

### <H1> Additional Objective Questions

#### <H2> Single Correct Choice Type

1.(B) For positive deviation,  $pV = nRT + npb$ . Thus, the factor  $npb$  is responsible for increasing the  $pV$  value above ideal value. The quantity  $b$  is actually the effective volume of molecule. So, it is the finite size of molecules that leads to the origin of  $b$  and hence positive deviation at high pressure.

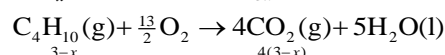
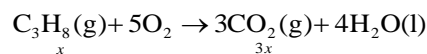
2.(B) We know

$$v_{\text{rms}} = \sqrt{\frac{1^2 + 0^2 + 2^2 + 3^2}{4}} = \sqrt{3.5}$$

3.(B) It is given that  $Z < 1$ , therefore

$$\frac{pV}{RT} < 1 \Rightarrow pV < RT \Rightarrow V < RT \Rightarrow V < 0.0821 \times 273 \Rightarrow V < 22.4 \text{ mL}$$

4.(C) The reactions involved are



The total amount of  $\text{CO}_2$  formed is 10 L, therefore,

$$3x + 4(3 - x) = 10$$

Solving, we get  $x = 2$ . Therefore, volume of  $\text{C}_3\text{H}_8 = 2 \text{ L}$  and of  $\text{C}_4\text{H}_{10} = 1 \text{ L}$ .

5.(C) Do remember that rms speed does not depend on the pressure.

$$\bar{C}_2 = \sqrt{\frac{3R \times 27}{M}} \Rightarrow \frac{\bar{C}_2}{\bar{C}_1} = \sqrt{3}$$

$$\bar{C}_1 = \sqrt{\frac{3R \times 27}{M}} \quad \text{or} \quad \bar{C}_2 = \sqrt{3} \bar{C}_1 = \sqrt{3} \times 10^4 \text{ cm s}^{-1}$$

6.(A) Molecular weight of  $\text{CO} = 28$ ; Molecular weight of  $\text{CO}_2 = 44$

Molecular weight of  $\text{N}_2\text{O} = 28 + 16 = 44$

Molecular weight of  $\text{CO}_2 = 12 + 32 = 44$

Since  $\text{CO}_2$  and  $\text{N}_2\text{O}$  have the same molecular weight, therefore, the rate of diffusion for both the gases is the same.

7.(A) Given that

$$(u_{\text{avg}})_i = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8R \times 300}{\pi M}} = 0.3 \text{ m s}^{-1} \quad (1)$$

Also, 
$$(u_{\text{avg}})_f = \sqrt{\frac{8R \times 1200}{\pi M}} \quad (2)$$

Dividing Eq. (1) by Eq. (2), we get

$$\frac{(u_{\text{avg}})_f}{(u_{\text{avg}})_i} = \sqrt{\frac{1200}{300}} \Rightarrow (u_{\text{avg}})_f = 2 \times 0.3 = 0.6 \text{ m s}^{-1}$$

**8.(A)** We know

$$\frac{u_{\text{O}_2}}{u_{\text{H}_2}} = \sqrt{\frac{M_{\text{H}_2}}{M_{\text{O}_2}}} \quad [\text{As } p_1 = p_2]$$

$$\frac{u_{\text{O}_2}}{u_{\text{H}_2}} = \sqrt{\frac{2}{32}} = \frac{1}{4}$$

**9.(A)** We know

$$u_{\text{avg}} \propto \sqrt{T}$$

$$\Rightarrow \frac{u_1}{u_2} = \sqrt{\frac{T_1}{T_2}}$$

$$\Rightarrow \frac{u_1}{u_2} = \sqrt{\frac{300}{1200}} = \frac{1}{2}$$

$$\Rightarrow u_2 = 2 \times 0.3 = 0.6 \text{ m s}^{-1}$$

**10.(D)** For an ideal gas,  $pV = nRT$ , therefore,

$$u_{\text{rms}} = \left(\frac{3RT}{M}\right)^{1/2} = \left(\frac{3pV/n}{M}\right)^{1/2}$$

$$u_{\text{rms}} = \left(\frac{3pV \times M/w}{M/V}\right)^{1/2} = \left(\frac{3p}{\rho}\right)^{1/2}$$

**11.(A)** We know

$$\frac{u_{\text{rms}}}{u_{\text{mp}}} = \frac{(3RT/M)^{1/2}}{(2RT/M)^{1/2}} = \left(\frac{3}{2}\right)^{1/2} = 1.22$$

**12.(C)** Given that  $u_{\text{mp}} = 400 \text{ m s}^{-1}$ , therefore,  $u_{\text{rms}} = 1.22 \times 400 = 488 \text{ m s}^{-1}$

**13.(D)** Assuming ideal gas behavior of gas A.

We have 
$$pV = (n_1 + n_2)RT \quad (1)$$

Substituting the given values in Eq. (1), we get

$$200 \times 10 = (0.5 + x)1000 \times R$$

$$2 = (0.5 + x)R$$

$$x = \frac{4 - R}{2R}$$

**14.(A)** The average kinetic energy of molecule is given by  $\frac{3}{2} \frac{RT}{N}$  where  $\frac{R}{N}$  is Boltzmann constant.

The Boltzmann constant is  $1.9872 \text{ cal mol}^{-1} \text{ K}^{-1}$ . Substituting, we get

$$1800 = 2 \times \frac{3}{2} \times 1.98 \times T \Rightarrow T = 300 \text{ K}$$

**15.(B)** The relative rates of diffusion of gases with respect to molecular weights is given by the expression.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}; \quad M_2 = 64 \text{ and } M_1 = 4$$

**16.(D)** We know

$$\frac{n_f}{n_i} = \frac{p_f}{p_i} \Rightarrow p_f = p_i \frac{n_f}{n_i} \quad \text{or} \quad 2 \times \frac{3}{2} = 3 \text{ atm}$$

**17.(B)** Boyle's temperature,

$$T_B = \frac{a}{Rb} = \frac{4 \text{ L}^2 \text{ atm. mol}^{-2}}{0.082 \text{ L atm. K}^{-2} \times 0.04 \text{ L}^2 \text{ mol}^{-1}} = 1219.5 \text{ K} = 1220 \text{ K}$$

**18.(A)** Given,  $p_{O_2,i} = \frac{3RT}{V_i}$   $p_{O_2,f} = \frac{2RT}{V_f}$

$$\frac{10RT}{V_i} = \frac{9RT}{V_f}$$

[Given,  $p_i = p_f$ ]

Or

$$V_f = \frac{9V_i}{10}$$

The partial pressure of  $O_2$  is

$$\frac{p_{O_2,i} - p_{O_2,f}}{p_{O_2,i}} = \frac{3 - \frac{20}{9}}{3} = 26\%$$

**19.(A)** As the peak of the curve decreases, temperature increases.

**20.(A)** We know

$$T_c = \frac{8a}{27Rb} = \frac{8 \times 371.843 \times 1000 \times 101325}{27 \times 0.0821 \times 40.8} = 324.8 \text{ K}$$

**21.(A)** Using the expression  $pV = nRT$ , we have

$$n = \frac{pV}{RT}$$

$$n = \frac{\left\{ \left( \frac{7.6 \times 10^{-10}}{760} \right) \text{ atm} \right\} (1 \text{ L})}{(0.0821 \text{ atm L K}^{-1} \text{ mol}^{-2})(273.15 \text{ K})} = 4.459 \times 10^{-14} \text{ mol}$$

$$N = nN_A = (4.459 \times 10^{-14} \text{ mol}^{-1}) = 2.686 \times 10^{10}$$

**22.(C)** Graphs **(A)**, **(B)**, and **(D)** represent Boyle's law, that is,  $pV = K$  (constant). Graph **(C)** does not represent Boyle's law since  $pV \neq \text{Constant}$ .

**23.(C)** Since  $\text{KE} = \frac{1}{2}Mu_{\text{rms}}^2$ , therefore,

$$u_{\text{rms}} = \sqrt{\frac{2E}{M}} \quad \left( \text{using } u_{\text{rms}} = \left( \frac{3RT}{M} \right)^{1/2} \text{ and } pV = \frac{1}{3}mnu^2 \right)$$

**24.(B)**  $b$  relates to the incompressible volume of the molecules and measures the effective size of gas molecules.  $Z = pV/RT = 1$  for an ideal gas.

**25.(C)** Ideal gas will obey all gas laws. So, as per Charles' law,

$$V_2 = \left( \frac{T_2}{T_1} \right) \times V_1 = \left( \frac{373}{273} \right) \times 22.4 = 30.6 \text{ L}$$

**26.(B)** We know

$$Z = \frac{pV_m}{RT}$$

$$Z = \frac{pV}{nRT} = \frac{pV \times M_0}{g \times RT} = \frac{pM_0}{\rho RT} \quad (1)$$

Substituting values in Eq. (1), we get

$$Z = \frac{0.5 \times 5}{0.3 \times 0.082 \times 300} = 0.34 < 1.$$

Therefore, attractive forces dominate, and gas is more liquefiable.

**27.(C)** In the expression,  $pV = RT \left[ A + \frac{B}{V} + \frac{C}{V^2} + \dots \right]$ , the first term, that is, A within the brackets is the main term and the rest are correction term for non-ideality. Therefore, A should be 1, since  $PV = RT$  per mole of an ideal gas. The van der Waals' equation for 1 mol is

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

Expanding the above equation, we get

$$\begin{aligned} pV &= RT + bp - \frac{a}{V} + \frac{ab}{V^2} \approx RT + bp - \frac{a}{V} \\ &= RT \left[ 1 + \frac{bp}{RT} - \frac{a}{RTV} \right] \end{aligned}$$

Applying the ideal gas equation in the correction term, we get

$$pV = RT \left[ 1 + \frac{b}{V} - \frac{a}{RTV} \right] = RT \left[ 1 + \frac{1}{V} \left( b - \frac{a}{RT} \right) \right]$$

Thus,

$$A = 1; B = \left( b - \frac{a}{RT} \right)$$

**28.(B)** Molecular speed of gases are as follows:

(I) The rms (root mean square) velocity

$$C = \sqrt{\frac{3RT}{M}}$$

(II) Average velocity =  $\sqrt{\frac{8RT}{\pi M}}$

$$v_{\text{rms}} = 1.085 \times v_{\text{avg}}$$

(III) Most probable velocity =  $\sqrt{\frac{2RT}{M}}$

The ratio is:

Most probable velocity: Average velocity : rms

$$\sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3} = 1 : 1.128 : 1.224$$

**29.(C)** At a given temperature,

$$\text{Rate of effusion} \propto p \text{ and Rate of effusion} \propto \frac{1}{\sqrt{M}}$$

$$\text{Rate of effusion of hydrogen gas} = 10 \text{ torr min}^{-1} \times \frac{2000}{1000} \sqrt{\frac{4}{2}} = 20\sqrt{2} \text{ torr min}^{-1}$$

**30.(C)** As per equation of state,

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

Given that  $p_2 = \frac{p_1}{2}$ ,  $p_2 = 2T_1$

So, 
$$V_2 = \frac{T_2}{T_1} \times \frac{p_1}{p_2} = \frac{2T_1}{T_1} \times \frac{p_1}{p_1/2} \times V_1 = 4V_1$$

**31.(B)** The measure of forces of attraction for  $n$  moles of real gas ( $n^2a/V^2$ )

$$\left(p + \frac{n^2a}{V^2}\right)(V - nb) = nRT$$

**32.(C)** This is one of the assumption of kinetic theory of gases.

**33.(C)**  $T_1 = 273 + 20 = 293$  K and  $T_2 = 273 + 40 = 313$  K. As the kinetic energy is proportional to the temperature, we have

$$\frac{E_1}{E_2} = \frac{T_1}{T_2} = \frac{293}{313} \Rightarrow \frac{E_2}{E_1} = \frac{313}{293}$$

**34.(C)** As  $pV = nRT$ , therefore,  $p \propto 1/V$  at constant  $T$ .

**35.(B)** As Rate  $\propto$  Temperature

Therefore, 
$$\frac{r_{\text{SO}_2}}{r_{\text{N}_2}} = \sqrt{\frac{M_{\text{N}_2}}{M_{\text{SO}_2}}} \times \frac{T_{\text{SO}_2}}{T_{\text{N}_2}} \quad (1)$$

Substituting the values in Eq. (1), we get

$$\begin{aligned} \frac{x}{1.625x} &= \sqrt{\frac{28}{64}} \times \frac{773}{T_{\text{N}_2}} \\ \Rightarrow T_{\text{N}_2} &= \sqrt{\frac{28}{64}} \times 773 \times 1.625 \\ \Rightarrow T_{\text{N}_2} &= \sqrt{7} \times \frac{773 \times 1.625}{4} = 830 \text{ K} \end{aligned}$$

**36.(C)** At constant  $p$  and  $V$ ,  $n_1T_1 = n_2T_2$ ; as  $pV = nRT$

$$4 \times T_1 = 3.2 \times (T_1 + 50)$$

On solving, we get  $T_1 = 200$  K

**37.(C)** Volume occupied by the molecules. The van der Waals equation for 1 mol of gas is:

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

where  $b$  is the measure of effective size of the gas molecules. It is also called co-volume or excluded volume.

**38.(B)** According to Maxwell's distribution, the most probable velocity increases as the fraction of molecules possessing most probable velocity decreases.

**39.(B)** We know

$$\frac{r_{\text{H}_2}}{r_{\text{gas}}} = \left(\frac{M_{\text{gas}}}{M_{\text{H}_2}}\right)^{1/2} \quad (1)$$

Now,

$$r_{\text{H}_2} = \frac{V_{\text{H}_2}}{\text{Time interval}} = \frac{X}{5}$$

$$\frac{r_{\text{H}_2}}{r_{\text{gas}}} = \frac{X/5}{X/t} = \frac{t}{5} \quad (2)$$

Working with the options one by one using Eq. (1), we find that the best fit is  $\text{O}_2$ .

$$\frac{r_{\text{H}_2}}{r_{\text{O}_2}} = \left(\frac{M_{\text{O}_2}}{M_{\text{H}_2}}\right)^{1/2} = 4$$

Substituting in Eq. (2), we get

$$4 = \frac{t}{5} \Rightarrow t = 20 \text{ s}$$

**40.(D)** In an ideal gas, there are no intermolecular forces of attraction. So,  $T_f = T_i$  for both reversible and irreversible processes.

**41.(A)** Since,  $r \propto$  area of orifice

$$r_1 \propto \pi r^2 \text{ (circle)}$$

$$r_2 \propto r^2 \text{ (square)}$$

Therefore,  $r_1/r_2 = \pi$ .

**42.(C)** The rms velocity is given by

$$u_{\text{rms}} = \left( \frac{3RT}{M} \right)^{1/2}$$

Given that

$$\frac{(u_{\text{rms}})_{\text{H}_2}}{(u_{\text{rms}})_{\text{N}_2}} = \sqrt{7}$$

Therefore,

$$\frac{(u_{\text{rms}})_{\text{H}_2}}{(u_{\text{rms}})_{\text{N}_2}} = \sqrt{7}$$

$$\Rightarrow \frac{(3RT_{\text{H}_2} / 2)^{1/2}}{(3RT_{\text{N}_2} / 28)^{1/2}} = \sqrt{7}$$

$$\Rightarrow 4T_{\text{H}_2} = T_{\text{N}_2}$$

**43.(B)** Using Boyle's law,  $p_1V_1 = p_2V_2$  where  $T = \text{constant}$ ,  $p_1 = 741.3 \text{ mm Hg}$  and  $V_1 = 0.357 \text{ L}$ ,  $p_2 = 758.1 \text{ mm Hg}$  and  $V_2 = ?$

$$V_2 = \frac{p_1V_1}{p_2} = \frac{741.3 \text{ mm Hg} \times 0.357 \text{ L}}{758.1 \text{ mm Hg}} = 0.349 \text{ L}$$

The volume of the balloon will decrease.

**44.(C)** This is based on Graham's law of diffusion.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_1}{M_2}} \quad (1)$$

Substituting values in Eq. (1), we get

$$\frac{45}{60} = \sqrt{\frac{40}{M_2}}$$

$$\Rightarrow M_2 = \frac{40 \times 60 \times 60}{45 \times 45} = 71.1 \text{ g mol}^{-1}$$

**45.(B)** We know

$$pV = \frac{w}{M} RT \quad (1)$$

Substituting the values in Eq. (1), we get

$$\frac{800}{760} \times \frac{380}{1000} = \frac{0.455}{M} \times 0.0821 \times 300$$

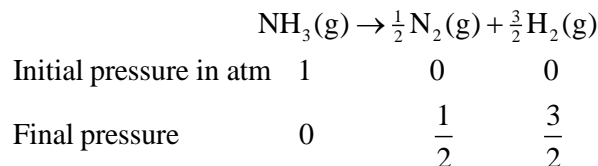
$$\Rightarrow M = \frac{0.455 \times 0.0821 \times 300 \times 5}{2} = 28.016 \text{ g mol}^{-1}$$



**46.(D)** Check for yourself that all are correct.

- Decrease with increase in concentration—means increase in the number of molecules.
- Increase with decrease in pressure at constant temperature—means increase in the volume available for gas molecules.
- Decreases with increase in molecular size—means decrease in the volume of available for gas molecules.

**47.(A)** Ammonia decomposes into  $N_2$  and  $H_2$  completely when a spark is passed through it.



The total pressure of  $N_2$  and  $H_2$  after complete decomposition is given by

$$p_T = p_{N_2} + p_{H_2} = 0.5 + 1.5 = 2 \text{ atm}$$

**48.(A)** We know

$$\frac{n_{\text{methane}}}{n_{O_2}} = \frac{w_{CH_4}}{w_{O_2}} \times \frac{M_{O_2}}{M_{CH_4}} = \frac{32}{16} = 2$$

According to Dalton's law of partial pressure  $p_{CH_4} = x_{CH_4} \times P_{\text{total}}$

$$\text{As } x_{CH_4} = \frac{2}{3} \text{ and } x_{O_2} = \frac{1}{3}, \text{ so } p_{O_2} = x_{O_2} \times P_{\text{total}} = \frac{1}{3} \text{ part}$$

**49.(D)** The van der Waals constant  $T_c$ ,  $p_c$ , and  $V_c$  are related by the expression

$$V_c = 3b; \quad T_c = a \frac{8a}{27Rb}; \text{ and } p_c = \frac{a}{27b^2}$$

**50.(D)** For gas:

$$V_{\text{gas}} = \frac{nRT}{p} \quad (1)$$

For  $H_2$ ,

$$V_{H_2} = \frac{n_{H_2} RT_{H_2}}{p_{H_2}} \quad (2)$$

Now,  $n = W / M$ . From Eqs. (1) and (2), we have

$$\frac{2.9 \times 368}{M_{\text{gas}}} = \frac{0.184 \times 290}{2} \Rightarrow M_{\text{gas}} = 40 \text{ g mol}^{-1}$$

Since volume and pressure of both the gases are the same.

## <H2> Multiple Correct Choice Type

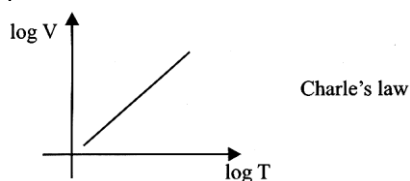
1.(A, B, C, D)

- (A) Collisions of gas molecules are perfectly elastic which means that the total energy of the molecules before and after the collision remains the same.
- (B) Heavier particles transfer more momentum because momentum depends on the mass.
- (C) Only a small number of molecules possess very high velocity.
- (D) Particles of gas move in all possible directions in straight lines. The distribution of speed remains constant at a particular temperature.

2.(A, C)  $V = (\text{const.})T = kT$

$$\log V = \log T + \log k$$

The graph based on the above equation is



3.(A, C, D) The van der Waals equation is

$$\left( p + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

At low pressure, when the sample occupies a large volume, then the intermolecular forces do not play any significant role and the gas behaves virtually perfectly ideal.

Constants  $a$  and  $b$  are called van der Waals constants and their value depends on the characteristic of a gas. Value of  $a$  is a measure of magnitude of intermolecular attractive forces within the gas and is independent of temperature and pressure.

The observed  $p$  correction term is added as

$$P_{\text{ideal}} = P_{\text{real}} + \frac{an^2}{V^2}$$

Pressure exerted by real gas is lower than the pressure exerted by the ideal gas.

4.(B, D) We know

$$\frac{r_{\text{He}}}{r_{\text{CO}}} = \left( \frac{M_{\text{CO}}}{M_{\text{He}}} \right)^{1/2} \quad (1)$$

Substituting the values in Eq. (1), we get

$$\frac{r_{\text{He}}}{r_{\text{CO}}} = \left( \frac{28}{4} \right)^{1/2} = 2.645$$

Similarly,

$$\frac{r_{\text{He}}}{r_{\text{SO}_2}} = \left( \frac{M_{\text{SO}_2}}{M_{\text{He}}} \right)^{1/2} = \left( \frac{64}{4} \right)^{1/2} = 4$$

5.(C, D) Concept based.

6.(A, C, D) When  $V_m$  approaches large value,  $p$  approaches zero, that is,  $Z = 1$ .

When  $p$  approaches large value, that is,  $Z > 1$  and increases.

van der Waals gas constant does not depend on the temperature and pressure and depends on the identity of the gas. A real gas exerts pressure less than an ideal gas due to the presence of the attractive force.

7.(C, D)  $p \propto T$  and  $V \propto T$ .

8.(B, C, D) According to kinetic gas equation;  $pV = 1/3 mNu^2$  based on kinetic gas postulates.

9.(A, D) As per kinetic theory of gases.

10.(B, C, D) Ideal gas follows all laws.

11.(A, B) Concept based.

12.(A, C) Concept based.

13.(A, B, C, D) Concept based.

14.(A, B, C) Concept based.

15.(A, B, D) Since, the diffusion of gases is directly proportional to square root of their masses, therefore, statement (C) is incorrect.

<H2> Assertion–Reasoning

1.(A) In case of H<sub>2</sub>, compressibility factor increases with the pressure. At 273 K,  $Z > 1$ , which shows that it is difficult to compress the gas as compared to ideal gas. In this case, repulsive forces dominate. The value of  $a$  is very small.

2.(A) Much before this temperature a gas liquefies, and then solidifies.

3.(A) Concept based.

4.(B) We know

$$p = \frac{nRT}{V}$$

5.(C) One mole of molecules of all gases occupies 22.4 L at 273 K and 760 mm Hg pressure (STP). Hence, molar volume of all gases at STP is 22.4 L.

6.(C)  $pV + \frac{a}{V} = RT \Rightarrow pV = RT - \frac{a}{V}$ ; As at low pressure  $V$  is large; therefore,  $V - b \sim V$

Therefore,

$$Z = \frac{pV}{RT} = \left[ 1 - \frac{a}{RTV} \right]$$

7.(A) Due to H-bonding in NH<sub>3</sub>, the intermolecular forces are stronger; hence, large value of ' $a$ ' and hence easily liquefiable.

### <H2> Comprehension Type

1.(A) The ratio of  $u_{\text{rms}}:u_{\text{avg}}:u_{\text{mp}}$  is

$$\sqrt{\frac{3RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{2RT}{M}} = \sqrt{3} : \sqrt{\frac{8}{\pi}} : \sqrt{2} = 1.732 : 1.596 : 1.414$$

Or

$$u_{\text{rms}} : u_{\text{avg}} : u_{\text{mp}} = 1 : 0.921 : 0.816$$

From their relationship, this can be worked out.

2.(A) We know that

$$\text{Kinetic energy} = \frac{3}{2} RT \Rightarrow 3RT = 4000$$

$$v_{\text{rms}} = \left( \frac{3RT}{M} \right)^{1/2} = \left( \frac{4000}{32} \right)^{1/2} = 11.18 \text{ m s}^{-1}$$

3.(D) We know

$$\frac{pV_m}{RT} = 1 + \frac{B}{V_m}$$

$$\Rightarrow \frac{1 \times V_m}{0.0821 \times 298} = 1 + \frac{6.4 \times 10^{-2}}{V_m}$$

Solving, we get a quadratic equation as

$$V_m^2 - (6.4 \times 10^{-2})(0.0821 \times 298) - (0.0821 \times 298)V_m = 0$$

Solving, we get  $V_m = 22.8 \text{ L mol}^{-1}$ . Therefore, the real volume is  $3 \times 22.8 = 68.4 \text{ L}$

**4.(A)** At Boyle's temperature,  $pV = RT$ .

**5.(A)** The Boyle's temperature is

$$T_B = \frac{a}{Rb}$$

On increase in the Boyle's temperature at constant value of  $b$ , the ease of liquefaction increases.

**6.(A)** Greater the attractive forces, higher the value of  $a$ .

**7.(D)** Check that 500 K is Boyle's temperature. So,

$$500 = \frac{a}{2 \times b} \Rightarrow \frac{a}{b} = 1 \text{ k cal atm}^{-1}$$

**8.(A)** Slope is high  $p$  region  $= \frac{b}{RT} = \frac{2.2 - 2}{1200 - 1000} = \frac{0.2}{200} = 10^{-3} \text{ atm}^{-1}$

**9.(C)** Given,  $Z = 2 = \frac{pV_m}{RT}$

Therefore,  $V_m = \frac{2RT}{p} = \frac{0.082 \times 200 \times 2 \text{ L}}{500} = 0.066 \text{ L}$

**10.(A)** At Boyle's temperature,  $Z$  versus  $p$  curve lies close to the curve for ideal gas at  $p = 0$  and rises above the ideal gas curve only very slowly with increase in pressure.

<H2> Integer Answer Type

1.(5) We know

$$\left(p + \frac{a}{V_2}\right)(V - b) = RT$$

From ideal gas equation,  $Z = \frac{pV}{RT}$

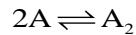
At high pressure, we have

$$Z = 1 + \frac{pb}{RT}$$
$$\text{Slope} = \frac{b}{RT} = \frac{\pi}{492.6} \Rightarrow b = \frac{\pi}{492.6} \times 0.0821 \times 300 \quad (1)$$

But  $b = \frac{4}{3} \pi r^3 \times 4N_A \quad (2)$

Solving, we get  $r = 2.5 \text{ \AA}$ , therefore,  $d = 5 \text{ \AA}$ .

2.(2) Since A and A<sub>2</sub> are two states in gaseous phase having their weight ratio 50%, that is, 1:1.



The number of moles ( $n$ ) =  $\frac{w}{M}$

$$\text{Moles of A} = \frac{96}{2} = \frac{1}{48} = 1$$

$$\text{Moles of A}_2 = \frac{96}{2} \times \frac{1}{96} = \frac{1}{2}$$

$$\text{So, total moles of A and A}_2 \text{ are} = 1 + \frac{1}{2} = \frac{3}{2}$$

Thus, from  $pV = nRT$

$$p \times 33.6 = \frac{3}{2} \times 0.0821 \times 546 = 2 \text{ atm}$$

3.(7) We know

$$\frac{p_1V_1 + p_2V_2}{V_{\text{Total}}} = p_{\text{Total}}$$

$$\Rightarrow p_{\text{Total}} = \frac{5 \times 9 + 10 \times 6}{15} = \frac{45 + 60}{15} = 7 \text{ atm}$$

4.(2) From the relation,  $p \propto \frac{1}{V}$

And  $V \propto T$

5.(5) Initial number of moles =  $\frac{4}{40} = 0.1$

Final number of moles =  $\frac{3.2}{40} = 0.08$

Since  $p$  and  $V$  are constant,  $n_1T_1 = n_2T_2$  as  $pV = nRT$

Therefore,  $nT = \text{Constant}$

$$0.1T_1 = 0.08(T_1 + 125) \Rightarrow T_1 = 500 \text{ K}$$

From  $pV = nRT$

$$p = \frac{0.1 \times 0.082 \times 500}{0.821} = 5 \text{ atm}$$

If  $p$  doubles,  $V$  becomes half, then when  $T$  doubles,  $V$  becomes its original value = 2 L.

6.(4) 1 mol of monoatomic gas,  $\gamma = 1.66$

1 mol of diatomic gas,  $\gamma = 1.40$

$$\gamma \text{ for the mixture} = \frac{1.66 + 1.40}{2} = 1.50$$

$$\gamma = \frac{C_p}{C_v} \Rightarrow 1.5 = \frac{C_p}{C_v} \text{ or } C_p = 1.5C_v$$

As  $C_p - C_v = R$ , so  $1.5C_v - C_v = 2 \text{ cal mol}^{-1} \text{ K}^{-1}$

$$C_v(1.5 - 1) = 2 \Rightarrow C_v(0.5) = 2 \Rightarrow C_v = \frac{2}{0.5} \Rightarrow C_v = 4 \text{ cal}$$

### <H2> Matrix–Match Type

1. **A** → (p, r, s); **B** → (q, r); **C** → (q); **D** → (p, r, s)

(A)  $p = 2V^2$

If the volume increases, the pressure will also increase.

(B)  $pV^2 = \text{constant}$ .

If volume increases, the temperature will decrease.

(C)  $C = C_v + 2R$

For expansion, heat will have to be supplied to the gas.

(D)  $C = C_v - 2R$

If volume increases, the temperature will decrease.

2. **A** → (r, t) **B** → (s, t) **C** → (q, r), **D** → (p, t)

$$\left(p + \frac{a}{V^2}\right)(V - b) = RT \text{ and } Z = \frac{V_m}{V_{\text{ideal}}}$$

3. **A** → (r), **B** → (p), **C** → (s,t), **D** → (q)

They are derived from kinetic gas equation.

