{729}{760} \times 0.300}{0.0821 \times 294} = 0.0119 \text{ mol } H_2$$

mass of gas (H<sub>2</sub>) =  $n_{H_2} \times M_{gas} = 0.0119 \times 2 = 0.24$  g

**14.**  $P_{H_2O}(g) + P_{O_2} = 1.1$ 

$$0.1 + \mathsf{P}_{\mathsf{O}_2} = 1.1 \implies \mathsf{P}_{\mathsf{O}_2} = 1 \text{ atm}$$

Vol is made 
$$\frac{1}{3}$$
 rd so  $P_{O_2}$  becomes 3 times but  $P_{H_2O}$  (•)

will be same

So 
$$P_{O_2} = 3 \text{ atm}$$

Total Pressure =  $P_{O_2} + P_{H_2O}$  (•) = 3 + 0.1 = 3.1 atm



Here 
$$n_T = \frac{2.8 \times 4}{0.0821 \times 273} = 0.5$$
 and  $n_x + n_{N_2} = 0.5$   
 $\Rightarrow n_x = 0.1$   
From (1)  
 $\frac{0.4}{0.1} = \sqrt{\frac{M_x}{28}} \Rightarrow M_x = 448$ .

16. Change in pressure = 2000-1500 = 500 torr; time-taken = 40 min

rate<sub>1</sub> = 
$$\frac{500}{40}$$
 torr/min; similarly rate<sub>2</sub> =  $\frac{500}{80}$  torr/min

$$\frac{\text{rate}_1}{\text{rate}_2} = \sqrt{\frac{M}{32}}$$

8. 
$$\frac{\mathbf{r}_{\text{mix}}}{\mathbf{r}_{\text{O}_2}} = \frac{\text{V}}{\text{V}} \times \frac{220}{245} = \sqrt{\frac{\text{M}_{\text{O}_2}}{\text{M}_{\text{mix}}}}$$
  $\therefore$   $M_{\text{mix}} = 39.6$ 

$$\Rightarrow \frac{1}{\sqrt{M_{mix}}} = \frac{X_{gas}}{\sqrt{M_{gas}}} + \frac{X_{O_2}}{\sqrt{M_{O_2}}}$$
$$\Rightarrow \frac{1}{\sqrt{39.6}} = \frac{1/5}{\sqrt{M_{gas}}} + \frac{4/5}{\sqrt{32}}.$$

$$M_{gas} = 133$$

20. 
$$U_{\text{MPS}} \propto \sqrt{T} \implies \frac{(U_{\text{MPS}})_2}{(U_{\text{MPS}})_1} = \sqrt{\frac{T_2}{T_1}}$$

$$\Rightarrow \frac{2}{1} = \sqrt{\frac{T_2}{400}}$$
$$T_2 = 1600 \text{ K} = 1327^{\circ}\text{C}.$$

21. 
$$u_{rms} = \sqrt{\frac{\sum n_1 u_1^2}{\sum n}} = \sqrt{\frac{u_1^2 \times n_1 + u_2^2 \times n_2 + u_3^2 \times n_3}{n_1 + n_2 + n_3}}$$
  
 $\therefore u_{rms} = \sqrt{\frac{10^{23} \times (10^4)^2 [2 + 2 \times 4 + 2 \times 9]}{6 \times 10^{23}}}$   
 $= 2.16 \times 10^4 \text{ cm/sec}$ 

23. Volume of one mole of given vapour =  $\frac{18}{0.36}$  L=50 L

volume of one mole of an ideal gas = 
$$\frac{RT}{P} = \frac{0.082 \times 500}{1}$$
  
= 41 L

so value of. 
$$Z = \frac{V_m, real}{V_m, ideal} = \frac{50}{41}$$



24. 
$$Z_1 = \frac{P_1 V_1}{RT_1}$$
 and  $Z_2 = \frac{P_2 V_2}{RT_2}$   
 $\frac{Z_1}{Z_2} = \frac{P_1}{P_2} \times \frac{T_2}{T_1} \times \frac{V_1}{V_2} \implies V_2 = \frac{1}{3}$  litre  
26.  $T_c = 273 + 31 = 304$  K,  $P_c = 728$  atm  
 $T_c = \frac{8a}{27 \text{ Rb}}$  and  $P_c = \frac{a}{27 b^2}$   
 $\therefore \frac{T_c}{P_c} = \frac{8a}{27 \text{ Rb}} \times \frac{27 b^2}{a} = \frac{8b}{R}$   
On substitution  $\frac{304}{728} = \frac{8b}{0.082}$   
 $\therefore b = \frac{304 \times 0.082}{728 \times 8} = 4.28 \times 10^{-3}$  litre/mole  
Now  $T_c = \frac{8a}{27 \text{ Rb}}$   
 $\therefore a = \frac{27 \text{ Rb}T_c}{8} = \frac{27 \times 0.0082 \times 4.28 \times 10^{-3} \times 304}{8}$   
 $= 0.36$  atm litre<sup>2</sup> mole<sup>-2</sup> Ans.  
27.  $\frac{P_r V_r}{T_r} = \frac{P}{P_c} \frac{V_m}{V_c} \cdot \frac{T_c}{T} = \frac{PV_m}{RT} \cdot \frac{RT_c}{P_c V_c}$   
 $Z = \frac{PV_m}{RT} = 2.4 \times \left(\frac{3}{8}\right)$ 

so 
$$V_{m,real} = \frac{0.0821 \times 200}{8.21} \times 0.9 = 1.8 L$$

so volume of two moles = 3.6L.

RT

28. Significance of constant 'b'. The constant 'b' is called co-volume or excluded volume per mole of a gas. Its units are litre mol<sup>-1</sup> The volume of 'b' is four times the actual volume of the molecules.

Significance of constant 'a'. The value of constant 'a' gives the idea of magnitude of attractive forces between the molecules of the gas. Its units are atm L<sup>2</sup> mol<sup>-2</sup>. Larger the value of a, larger will be the intermolecular attraction among the gas molecules.

30. 
$$CO_2 + C \longrightarrow 2CO$$
  
At t = 0 x L  $(1-x)$  L  
At = t 0  $2x$  L  
 $1-x+2x=1.6$   
 $1+x=1.6$   
 $x=0.6$  L  
 $1-x=0.4$  L

**31.** In the mix., volume of  $O_2 = 20$  ml Volume of  $O_2 = 80 \text{ ml}$ on heating  $2O_3 \longrightarrow 3O_2$ 20 ml O<sub>3</sub> will give  $\frac{3}{2} \times 20 = 30$  ml O<sub>2</sub> total volume = 80 + 30 = 110 ml increase in volume = 110 - 100 = 10 ml. 34. HCOOH  $\longrightarrow$  H<sub>2</sub>O + CO a mole 0 а  $\begin{array}{c} H_2C_2O_4 \longrightarrow H_2O + CO + CO_2 \\ b \text{ mole } 0 & 0 & 0 \end{array}$ H<sub>2</sub>O absorb by H<sub>2</sub>SO<sub>4</sub> and CO<sub>2</sub> absorbed by KOH volume of CO<sub>2</sub> / total volume = b/a + 2b = 1/6a/b = 4/1the molar ratio of HCOOH and  $H_2C_2O_4$  is 4 : 1. **35.** Balanced chemical equation:  $C_{y}H_{y} + (x + y/4)O_{2} \rightarrow x CO_{2} + y/2 H_{2}O_{3}$ 22.4 Lt of CxHy gives = 44 a gram CO<sub>2</sub>  $1.12 \text{ Lt of CxHy gives} = 44 \text{ a} \times 1.12 / 22.4 \text{ gram CO}_{2}$  $44 a \times 1.12 / 22.4 = 2.2$ a = 122.4 Lt of CxHy gives =  $18 \times b/2$  gram H<sub>2</sub>O 1.8 Lt of CxHy gives =  $18 \times b/2 \times 1.12 / 22.4$  gram H<sub>2</sub>O  $18 \text{ b/2} \times 1.12 / 22.4 = 1.8$ b = 4Hydrocarbon is CH<sub>4</sub> wt of 1.12 Lt CH<sub>4</sub> at NTP =  $16 \times 1.12 / 22.4 = 0.8$  gram mole of O<sub>2</sub> used in the 22.4 Lt. hydrocarbon combustion = a + b/4 = 2 mole mole of O<sub>2</sub> used in the 1.12 Lt. hydrocarbon combustion  $= 2 \times 1.12 / 22.4 = 0.1$ volume of O<sub>2</sub> used in the 1.12 Lt. hydrocarbon combustion  $=0.1 \times 22.4 = 2.24$  Lt. **36.** Initially at 300 K, length of Hg column =  $\frac{152}{2}$  = 76 cm = 1 atm.

Let the volume of the cylinder be V litres

Applying ideal gas equation : (1 + 1).  $\frac{V}{2}$ = nR.300. ...(1)

Finally, say at a temp. T : length of Hg column =  $\frac{76}{2}$  cm = 0.5 atm.

:. 
$$(1+0.5) \frac{3V}{4} = nRT.$$
 ... (2)  
From equation (1) and (2), T = 337.5 K.



37.	$P_0 = 76 \text{ cm}$	X	$P_0 = 76 \text{ cm}$	
	45 cm	10cm	45 cm	•

Initially: (A)  $76 \times 45 \times A = 76 \times 45 \times A$  (B)

A = Area of cross section.When tube is made vertical, let Hg column gets displaced by x cm towards A.



For A side :  $P_1 \times (45 - x) \times A = 76 \times 45 \times A$ For B side :  $P_2 \times (45 + x) \times A = 76 \times 45 \times A$ also  $P_1 = P_2 + 10$ 

$$\frac{76 \times 45}{(45 - x)} = \frac{76 \times 45}{(45 + x)} + 10$$
  

$$\Rightarrow 76 \times 45 \left[ \frac{1}{(45 - x)} - \frac{1}{(45 + x)} \right] = 10$$
  

$$76 \times 45 \left[ \frac{45 + x - 4s + x}{(45 - x)(45 + x)} \right] = 10$$

**39.** (a) By Dalton's partial pressure

$$P_{N_2} + P_{H_2O} = 760 \, \text{mm}$$

From given data  $P_{N_2} = 745 \text{ mm}$ 

So 
$$P_{H_2O} = 760 - 745 = 15 \text{ mm}$$

% Mole of N<sub>2</sub> = % of pressure of N<sub>2</sub> =  $\frac{745}{760} \times 100 = 98.02$ 

... Mole % of  $H_2O = 100 - 98.02 = 1.98\%$  Ans. (b) Increase weight of drying agent due to absorption of water ( $H_2O$ ).

Hence, Wt. of 
$$H_2O = 0.15 \text{ g}$$
  

$$\therefore \text{ Mole of } H_2O = \frac{0.15}{18}$$

Pressure of H<sub>2</sub>O(( $P_{N_2}$ ) = 15 mm =  $\frac{15}{760}$  atm From gas equation PV = nRT

$$\frac{15}{760} \times V = \frac{0.15}{18} \times 0.0821 \times (273 + 20)$$
  
V = 10.156 litres Ans.

**41.** 
$$\left(\frac{P_1V_1}{T_1}\right)_{\text{Inside cylinder}} = \left(\frac{P_2V_2}{T_2}\right)_{\text{Outside cylinder}}$$

$$\frac{20 \times 2.82}{300} = \frac{1 \times V_2}{273}$$

 $V_2 = 51.324 L$ 

Volume of gas at STP in cylinder = 51.324 L Volume of gas left inside cylinder = 2.82 L Volume of gas available to be filled in balloon = 48.504 L Let n balloons are filled

$$\therefore \quad \frac{4}{3} \pi \times \left(\frac{21}{2}\right)^3 \times \frac{n}{1000} = 48.504$$
$$\implies \qquad n = 10$$

$$u_{\rm AV} = \sqrt{\frac{8\,\rm RT}{\pi\rm M}}$$

For He: 
$$4 \ge 10^2 = \sqrt{\frac{8 \text{RT}}{\pi \ge 4 \ge 10^{-3}}}$$

so 
$$RT = \frac{16 \times 10^4 \times \pi \times 1 \times 10^{-3}}{8} = 80 \,\pi$$

for Ne: 
$$4 \times 10^2 = \sqrt{\frac{8RT}{\pi \times 20 \times 10^{-3}}}$$

so RT = 
$$\frac{16 \text{ x } 10^4 \text{ x } \pi \text{ x } 20 \text{ x } 10^{-3}}{8} = 400\pi$$

$$\Rightarrow KE = \frac{3}{2} nRT$$

 $\therefore \quad \text{Total KE of He} = \frac{3}{2} \times \frac{6}{4} \times 80\pi = 565.71 \text{ J/mol}$ 

:. Total KE of Ne =  $\frac{3}{2} \times \frac{12}{20} \times 400\pi = 1131.42 \text{ J/mol}$ Total K.E. of mixture = 565.71 + 1131.42 J = 1697.13 J/mol

Total mole in mixture = 
$$\frac{6}{4} + \frac{12}{20} = 1.5 + 0.6 = 2.1$$
.

KE / mole of mixture = 
$$\frac{1697.13}{2.1}$$
 = 808.16 J/mol



45. (a) Under low pressure region, V is high

$$(V-b) \approx V \implies \left(P + \frac{a}{V^2}\right) V = RT$$
$$\implies PV + \frac{a}{V} = RT$$
$$\implies \frac{PV}{RT} + \frac{a}{RTV} = 1$$
$$\therefore \quad Z = \frac{PV}{RT} = \left(1 - \frac{a}{RTV}\right)$$
$$Z = 1 - \frac{20.39}{0.0821 \times 350 \times 35} = 0.98$$

(b) Under high pressure region, P is high,

$$\left(\frac{P + \frac{a}{V^2}}{V^2}\right) \approx P \qquad \therefore \quad P(V-b) = RT$$

$$\Rightarrow PV - Pb = RT \qquad \Rightarrow \qquad \frac{PV}{RT} - \frac{Pb}{RT} = 1$$

$$\therefore \quad Z = \frac{PV}{RT} = 1 + \frac{Pb}{RT}$$

$$Z = 1 + \frac{b}{V} \quad \text{by approximation} \left(Q\frac{PV}{RT} = 1, \frac{P}{RT} = \frac{1}{V}\right)$$

$$= 1 + \frac{0.1383}{35} = 1 + 0.004 = 1.004$$

46. Mole of H<sub>2</sub> diffused = 0.7 in 20 min Mole of gas diffused = n<sub>1</sub> in 20 min for gaseous mixture after diffusion PV = nRT  $n = 6 \times 3 / 0.0821 \times 300 = 0.731$ Mole of mixture (n) = mole of H<sub>2</sub> + mole of gas  $0.731 = 0.7 + n_1 \implies n_1 = 0.0308$ 

$$\mathbf{r}_{\mathrm{H}_{2}}/\mathbf{r}_{\mathrm{g}} = \sqrt{\frac{\mathrm{Mg}}{\mathrm{M}_{\mathrm{H}_{2}}}} \implies 0.7/0.0308 = \sqrt{\frac{\mathrm{Mg}}{2}}$$

47. At constant volume and temperature P  $\alpha$  W (here, volume of balloon is assumed to be constant) Thus, for N<sub>2</sub>: P<sub>1</sub> = 2 atm P<sub>2</sub> = 1/2 atm at t = 1 hr

$$\frac{P_1}{P_2} = \frac{w_1}{w_2} \quad \text{or} \quad w_2 = \frac{P_2}{P_1} \times w_1 \quad \text{or} \quad w_2 = \frac{\frac{1}{2}}{2} \times 14 = \frac{14}{4}$$
  
$$\therefore \quad \text{wt of } N_2 \text{ diffused} = 14 - \frac{14}{4} = \frac{21}{4} \text{ kg}$$
  
For  $H_2: P_1 = 2 \text{ atm } P_2 = \frac{1}{2} \text{ atm} \qquad \text{at } t = t \text{ hr}$   
$$w_1 = 1 \text{ kg} \qquad w_2 = ?$$

$$w_2 = \frac{P_2}{P_1} \times w_1 = \frac{1}{4} \text{ kg}$$

Hence wt of H<sub>2</sub> diffused =  $1 - \frac{1}{4} = \frac{3}{4}$  kg Now, we are to conclude one point as

$$\frac{r_{A}}{r_{B}} = \sqrt{\frac{\rho_{B}}{\rho_{A}}}$$
, as per Graham law

$$\frac{\overline{V_A}}{\overline{V_B}} = \sqrt{\frac{\rho_B}{\rho_A}} \quad \text{or} \quad \frac{\overline{V_A}}{\overline{V_B}} \times \frac{t_B}{t_A} = \sqrt{\frac{\rho_B}{\rho_A}} \quad \text{or}$$

$$\frac{V_A}{V_B} \times \frac{t_B}{t_A} = \frac{\rho_B}{\rho_A} \sqrt{\frac{\rho_A}{\rho_B}}$$

or 
$$\frac{V_A \rho_B t_B}{V_B \rho_B t_A} = \sqrt{\frac{\rho_A}{\rho_B}}$$
 or  $\frac{w_A}{w_B} \times \frac{t_B}{t_A} = \sqrt{\frac{\rho_A}{\rho_B}}$ 

For our problem we can write,

$$\frac{W_{H_2} \times t_{N_2}}{W_{N_2} \times t_{H_2}} = \sqrt{\frac{M_{H_2}}{M_{N_2}}} \quad \text{or} \quad \frac{\frac{3}{4} \times 1}{\frac{21}{2} \times t} = \sqrt{\frac{2}{28}} = \sqrt{\frac{1}{14}}$$
  
or 
$$\frac{6}{21 \times 4} \times \frac{1}{t} = \sqrt{\frac{1}{14}} \quad \text{or} \quad \frac{1}{14t} = \frac{1}{\sqrt{14}} \quad \text{or}$$
$$\sqrt{\frac{14}{21}} = \frac{1}{\sqrt{14}} \quad \frac{60}{21} = \frac{60}{21} \quad \frac{60}{21} = \frac{1}{\sqrt{14}} \quad \frac{1}{\sqrt{14}} \quad \frac{1}{\sqrt{14}} = \frac{1}{\sqrt{14}} \quad \frac{1}{\sqrt{14}} = \frac{1}{\sqrt{14}} \quad \frac{1}{\sqrt{14}} \quad \frac{1}{\sqrt{14}} = \frac{1}{\sqrt{14}} \quad \frac{1}{\sqrt{14}}$$

$$t = \frac{\sqrt{14}}{14} = \frac{1}{\sqrt{14}}$$
 hr  $= \frac{300}{\sqrt{14}}$  mins  $= \frac{300}{3.741}$  mins  $= 16$  mins

 $\therefore$  for H<sub>2</sub>, 16 mins are required **Ans.** 

48. For the gaseous mixture PV = nRT  $1 \times 40 = n \times 0.0821 \times 400$  n = 1.219Total mole of mix. = 1.219 Suppose moles of C<sub>2</sub>H<sub>6</sub> and C<sub>2</sub>H<sub>4</sub> are a and b respectively a+b=1.219 ......(1)

$$C_2H_6 + \frac{7}{2}O_2 \rightarrow 2CO_2 + 3H_2$$
  
a  $7a/2$   
b  $3b$   
b  $2b$ 

moles of  $O_2$  required for the combustion of mix. = 7a/2 + 3b7a 130

$$\frac{74}{2} + 3b = \frac{100}{32}$$
 .....(2)

by solving eq. (1) and (2) we get a=0.808 b=0.4115

mole fraction of  $C_2H_6 = \frac{0.808}{1.219} = 0.66$ mole fraction of  $C_2H_4 = \frac{0.4115}{1.219} = 0.34$ 

49. Volume of H<sub>2</sub>O produced = 17 mL; Volume of CO<sub>2</sub>  
produced = 14 mL  
Total volume of CO<sub>2</sub> = volume of CO<sub>2</sub> initially + volume  
of CO<sub>2</sub> produced = 14 ml  
suppose volume of CH<sub>4</sub> and C<sub>2</sub>H<sub>4</sub> in the mixture x and y  
ml respectively  
volume of CO<sub>2</sub> produced on explosion = 14 - (10 - x - y)  
= (4 + x + y) ml  
POAC for C, H and O  
for C, 1 × mole of CH<sub>4</sub> + 2 × mole of C<sub>2</sub>H<sub>4</sub> = 1 × mole of CO<sub>2</sub>  
$$x + 2y = 4 + x + y$$
 y = 4  
for H, 4x + 4y = 2 × mole of H<sub>2</sub>O .......(i)  
for O, 2 × mole of O<sub>2</sub> = 2 × (4 + x + y) + mole of H<sub>2</sub>O  
.......(ii)  
for equation (i) and (ii)  
mole of O<sub>2</sub> = (4 + 2x + 2y) (use in explosion)  
in explosion reaction  
volume of reactant - volume of product = 17 ml  
(x + y) + (4 + 2x + 2y) - (4 + x + y) = 17  
2x + 2y = 17 .......(iii)  
from equation (i) and (iii)  
(volume of CH<sub>4</sub>) = 4.5  
(volume of C<sub>2</sub>H<sub>4</sub>) = 4 ml  
(volume of CO<sub>2</sub>) = 4 ml

**1.** K.E. = 
$$\frac{3}{2}$$
 RT

$$E_1 = \frac{3}{2} R 293$$
 and  $E_2 = \frac{3}{2} R 313 \implies E_2 = \frac{313}{293} \times E_1$ 

- 2.  $b = 4N \times v$  i.e., total volume occupied by molecules of one mole of gas in motion.
- 3. Helium is not used to produced and sustain powerful superconducting magnets. All others are the uses of helium.
- 4. Distribution of molecules (N) with velocity ( $\mu$ ) at two temperature T<sub>1</sub> and T<sub>2</sub> (T<sub>2</sub> > T<sub>1</sub>) is show below:



At both temperatures, distribution of molecules with

increase in velocity first increases, reaches a maximum value and then decreases.

5. Let the mass of methane and oxygen be m gm. Mole fraction of oxygen  $x_{o_2}$ 

$$=\frac{\frac{m}{32}}{\frac{m}{32}+\frac{m}{16}}=\frac{m}{32}\times\frac{32}{3m}=\frac{1}{3}$$
 Let the total pressure be P.

$$\therefore \text{ Partial pressure of } O_2, P_{o2} = P \times x_{o2} \qquad P \times \frac{1}{3} = \frac{1}{3} P.$$

6. a b  

$$Cl_2$$
 6.579 L<sup>2</sup> bar mol<sup>-2</sup> 0.05622 L mol<sup>-1</sup>  
 $C_2H_5$  5.562 L<sup>2</sup> bar mol<sup>-2</sup> 0.06380 L mol<sup>-1</sup>

$$r \propto \frac{p}{\sqrt{M}}$$
$$\frac{r_{A}}{r_{B}} = \frac{p_{A}}{p_{B}} \sqrt{\frac{M_{B}}{M_{A}}}$$
$$v = \sqrt{\frac{8RT}{\pi M}}$$

v ∝ √T

7

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

at high pressure  $\frac{a}{V^2}$  can be neglected.

$$PV - Pb = RT$$
$$PV = RT + Pb$$
$$\frac{PV}{RT} = 1 + \frac{Pb}{RT}$$

$$Z = 1 + \frac{10}{RT} \qquad ;$$

Z > 1 at high pressure

10. C\* = most probable speed = 
$$\sqrt{\frac{2RT}{M}}$$

$$\overline{C}$$
 = average speed =  $\sqrt{\frac{8RT}{\pi M}}$ 

C = Mean square speed corrected as rms = 
$$\sqrt{\frac{3RT}{M}}$$

$$C^* < \overline{C} < C$$



$$C^*: \overline{C}: C = 1: \sqrt{\frac{4}{\pi}}: \sqrt{\frac{3}{2}} = 1: 1.128: 1.225$$

- 11. 18g H₂O contains 2g H
   ∴ 0.72 g H₂O contains 0.08 gH.
   44 g CO₂ contains 12g C
  - $\therefore$  3.08 g CO<sub>2</sub> contains 0.84 g C

$$\therefore C: H = \frac{0.84}{12}: \frac{0.08}{1} = 0.07: 0.08$$
$$= 7: 8$$

 $\therefore$  Empirical formula = C<sub>7</sub>H<sub>8</sub>

$$12. \quad \left(P + \frac{n^2 a}{V^2}\right) (V - b) = nRT$$

For 1 mole, 
$$\left(P + \frac{a}{V^2}\right)(V-b) = RT$$
  
 $PV = RT + Pb - \frac{a}{V} + \frac{ab}{V^2}$ 

at low pressure, terms Pb &  $\frac{ab}{V^2}$  will be negligible as

compared to RT.

So, 
$$PV = RT - \frac{a}{V}$$
$$Z = 1 - \frac{a}{RTV}$$

**13.** Moles of  $O_2 = \frac{W}{32}$ 

Moles of N<sub>2</sub> =  $\frac{4w}{28}$ 

 $\frac{n_{O_2}}{n_{N_2}} = \frac{w}{32} \times \frac{28}{4w} = \frac{7}{32}$ 

14. Applying Mole conservation

$$\frac{p_1^0 V}{RT_1} + \frac{P_i V}{RT_1} = \frac{P_f V}{RT_1} + \frac{P_f V}{RT_2}$$
$$\frac{2P_i}{T1} = P_f \left\{ \frac{1}{T_1} + \frac{1}{T_2} \right\}$$
$$P_f = \frac{2P_i}{T_1} \left\{ \frac{T_1 T_2}{T_1 + T_2} \right\} = \frac{2P_1^0 T_2}{T_1 + T_2}$$

# Part # II : IIT-JEE ADVANCED

 $rac{r_{
m He}}{r_{
m CH_4}}$ 

16

- 1.  $\frac{\text{rate of diffusion}_2}{\text{rate of diffusion}_1} = \sqrt{\frac{M_1}{M_2}}$
- 2. The vander waals equation of state is (for 1 mole of gas)

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

When a is negligible, then

$$Z = \frac{pV_m}{RT} = 1 + \frac{b}{RT}P$$

that is Z increases with increaser in p. When b is negligible, then

$$Z = \frac{pV_m}{RT} = 1 - \frac{a}{VRT}$$

increasing p implies decrease in V, which is turn, implies increase in the value of a/VRT and hence decrease in the value of Z.

The curve IV is applicable provided temperature of the gas is near but larger than it's critical temperature Hence, the choice (a), (b) and (c) are correct.

- 3. (A) For  $H_2$  gas at high pressure Z > 1.
  - (B) For any gas at  $P \sim 0$ ,  $Z \sim 1$  i.e. ideal behaviour.
  - (C) For  $CO_2$  gas at normal pressure and temperature Z < 1. (D) For any gas at very large molar volume i.e.  $P \sim 0, Z \sim 1$

i.e. ideal behaviour of gas.

4. (A) At very large molar volume

$$P + \frac{a}{V_m^2} \approx P \text{ and } V_m - b = V_m$$

(C) According to van der Waals equation 'a' and 'b' are independent of tempt.

5. Correction factor for attractive force in to the real gas is

given by 
$$\frac{an^2}{V^2}$$

 $6. V_{\rm rms} = V_{\rm mp}$ 

$$\sqrt{\frac{3RT}{M_{X}}} = \sqrt{\frac{2RT}{M_{Y}}} \qquad \Rightarrow \quad \sqrt{\frac{3R \times 400}{40}} = \sqrt{\frac{2R \times 60}{M_{Y}}}$$
$$M_{Y} = 4.$$

7. (A) Fact

**(B)** 
$$P = MV = M\sqrt{\frac{3RT}{M}} = \sqrt{3MRT}$$





8.  $P_{He} = 1 - 0.68 = 0.32$  atm V=?

n=0.1  

$$V = \frac{nRT}{P} = \frac{0.1 \times 0.0821 \times 273}{0.32} = 7$$
9. PV 21.6  
20.1  
21.6  
20.1  
20.1  
2.0 1/v  $\rightarrow$  3.0

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

PV + a/V = RTPV = RT - a(v)y = RT - a(x)

So, slope = 
$$a = \frac{21.6 - 20.1}{3 - 2} = \frac{1.5}{1} = 1.5$$

**10.** 
$$\lambda = \frac{h}{\sqrt{2m(KE)}}$$
 KE  $\propto$  7

$$\frac{\lambda_{\text{He}}}{\lambda_{\text{Ne}}} = \sqrt{\frac{m_{Ne}KE_{Ne}}{m_{He}KE_{He}}} = \sqrt{\frac{20\times1000}{4\times200}} = 5.$$

 $K \rightarrow L \Longrightarrow V \uparrow at constant P$ 

Hence T  $\uparrow$  (Heating)

$$L \rightarrow M \Longrightarrow P \downarrow \text{ at constant V}$$

Hence T  $\downarrow$  (Cooling)

 $M \rightarrow N \Rightarrow V \downarrow at constant P$ 

Hence T  $\downarrow$  (Cooling)

 $N \rightarrow K \Rightarrow P \uparrow at constant V$ 

Hence T 
$$\uparrow$$
 (Heating)

13. 
$$L \rightarrow M$$

 $M \rightarrow K$ 

Both are having constant volume therefore these processes are isochoric.

16.  

$$\begin{array}{c|c}
1 & m^{3}, 5 \text{ bar}, \\
400 \text{ k} \\
A \\
\end{array} & 3 & m^{3}, 1 \text{ bar}, 300 \text{ K} \\
B \\
\end{array}$$
Finally,  $P_{A} = P_{B}$  also  $T_{A} = T_{B}$   
So  $\frac{n_{A}}{n_{B}} = \frac{V_{A}}{V_{B}}$   
 $\frac{5}{400\text{ R}} = \frac{V_{A}}{V_{B}}$   $\Rightarrow \frac{V_{A}}{V_{B}} = \frac{5}{4}$   
 $\Rightarrow V_{A} = \frac{5}{9} \times 4 = \frac{20}{9} = 2.22$   
**NOCK TEST**  
1.  $V_{real} = \frac{\text{Molar mass}}{\text{density}} = \frac{18}{0.36}$   
 $V_{ideal} = \frac{n\text{RT}}{P} = \frac{1 \times 0.082 \times 500}{1}$   
So,  $Z = \frac{V_{real}}{V_{ideal}} = \frac{1 \times 0.082 \times 500}{1} = \frac{50}{0.082 \times 500} = \frac{50}{41}$   
2.  $PM = dRT$   
So,  $M = \frac{dRT}{P} = 75 \text{ (approx.)}$ 

Normal molecular wt. of HF = 18

i.e. HF is tetramer  $(HF)_4$  in gaseous state due to H - bonding.

3. 
$$E_1 = \frac{3}{2} \times \frac{M}{16} RT_1$$
  
 $E_2 = \frac{3}{2} \times \frac{M}{30} RT_2 \implies \frac{E_1}{E_2} = \frac{30}{16} \frac{T_1}{T_2}$   
 $\frac{3}{1} = \frac{30}{16} \frac{T_1}{T_2} \implies \frac{T_1}{T_2} = \frac{8}{5}$ 

4. At constant pressure PV=nRT

$$V = \left(\frac{nR}{P}\right)T \text{ So, } \log V = \log\left(\frac{nR}{P}\right) + \log T$$

y = C + mx.So answer is **(B)**.

- 5. For constant pressure  $V \propto T$ , hence linear dependence with the slope.
- Let suppose we use x-litre water gas at S.T.P. So it contain 0.5x litre H<sub>2</sub>. Now according to the given equation

 $\rightarrow$  CO<sub>2</sub> + CO +H,O Η, 0.45x Excess 0.5x 0.05x ( 0 (0.05x + 0.45x) (0.5x + 0.45x)So total volume of  $H_2$  gas = 0.95x But according to question It is 5 litre So 0.95x = 5x = 5.263 litre

7. 
$$\frac{r_{O_3}}{r_{O_2}} = \sqrt{\frac{M_2}{M_1}}$$

$$0.98 = \sqrt{\frac{M_{O_2}}{M_{O_3}}}$$
$$0.9604 = \frac{32}{M_{O_3}}$$
$$M_{O_3} = 33.32$$

Now ozonised gas pure oxygen x% 100-x

Avg. molar mass =  $\left(\frac{M_1 x_1 + M_2 x_2}{x_1 + x_2}\right)$  $33.32 = \frac{48x + 32(100 - x)}{100}$ 100 x = 8.25. В А 88 g H. CO<sub>2</sub> 10 Litre 10 Litre 8. Thermo state at 200 K Thermo state at 500 K No. of mole of H<sub>2</sub> in flask A = 20/2 = 10 mole No. of mole of CO<sub>2</sub> in flask B = 88/44 = 2 mole Now pressure of gas in flask A PV = nRT $P_{\star} \times 10 = 2 \times R \times 200$  $P_A = 500R$ Now pressure of gas in flask B  $P_{\rm p} \times 10 = 2 \times R \times 200$  $P_{\rm B} = 40R$ Beacause flask A is on higher pressure that's why H<sub>2</sub> flows from flask A to B. Let suppose x mole of H<sub>2</sub> move from flask A to B. So mole of H<sub>2</sub> remain in A = (10 - x) and total mole in B = (2 + x)Now after opening stop cock pressure of both flask become equal.  $n_A T_A = n_B T_B$  $(10-x) \times 500 = (2+x) \times 200$ x = 6.57Composition of H<sub>2</sub> in B =  $\frac{6.57 \times 2}{101.14} \times 100 = 13\%$ 2 litre  $n_A : n_B = 2 : 3$  $n_{A'} : n_B = 3 : 5$ 12. and  $n_{A} = 3/5 n_{B}$ So  $n_{A} = 2/3 n_{B}$ Put the values of  $n_A^{}$ ,  $n_B^{}$ ,  $n_{A^{\prime}}^{}$ ,  $n_{B^{\prime}}^{}$  and get

mean molar mass =  $\frac{(n_{A} + n_{A'})m_{A} + (n_{B} + n_{B'})m_{B}}{(n_{A} + n_{B} + n_{A'} + n_{A})}$ 

13. Applying Boyle's law.  $P_1V_1 = P_2V_2$  $10P_w.g = (10 + n) \times (2 - h) P_w.\& [A = Area of cross]$ 



section ]  $\times$  2. A h=1.71m

$$14. PV = nRT$$

$$V = \frac{nR}{P}.T \implies V = KT$$

 $\therefore \quad \log P = \log T + \log K \\ \text{Linear dependence with positive slope} \\ \text{At the y-intercept } P/_{t \to 0} \text{ inplying } P \to 0, \text{ under such conditions all gases show ideal behaviour.}$ 

15. 
$$P_{mix} = \frac{PM_{mix}}{RT} \implies M_{mix} = \frac{1.7 \times 0.0821 \times 273}{1}$$
  
= 1.7 × 22.4 gms  
 $M_{mix} = X_{co} \times M_{co} + (1 - X_{co}) \cdot M_{co2}$   
=  $X_{co} \times 18 + (1 - X_{co}) \times 44$   
 $\Rightarrow X_{co} = 0.37$   
16.  $b = \frac{16}{3} \pi r^3 N_a$ 

**18.** 
$$b = \frac{10}{3}\pi r^3 N_a$$
  
 $0.072 \times 1000 = \frac{16}{3} \times 6.023 \times 10^{23} \times \pi r^3$   
then calculate r.

### **21.** V - T curve for a given mass of the gas is given

$$V = \frac{P_{1} = \text{constant} \quad (P_{2} > P_{1})}{T}$$
34. 24.  

$$P = P_{0}e^{-kt}$$

$$P_{0} = \frac{228}{3} + 76 = 152 \text{ cm Hg.}$$
At t = 7 min.  

$$P = \frac{114}{3} + 76 = 114 \text{ cm Hg.}$$

$$P = \frac{152}{3}e^{-kt} = 114$$



$$\Rightarrow kt = \frac{1}{t} \ln \frac{152}{114} = \frac{1}{7} \times 0.28 = 0.04 \text{ min}$$
$$= 0.04 \times 60 \text{ hr}^{-1} = 2.4 \text{ hr}^{-1}$$