# HINTS & SOLUTIONS

## TOPIC : GASEOUS STATE

### EXERCISE # 1

**Section (A)**

**1.** According to Boyle's law $V \propto \frac{1}{P}$

$V = \frac{\text{Constant}}{P}$ **;** $VP$ = Constant

**2.** At sea level, because of compression by air above the proximal layer of air, pressure increases hence volume decreases *i.e.* density increases. It is Boyle's law

**3.** At constant T, $P_1V_1 = P_2V_2$ ⇒ $1\times 20 = P_2 \times 50$ **;** $P_2 = \frac{20}{50}\times 1$

**4.** $V \propto \frac{T}{P}$

So no change in volume i.e. 2 litres.

**5.** The value of universal gas constant can be expressed in different units and its value would depend only on the units of the measurement.

**6.** (atm. $K^{-1}$ $mol^{-1}$) is not a unit of R

**7.**✍ $V_1$ = 100 ml $V_2$ = 80 ml
$T_1$ = 300 K $T_2$ = ?
$P_1$ = 740 mm $P_2$ = 740 mm

Applying charles law $V \propto T$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{100}{300} = \frac{80}{T_2}$$

$T_2 = \frac{300\times 80}{100}$ = 240 K = 24 – 273 = 240 – 273ºC = – 33ºC.

**8.** $V_2 = \frac{P_1V_1T_2}{P_2T_1}$ ⇒ $P_1 = P$ ⇒ $T_1 = 273^o K$

$P_2 = \frac{3}{2}P$ **;** $T_2 = T_1 + \frac{T_1}{3} = \frac{4}{3}\times 273^o K$ ⇒ $V_2 = \frac{2P}{3P}\times\frac{4}{3}\times\frac{273}{273}\times 100cc = \frac{800}{9}cc = 88.888cc$ = 88.9 cc

**9.**✍
$$\frac{\text{Molecular weight of } O_2}{\text{Molecular weight of } SO_2} \Rightarrow \frac{M_1}{M_2} \Rightarrow \frac{32}{64} = \frac{1}{2}$$

**10.** In gases molecules are in continuous random motion thus K.E. of molecules is highest.

**11.** In 22.4 L of $H_2$ maximum number of molecules = $6.023 \times 10^{23}$

In 1 L of $H_2$ maximum number of molecules = $\frac{6.023\times 10^{23}}{22.4}$

In 15 L of $H_2$ maximum number of molecules = $\frac{6.023\times 10^{23}}{22.4}\times 15$ = $4.03 \times 10^{23}$ molecules

**12.** $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ ; $\frac{P_1\times 4}{T_1} = \frac{2P_1\times V_2}{2T_1}$ ⇒ $8 = 2\times V_2$ so $V_2 = 4dm^3$

**13.**✍ $P \propto \frac{1}{V}$

$P \times 250 = 1000 \times P'$

$P' = \frac{1}{4} P$

**14.** $V \propto T$ (at constant n and P).

**15.**✍ $6.02 \times 10^{22}$ molecules of $N_2, O_2$ each and $H_2 = \frac{6.02 \times 10^{22}}{6.02 \times 10^{23}}$ *moles* of each

Weight of mixture = weight of 0.1 *mole* $N_2 +$ weight of 0.1 *mole* $H_2$ + weight of 0.1 mole of $O_2$

$= (28 \times 0.1) + (2 \times 0.1) + (32 \times 0.1)$

**16.**✍ No. of moles of $CO_2$ present in 200 *ml* solution = molarity × Volume (in *lt.*) $= 0.1 \times \frac{200}{1000} = .02$

Volume of 0.02 mole of $CO_2 = 22.4 \times .02 lt. = 0.448 lit.$

**17.** HI for has the least volume because of greater molecular weight $V \propto \frac{1}{M}$

**18.**✍ $\frac{V_1}{V_2} = \frac{T_1}{T_2} \because V_2 = \frac{T_2}{T_1} V_1 \quad = \frac{546^o K}{273^o K} \times 0.2L = 0.4L.$

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

**21.** $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \therefore n_2 = \frac{P_2V_2T_1}{P_1V_1T_2} n_1$

at STP $n_1 =$ one mole. at $T = 273^o + 30^o = 303^o K$

$P_1$ = 1 *atm.* $P_2 =$ 1 *atm*

$V_1$ = 22.4 *lt* $V_2 = 22.4$

$T_1 = 273^o K$

$n_2 = \frac{1}{1} \times \frac{22.4}{22.4} \times \frac{273}{303} \times 1 = 0.9$ moles

**22.** $n = \frac{PV}{RT} = \frac{m}{M} \quad \Rightarrow \quad m = \frac{MPV}{RT} = \frac{34 \times 2 \times 100}{0.082 \times 293} = 282.4 gm$

**23.**✍ $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \therefore T_2 = \frac{P_2}{P_1} \frac{V_2}{V_1} T_1 \frac{m_1}{m_2} \quad = \frac{0.75}{1} \times \frac{1}{1} \times \frac{2}{1} \times 300^o K = 450^o K$

**24.** (3) At constant pressure

$V \propto nT \propto \frac{m}{M} T$

$\frac{V_1}{V_2} = \frac{m_1T_1}{m_2T_2} \therefore \frac{T_1}{T_2} = \frac{V_1}{m_1} \times \frac{m_2}{V_2} = \frac{d_2}{d_1} \quad \Rightarrow \frac{300^o K}{T_2} = \frac{0.75d}{d}$

$T_2 = \frac{300}{0.75} = 400^o K$

**27.** Here $V_1$ = 10 L, $V_2$ = 2 L

$P_1$ = 1 atm, $P_2$ = 1 atm

$T_1$ = 300 K, $T_2$ = ?

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \Rightarrow \quad \frac{1 \times 10}{300} = \frac{1 \times 2}{T_2} \quad \Rightarrow \quad T_2 = \frac{300 \times 2}{10} = 60$ K

**28.** $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \Rightarrow \frac{P_1 \times 4}{T_1} = \frac{2P_1 \times V_2}{2T_1}$

$2V_2 = 8 \Rightarrow V_2 = 4dm^3$

**29.** We know that $pV = nRT$ or $pV = \frac{w}{M}RT$

or $M = \frac{w}{V}\frac{RT}{P}$ or $M = d\frac{RT}{P}$

$d = 1.964\ g/dm^3 = 1.964 \times 10^{-3}\ g/cc$

$p = 76\ cm$

**30.** PV = nRT ideal gas equation $P = \frac{nRT}{V}$.

**Section (B)**

**2.**✍ Equal masses of CO and $N_2$ contain equal moles also on applying charles law :

$P \propto n$ (At const T and v) $\therefore$ $P_{N_2} = P_{CO}$

**3.**✍ After joining vessel

Total number of moles $= n_1 + n_2 = \frac{100\ \ V}{RT} + \frac{400\ \ V}{RT}$ $\therefore$ Applying PV = nRT

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

Volume become double $P = \frac{\left(\frac{100\ \ V}{RT} + \frac{400\ \ V}{RT}\right) RT}{2V} = \frac{500\ \ V}{2\ \ V} = 250\ mm$

**4.** $P.P_{CH4} = \frac{\frac{16}{16}}{\frac{56}{28}+\frac{44}{44}+\frac{16}{16}} \times 720$ (where P.P = Partial pressure) $= \frac{1}{4} \times 720 = 180\ mm.$

**5.**✍ Mole Fraction $= \frac{n_{H_2}}{n_{H_2} + n_{C_2H_6}} \Rightarrow \frac{\frac{w}{2}}{\frac{w}{2}+\frac{w}{30}} \Rightarrow \frac{\frac{w}{2}}{\frac{16w}{30}} \Rightarrow \frac{15}{16}$

(Here w = weight).

**6.**✍ $\frac{751-21}{760} = 0.96\ atm.$

**Section (C)**

**1.**✍ $\frac{(dn/dt)_{O_2}}{(dn/dt)_{H_2}} = \sqrt{\frac{M_{H_2}}{M_{O_2}}}$

Here, x = wt. of $H_2$

$\frac{\frac{4/32}{dt}}{\frac{x/2}{dt}} = \sqrt{\frac{2}{32}} \Rightarrow \frac{8}{32x} = \frac{1}{4}$

$x = 1\ g.$

**2.** $\dfrac{\left(\dfrac{dn}{dt}\right)_{O_2}}{\left(\dfrac{dn}{dt}\right)_{H_2}} = \sqrt{\dfrac{M_{H_2}}{M_{O_2}}} \Rightarrow \dfrac{\dfrac{x/32}{dt}}{\dfrac{2/2}{dt}} = \sqrt{\dfrac{2}{32}} \Rightarrow \dfrac{x}{32} = \dfrac{1}{4} \Rightarrow$ x = 8g. (Here x = wt. of $O_2$).

**3.**✍ $\dfrac{r_{element}}{r_{He}} = \sqrt{\dfrac{M_{He}}{M_{element}}} \Rightarrow \dfrac{1}{4} = \sqrt{\dfrac{4}{M_{element}}}$

$M_{element} = 64$.

**4.** $\dfrac{1}{6} = \sqrt{\dfrac{2}{x}}$ (Where X is molecular weight of gas)

$\dfrac{1}{36} = \dfrac{2}{x}$

x = 72

**5.** $\dfrac{\dfrac{V_x}{2}}{\dfrac{V_x}{5.65}} = \sqrt{\dfrac{32}{M_x}} \Rightarrow \left(\dfrac{5.65}{2}\right)^2 = \dfrac{32}{M_x}$

$M_x = 4$.

**6.**✍ $r \propto \dfrac{1}{\sqrt{M}}$. So, $SO_2 : O_2 : CH_4$

$\dfrac{1}{\sqrt{64}} : \dfrac{1}{\sqrt{32}} : \dfrac{1}{\sqrt{16}} \Rightarrow 1 : \sqrt{2} : \sqrt{4} \Rightarrow 1 : \sqrt{2} : 2$

**7.** $H_2$ will be filled first because of lower molecular weight

**8.** $\dfrac{r_{H_2}}{r_{O_2}} = \sqrt{\dfrac{d_{O_2}}{d_{H_2}}} \Rightarrow \dfrac{1}{r_{O_2}} = \sqrt{\dfrac{1.44}{0.09}} \Rightarrow r_{O_2} = \sqrt{\dfrac{1}{16}} \Rightarrow r_{O_2} = \dfrac{1}{4}$

**9.**✍ $\dfrac{r_{gas_1}}{r_{gas_2}} = \sqrt{\dfrac{d_{gas_2}}{d_{gas_1}}} \Rightarrow \dfrac{r_{gas_1}}{r_{gas_2}} = \sqrt{16} \Rightarrow \dfrac{r_{gas_1}}{r_{gas_2}} = 4$

**11.** Rate of diffusion $\propto \dfrac{1}{\sqrt{\text{Molecular Mass}}}$ that is why $H_2$ gas diffuse first

**12.**✍ $\dfrac{r_1}{r_2} = \dfrac{t_2}{5} = \sqrt{\dfrac{M_2}{2}}$

**Section (D)**

**1.**✍ $\dfrac{V_{rms}}{V_{av.}} = \dfrac{\sqrt{\dfrac{3RT}{M}}}{\sqrt{\dfrac{8RT}{\pi m}}} = \sqrt{\dfrac{3\pi}{8}} \Rightarrow \dfrac{V_{rms}}{V_{av.}} = \dfrac{1.086}{1}$

**3.** At constant temperature K.E. also remain constant.

**4.** K.E. $\propto$ T

(T will same)

**5.**✍ $V_{mps} : V_{av} : V_{rms} \quad \Rightarrow \quad \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}} \quad \Rightarrow \quad \sqrt{2} : \sqrt{8/\pi} : \sqrt{3}$

**6.**✍ $V_{rms} = \sqrt{\frac{3RT}{M}}$

(T will same at STP)

So, $V_{rms} \propto \frac{1}{\sqrt{M}}$ $\therefore$ Order for $V_{rms}$ are $HBr < O_2 < N_2 < H_2$

**8.** $V_{rms} = \sqrt{\frac{3RT}{M}} \Rightarrow V_{rms} = \sqrt{T}$

Given $V_1 = V$ , $T_1 = 300K$ , $V_2 = 2V$ , $T_2 = ?$

$\frac{V_1}{V_2} = \sqrt{\frac{T_1}{T_2}}$ ; $\left(\frac{V}{2V}\right)^2 = \frac{300}{T_2} \Rightarrow T_2 = 300 \times 4 = 1200K$

**9.**✍ $$\frac{\frac{3}{2} \times \frac{N}{N_A} \times R \times 150}{\frac{3}{2} \times \frac{N'}{N_A} \times R \times 300} = \frac{x}{2x}$$

$\boxed{N = N'}$ Therefore, (1) option is correct.

**10.**✍ Average K.E. for one mole = $\frac{3}{2}$ RT

Average K.E. for 14 g of $N_2$ = $\left(\frac{1}{2}\text{mole}\right) \times \frac{3}{2} \times \frac{8.314}{2}$ 400 = 2494 J.

**11.** $d \propto M \Rightarrow \frac{d_1}{d_2} = \frac{M_1}{M_2}$ ; $\frac{3d}{d} = \frac{M}{M_2}$ ; $M_2 = \frac{M}{3}$.

**12.** $KE = \frac{3}{2}RT = \frac{3}{2}PV$ $\quad \therefore P = \frac{2}{3}\frac{E}{V}$ for unit volume ($V$ = 1) $\Rightarrow$ $P = \frac{2}{3}K.E$

**13.**✍ $\frac{E_{He}}{E_{H_2}} = \sqrt{\frac{T_{He}}{T_{H_2}}}$ so energies will be same for $He \,\&\, H_2$ at same temperature.

**14.** $\frac{U_{O_3}}{U_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{O_3}}} = \sqrt{\frac{32}{48}} = \sqrt{\frac{2}{3}}$

**Section (E)**

**1.**✍ When we increase the value of 'a' i.e. easily liquified.

**2.** The pressure of real gases is less than that of ideal gas because of intermolecular attractions.

**3.** At lower tempertures, all gases show negative deviations.

**6.**✍ (4) $PV = nRT$ is a ideal gas equation it is allowed when the temperature is high and pressure is low.

**7.** b = L $mol^{-1}$

**8.**✍ For the non-zero values of force of attraction b/w gas molecules b = 0.
Vander Wall's equation

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT \quad \Rightarrow \quad \left(P + \frac{n^2a}{V^2}\right)(V - 0) = nRT \quad \Rightarrow \quad PV = nRT - \frac{n^2a}{V}$$

**9.**✍ At low pressure,

$$\left[P+\frac{a}{V^2}\right]V = RT \quad \Rightarrow \quad PV + \frac{a}{V} = RT \quad \Rightarrow \quad \frac{PV}{RT} + \frac{a}{VRT} = 1 \quad \Rightarrow \quad Z = 1 - \frac{a}{VRT}$$

**10.**✍ Gases deviate from ideal behaviour because of intermolecular forces of attraction.

**11.**✍ A real gas approaches the behaviour of ideal gas when the pressure is low and the temperature is high.

**12.** It has been experimentally found that above a particular temperature known as critical temperature ($T_c$), the liquification of a gas does not take place at all, howsoever high pressure may be applied. It is because above critical temperature of a gas, the kinetic energy of the gas molecules is sufficient enough to over-come attractive forces and keep the molecules apart.

**14.** At Boyle temperature real gas is changed into ideal gas

**15.** At high temperature and low pressure, Vander Waal's equation is reduced to ideal gas equation.
$PV = nRT$
$PV = RT$ (For 1 mole of gas)

# EXERCISE # 2

**1.** $\frac{d_1}{d_2} = \frac{1}{2}$, $\frac{T_1}{T_2} = \frac{2}{1} \because \frac{P_1}{P_2} = \frac{V_2}{V_1} \times \frac{T_1}{T_2} = \frac{T_1.d_1}{T_2.d_2}$ $\Rightarrow$ $\frac{P_1}{P_2} = \frac{2}{1}.\frac{1}{2} = \frac{1}{1}$

**3.** $d_a = 2d_b$ ; $2M_a = M_b$

$PV = nRT = \frac{m}{M}RT$ ; $P = \frac{m}{V}.\frac{RT}{M} = \frac{dRT}{M}$ $\Rightarrow$ $\frac{P_a}{P_b} = \frac{d_a}{d_b}\frac{M_b}{M_a} = \frac{2d_b}{d_b} \times \frac{2M_a}{M_a} = 4$

**4.**✍ $V_2 = \frac{P_1}{P_2}\frac{T_2}{T_1}.V_1 = \frac{1}{0.5} \times \frac{250}{300} \times 12000 lit. = 20000 lit.$

**5.** $PV = nRT$

$P = \frac{n}{V}RT \because \frac{n}{V} = C \Rightarrow P = CRT$ $\Rightarrow$ $T = \frac{P}{CR} = \frac{1}{1 \times .821} = 12^o K$

**6.**✍ M. wt. = V.d. × 2 = 70 × 2 = 140 ⇒ $x = \frac{m.wt.}{wt.of[CO]} = \frac{140}{[12+16]} = 5$

**7.**✍ At same temperature and pressure, equal volumes have equal number of molecules. If 1lit. of oxygen consists N molecules then at same temperature and pressure 1 lit of $SO_2$ will consists N molecules. So 2 lit. of $SO_2$ will contain 2N molecules.

**8.** $n(H_2) = \frac{2}{2} = 1$, $n(N_2) = \frac{14}{28} = 0.5$ $\Rightarrow$ $n(O_2) = \frac{16}{32} = 0.5$ $\Rightarrow$ $p(H_2) = \frac{1}{1+0.5+0.5}P = \frac{1}{2}P$

**9.** No. of moles of lighter gas $= \frac{m}{4}$ $\Rightarrow$ No. of moles of heavier gas $= \frac{m}{40}$

Total no. of moles $= \frac{m}{4} + \frac{m}{40} = \frac{11m}{40}$ $\Rightarrow$ Mole fraction of lighter gas $= \frac{\frac{m}{4}}{\frac{11m}{40}} = \frac{10}{11}$

Partial pressure due to lighter gas $= P_o \times \frac{10}{11} = 1.1 \times \frac{10}{11} = 1 atm.$

**10.**✍ $r \propto \sqrt{\frac{1}{M}}$.

**11.** Because both NO and $C_2H_6$ have same molecular weights $\lfloor M_{NO} = M_{C_2H_6} = 30 \rfloor$ and rate of diffusion $\propto$ molecular weight.

**12.**✍ *P.V* = constant at constant temperature. As temperature changes, the value of constant also changes.

**13.** Root mean square speed $= \left[\frac{n_1c_1^2 + n_2c_2^2 + n_3c_3^2 + ....}{n_1 + n_2 + n_3 + ......}\right]^{1/2}$

**14.** If 40 *ml* $O_2$ will diffuse in *t* min. then. $r_{O_2} = \frac{40}{t}$

$r_H = \frac{50}{20} \Rightarrow r_O = r_H \, 2\sqrt{\frac{M_{H_2}}{M_{O_2}}} = \frac{50}{20}\sqrt{\frac{2}{32}} = \frac{50}{20}.\frac{1}{4}$

$\Rightarrow \quad \frac{40}{t} = \frac{50}{80} \therefore t = \frac{40 \times 80}{50} = 60$ *min.*

**16.** $\frac{u\ (SO_2)}{u\ (O_2)} = \sqrt{\frac{T\ (SO_2)}{M\ (SO_2)} \times \frac{M\ (O_2)}{T\ (O_2)}}$ i.e $1 = \sqrt{\frac{T\ (SO_2)}{64} \times \frac{32}{303}}$ or $T\ (SO_2) = 606$ K

**17.**✍ At constant *V* of a definite mass

$\frac{P_1}{T_1} = \frac{P_2}{T_2} \therefore \frac{P_1}{P_2} = \frac{300}{600} = \frac{1}{2}$ *i.e.* pressure increases and on increasing temperature energy of molecules increases so the rate of collisions also increases and number of moles remains constant because there is neither addition nor removal of gas in the occurring.

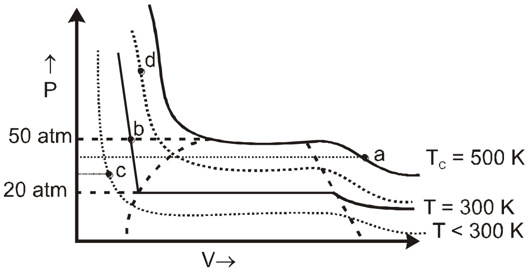
**18.**✍ $\frac{U_1}{U_2} = \sqrt{\frac{T_1}{T_2}} \qquad \therefore \frac{T_1}{T_2} = \left(\frac{5 \times 10^4}{10 \times 10^4}\right)^2 = \frac{1}{4}$

**19.** $\frac{U_1}{U_2} = \sqrt{\frac{n_1T_1}{n_2T_2}} = \sqrt{\frac{n \times T}{2n \times 2T}} = \sqrt{\frac{1}{4}} = \frac{1}{2}$

$U_2 = 2U_1 = 2U$

**23.** At low temperature and high pressure.

**24.**✍



(1) at T = 500 K, P = 40 atm corresponds to 'a' substance - gas
(2) at T = 300 K, P = 50 atm corresponds to 'b' substance - liquid
(3) at T < 300 K, P > 20 atm corresponds to 'c' substance - liquid
(4) at T < 500 K, P > 50 atm corresponds to 'd' substance - liquid
So, Answer **(4)**

**25.**✍ $T_c = \frac{8a}{27\ Rb}$. Thus $T_c \propto \frac{a}{b}$

**26.** Absence of inter molecular attraction ideal gas cannot be liquefied at any volume of *P* and *T*.

**30.** Compressibility of an ideal gas (Z) = $\frac{P \times V}{nRT} \Rightarrow$ For a mole of ideal gas at STP, (Z) = $\frac{P \times 22.4}{RT}$

For other gases Z < 1. It is equal to $\frac{P \times V_m}{RT} \Rightarrow$ Hence $V_m$ < 22.4 litre.

**31.** Given that : $\sqrt{\frac{3RT_{H_2}}{2}} = \sqrt{7}\sqrt{\frac{3RT_{N_2}}{28}}$ (because $U_{rms} = \sqrt{\frac{3RT}{M}}$ )

So $\frac{3RT_{H_2}}{2} = 7 \times \frac{3 \times RT_{N_2}}{28}$ $\therefore$ $T_{N_2}$ = or $T_{N_2} > T_{H_2}$ .

**32.** Mass of steam = 1000 × 0.0006 = 0.6 gm
Mass of water ($\ell$) = 0.6 gm
Volume of liquid water = 0.6 $cm^3$

**33.** Correction factor for attractive force in to the real gas is given by $\frac{an^2}{V^2}$ .

# EXERCISE # 3

## PART - I

**1.** As the temperature is constant. Boyle's law is applicable :

$$p_1V_1 = p_2V_2$$

$V_1$ = 380 mL, $p_1$ = 720 mm, $V_2$ = ?, $p_2$ = 760 mm

$$730 \times 380 = 760 \times V_2$$

$$V_2 = \frac{730 \times 380}{760} = 365 \text{ mL}$$

**2.** Mass of gas can be determined by weighing a container in which it is enclosed as follows
Mass of the gas = mass of the cylinder including gas – mass of empty cylinder.

**3.** KE = $\frac{3}{2}$ RT (for one mole of gas)
The temperature is constant and kinetic energy is independent of molecular weights. So,

$$\overline{KE}_{CO} = \overline{KE}_{N_2}$$

**4.** $r_x / r_{N_2} = \sqrt{M_{N_2}/M_x} = \sqrt{28/56} = \sqrt{1/2} = 1/\sqrt{2}$. or $r_{N_2} = r_x \times \sqrt{2} = 56 \times \sqrt{2}$ = 79.19 ml $s^{-1}$.

**5.** The beans are cooked earliker in pressure cooker because boiling point increases with increasing pressure.

**6.** At higher temperature and low pressure real gas acts as an ideal gas and obey pV = nRT relation.

**7.** Surface tension of $H_2O$ is maxiumum due to maxiumum hydrogen bonding in comparsion to $C_6H_6$, $CH_3OH$, $C_2H_5OH$. The order of H-bonding is :
$H_2O > CH_3OH > C_2H_5OH$ (Benzene does not form H-bond).

**8.** KE = $\frac{3}{2}$ RT (for one mole of a gas) As, the kinetic energy of a gaseous molecule depends only on temperature, thus at constant temperature, the kinetic energy of the molecule remains the same.

**9.** $\frac{V_A}{t_A} \Big/ \frac{V_B}{t_B} = \sqrt{\frac{M_B}{M_A}}$

**10.** $r \propto \sqrt{\frac{1}{M}} \quad \Rightarrow \quad \frac{r_2}{r_1} = \sqrt{\frac{M_2}{M_1}} \quad \Rightarrow \quad \frac{\frac{V_g}{3t}}{\frac{V_{He}}{t}} = \sqrt{\frac{4}{M}} \quad \Rightarrow \quad \frac{1}{9} = \frac{4}{M}$ M = 36 g/mole

**11.** Molar mass ↑, 'a' increases
size of molecule ↑, 'b' increase

| | b (L/mol) | | a (bar.L²/mol²) |
|---|---|---|---|
| $H_2$ | → 0.02661 | $CH_4$ | → 2.283 |
| He | → 0.0237 | $O_2$ | → 1.378 |
| $O_2$ | → 0.03183 | $H_2$ | → 0.2476 |
| $CO_2$ | → 0.04267 | | |

**12.** Most polar molecule amongs the following.

**14.** Equal moles are given so partial pressure is equal (let = x)

$$\frac{r_{O_2}}{r_{H_2}} = \sqrt{\frac{M_{H_2}}{M_{O_2}}}$$

$$\frac{n_{O_2}/t}{\frac{x}{2}/t} = \sqrt{\frac{2}{32}} = \frac{1}{4}$$

$$\frac{n_{O_2}/t}{\frac{x}{2}/t} = \frac{1}{4}$$

$$\frac{n_{O_2}}{x} = \frac{1}{8}$$

Fraction of oxygen escaped = $\frac{1}{8}$

**15.** Correction factor is corresponding to force of attraction between the gas molecule.

**16.** For $NH_3$ a is high so it is most easily liquefied.

**17.** $V_i = V$,
$V_r = V - 0.2\ V \qquad = 0.8\ V$

if value of Z < 1 then attractive forces are dominant. $\left(Z = \frac{V_r}{V_i} = 0.8\right)$

**18.** A gas expands against a constant external pressure is irreversible process.
The work done in irreversible process
$= -P_{ext}\Delta V = -P_{ext}(V_2 - V_1) = -2\ (0.25 - 0.1) = -2 \times 0.15$ bar – L/L-bar = –.30 × 100J = –30 J and work done by the gas is +30 J

**19.** The total pressure = $P_{H_2O} + P_{air}$

$1.2 = P_{H_2O} + P_{air}$

$P_{H_2O}$ = mole fraction of $H_2O$ + $P_{total}$

$P_{H_2O} = 0.02 \times 1.2 = 0.024$

$P_T = P_{H_2O} + P_{air}$

$1.2 = 0.024 + P_{air}$

$P_{air} = 1.176$ atm

**20.** Given : $W_{H_2O} = 1.8$ g

$P = 1$ bar

$T = 374 + 273 \quad = 647$ K

from ideal gas :

$$PV = nRT$$

$$v = \frac{nRT}{P} \qquad = \frac{{}^{1.8}\!/_{18} \times 0.083 \times 647}{1}$$

$$v = 5.37 \text{ L}$$

## PART - II

**1.** The correct order of pressure is

$p_1 > p_3 > p_2$ (For same volume & moles p T ).

**2.** Average kinetic energy $E = \frac{3}{2}$ RT

Root mean square velocity $U_{rms} = \sqrt{\frac{3\,RT}{M}}$

From eq. (i) $RT = \frac{2\,E}{3}$

On putting the value of RT in the expression of $U_{rms}$, we get $U_{rms} = \sqrt{\frac{3 \times 2\,E}{3M}} = \sqrt{\frac{2\,E}{M}}$

**3.** In vander Waal's equation, a signifies the intermolecular force of attraction.

**4.** For effusion of same volume, $\frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}} \Rightarrow \frac{t_1}{\sqrt{M_1}} = \frac{t_2}{\sqrt{M_2}}$

This is clearly seen from the options that the ratio of $\frac{t}{\sqrt{M}}$ is same for $H_2$ and $O_2$ $\left(\frac{5}{\sqrt{2}} = \frac{20}{\sqrt{32}} = \frac{5}{\sqrt{2}}\right)$.

**5.** Rate $\propto \frac{1}{\sqrt{M}}$ The smaller the value of M the more is rate of diffusion

**9.** From the given equation

$$\left[p + \frac{a}{V^2}\right]V = RT \quad \Rightarrow \quad pV + \frac{a}{V} = RT \quad \Rightarrow \frac{pV}{RT} = \frac{a}{RTV} = 1 \Rightarrow \frac{pV}{RT} \text{ Z (compressibility factor)} = \left(1 - \frac{a}{RTV}\right)$$

**10.** Critical temperature of a gas may be defined as that temperature above which it cannot be liquefied however high pressure may be applied on the gas. Hence, $CO_2$ cannot be liquefied above its critical temperature. According to Boyle's law $p \propto \frac{1}{V}$ or $V \propto \frac{1}{P}$ (at constant T and n)

**12.** $\frac{n_1}{p_1} = \frac{n_2}{p_2} \quad \Rightarrow \quad \frac{\frac{1}{M_A}}{4} = \frac{\frac{1}{M_A} + \frac{2}{M_B}}{6} \quad \Rightarrow \quad \frac{3}{M_A} = \frac{2}{M_A} + \frac{4}{M_B} \quad \Rightarrow \quad \frac{1}{M_A} = \frac{4}{M_B} \quad \Rightarrow \quad M_B = 4M_A$

**13.** Strong hydrogen bonding intermolecular forces results in greater surface tension of water.

**14.** $\left(P + \frac{an^2}{V^2}\right)(V) = nRT \quad \Rightarrow \quad \left(P + \frac{1.4}{(0.1)^2}\right)(0.1) = 1 \times 0.082 \times 3600$

$(P + 140) \times 0.1 = 4.6 \quad \Rightarrow \quad 0.1\,P + 14 = 246 \quad \Rightarrow \quad 0.1\,P = 10.6 \quad \Rightarrow \quad P = 106$ atm

**15.** $\frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \Rightarrow \quad \frac{10}{300} = \frac{15}{T_2} = T_2 = 450$ K

## PART - III

**1.** $R = 8.3\ J\ mol^{-1}\ K^{-1} = 2\ cal\ mol^{-1}\ K^{-1} = 8.314 \times 10^7\ erg\ mol^{-1}\ K^{-1} = 0.082$ litre atm $K^{-1}\ mol^{-1}$

**2.** Various gas law can be easily derived from kinetic equation obtained as a result of kinetic theory of gases

**3.** $\frac{n}{V} = \frac{P}{RT}$

**4.** It is one of the assumption of kinetic theory of gas

**5.** $\underset{2\ mol}{2BCl_3} + \underset{3\ mol}{3H_2} \longrightarrow \underset{2\ mol}{2B} + 6HCl$

21.6 g B = 2 mol B = 3 mol $H_2$

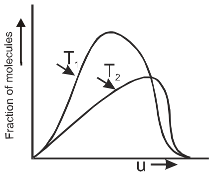
$PV = nRT \quad \therefore \quad V = \frac{nRT}{P} = \frac{3 \times 0.0821 \times 273}{1} = 67.2$ L

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

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

**7.** $b = 4N_A \times v$ i.e., total volume occupied by molecules of one mole of gas in motion.

**8.** Distribution of molecules (N) with velocity ($\mu$) at two temperature $T_1$ and $T_2$ ($T_2 > T_1$) is show below :



At both temperatures, distribution of molecules with increases in velocity first increases, reaches a maximum value and then decreases.

**9.** C* = most probable speed = $\sqrt{\frac{2RT}{M}}$ $\Rightarrow$ $\bar{C}$ = average speed = $\sqrt{\frac{8RT}{\pi M}}$

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

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

**Note :** As no option correspond to mean square speed, it is understood as misprint. It should be root means square speed. So, ans is (3)

**10.** In low pressure region, vanderwaal equation becomes $Z = 1 - \frac{a}{RTV_m}$

**11.** Initial moles = final moles

$$\frac{P_i \times V}{RT_1} + \frac{P_i \times V}{RT_1} = \frac{P_f \times V}{RT_2} + \frac{P_f \times V}{RT_1} \Rightarrow \frac{P_i}{T_1} + \frac{P_i}{T_1} = \frac{P_f}{T_2} + \frac{P_f}{T_1} \Rightarrow \frac{2P_i}{T_1} = P_f\left[\frac{1}{T_2} + \frac{1}{T_1}\right]$$

$$\frac{2P_i}{T_1} = P_f\left[\frac{T_1 + T_2}{T_1T_2}\right] \Rightarrow P_f = 2P_i \times \left(\frac{T_2}{T_1 + T_2}\right)$$

**12.** PV = nRT

200 × 10 = (0.5 + x) R × 1000

$0.5 + x = \frac{2}{R}$ $\Rightarrow$ $x = \frac{4 - R}{2R}$

**13.** $n_1T_1 = n_2T_2$

$\Rightarrow$ $n \times 300 = \left(n - \frac{2n}{5}\right)T_2$

$\Rightarrow$ $300 = \frac{3}{5}T_2$

$\Rightarrow$ $T_2 = 500$ K

**14.** PV = ZnRT

$\Rightarrow$ $P = \frac{ZnRT}{V}$ at constant T and mol $P \propto \frac{Z}{V}$

$$\frac{P_A}{P_B} = \frac{Z_A}{Z_B} \times \frac{V_B}{V_A} = \left(\frac{3}{1}\right) \times \left(\frac{1}{2}\right) = \frac{3}{2} \quad \therefore 2P_A = 3P_B$$