5 States of Matter

Section-A : JEE Advanced/ IIT-JEE

A **1.** 900 **2.** R **3.** inversely, time **4.** 0.25 **5.** 1 : 16

B **1.** F **2.** F **3.** F **4.** F

C **1.** (a) **2.** (b) **3.** (a) **4.** (b) **5.** (b) **6.** (d) **7.** (a)
8. (c) **9.** (b) **10.** (c) **11.** (b) **12.** (a) **13.** (d) **14.** (a)
15. (a) **16.** (c) **17.** (c) **18.** (b) **19.** (b) **20.** (b) **21.** (c)
22. (c) **23.** (b) **24.** (c) **25.** (d) **26.** (c) **27.** (b) **28.** (c)
29. (c) **30.** (b) **31.** (a) **32.** (b) **33.** (b) **34.** (c) **35.** (d)

D **1.** (b) **2.** (a, b) **3.** (d) **4.** (c) **5.** (b) **6.** (a, c) **7.** (a, b, c, d) **8.** (c)

E **1.** 3.42 g/litre **2.** 39.13 **3.** 85.2 cm **4.** 3 : 2
5. 2.197 atm **6.** 6.21×10^{-21} joules/molecule **7.** 2.7×10^{10} molecules **8.** 1 : 3
9. 3.9×10^{4} cm sec^{-1} **10.** 10 **11.** 1682.5 K, 2143.4 K **12.** 5.23 L
13. 1033 **14.** 0.0043, 0.221 atm **15.** 494.16 m/s, 274.13K, 405.2 m/sec **16.** 8 : 1
17. 2.463 m^3, 1.478 atm **18.** 33.75% and 66.25% **19.** 13.7%
20. CH_4-20%, CO-50%, He-30% **21.** 123 **22.** 4.53 g/litre **23.** 6.46 atm $\text{litre}^2\ \text{mol}^{-2}$
24. 0.4 **25.** XeF_6 **26.** 100 K, 0.82 litre **27.** 0.9922 atm **28.** 1.253 atm $L^2\ \text{mol}^{-2}$
29. (a) (i) 18.09, (ii) 41.025L, (iii) 1.224, (b) 2.07×10^{-20} **30.** 434 ms^{-1} **31.** RT

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

G **1.** (c) **2.** (d)
H **1.** (a) **2.** (a)
I **1.** 4 **2.** 7 **3.** 4

Section-B : JEE Main/ AIEEE

1. (c) **2.** (c) **3.** (d) **4.** (b) **5.** (a) **6.** (a) **7.** (c)
8. (d) **9.** (d) **10.** (c) **11.** (c) **12.** (b) **13.** (b) **14.** (b) **15.** (a)

Section-A JEE Advanced/ IIT-JEE

A. Fill in the Blanks

1. **900;** Energy of one mole of an ideal monoatomic gas $= \frac{3}{2}$ RT

$R = 8.314\ J\ K^{-1}\ \text{mole}^{-1} = 1.99\ \text{cal}\ K^{-1}\ \text{mole}^{-1}$
$T = 27°C = 273 + 27 = 300\ K$

$\Rightarrow\ E = \frac{3}{2} \times 1.99 \times 300 = 900$ cal

2. **R.** $[\because C_p - C_v = R]$

3. **Inversely, time;** $\left[\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}\right]$.

4. **0.25;** $\left(\because 5.6\,l \equiv \frac{1}{4} \text{ mole of an ideal gas}\right)$

Using ideal gas equation

$$PV = nRT = \frac{1}{4}RT = 0.25\ RT$$

5. **1 : 16;** $\frac{K.E._{O_2}}{K.E._{H_2}} = \frac{(3/2)n_{O_2}RT}{(3/2)n_{H_2}RT} = \frac{n_{O_2}}{n_{H_2}} = \frac{8/32}{8/2} = \frac{1}{16}$

B. True / False

1. **False :** K.E. = $\frac{3}{2}$ KT, and cannot be zero at 0°C or 273 K.
2. **False :** The pressure exerted by the gas is the result of collisions of the molecules on the walls of the container.
3. **False :** The constant '*a*' reflects the intermolecular attraction between gaseous molecules. The constant '*b*' reflects the actual volume of one mole of gaseous molecules.
4. **False :** An ideal gas cannot be liquefied as there exists no intermolecular attraction between molecules.

C. MCQs with One Correct Answer

1. **(a)** **TIPS/Formulae :**

$$\text{Mole fraction of } O_2 = \frac{\text{Moles of } O_2}{\text{Total moles}}$$

Partial pressure of O_2 = Mole fraction of O_2

$$\text{Mole fraction of } O_2 = \frac{\frac{W}{32}}{\frac{W}{16}+\frac{W}{32}} = \frac{1}{3}$$

2. **(b)** The temperature at which a real gas behaves like an ideal gas is called Boyle's temperature or Boyle's point.

3. **(a)** $U_{rms} : U_{av} = \sqrt{\frac{3RT}{M}} : \sqrt{\frac{8RT}{\pi M}}$ *or* $\sqrt{3} : \sqrt{\frac{8}{\pi}} = 1.086 : 1$

4. **(b)** Average kinetic energy depends only on temperature and does not depend upon the nature of the gas. ($\because$ K.E. = 3/2 KT)

5. **(b)** Pressure exerted by hydrogen will be proportional to its mole fraction.

$$\text{Mole fraction of } H_2 = \frac{\frac{w}{2}}{\frac{w}{16}+\frac{w}{2}} = \frac{8}{9}$$

6. **(d)** $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$

7. **(a)** $U_{av} = \sqrt{\frac{8RT}{\pi M}}$; $\left\{\begin{matrix} T_1 = 27+273 = 300\text{ K} \\ T_2 = 927+273 = 1200\text{ K} \end{matrix}\right\}$

$$\frac{U_{av_1}}{U_{av_2}} = \sqrt{\frac{T_1}{T_2}} \text{ or } \frac{0.3}{U_{av_2}} = \sqrt{\frac{300}{1200}} \text{ or } \frac{0.3}{U_{av_2}} = \sqrt{\frac{1}{4}}$$

or U_{av_2} **= 0.6 m/sec.**

8. **(c)** $\left(P+\frac{a}{V^2}\right)(V-b) = RT$; Here $\left(P+\frac{a}{V^2}\right)$ represents the intermolecular forces.

9. **(b)** **TIPS/Formulae :**

$$\text{Rate of diffusion} \propto \sqrt{\frac{1}{\text{Molecular mass}}}$$

$\because$ Molecular mass of HCl > molecular mass of NH_3

$\therefore$ HCl diffuses at slower rate and white ammonium chloride is first formed near HCl bottle.

10. **(c)** '*a*' is directly related to forces of attraction. Hence greater the value of '*a*', more easily the gas is liquified.

11. **(b)** **TIPS/Formulae :**

$$d = \frac{PM}{RT}$$

It means density of gas is directly proportional to pressure and inversely proportional to temperature. Density of neon will be maximum at highest pressure and lowest temperature.

$\therefore$ (b) is correct answer.

12. **(a)** $\frac{r_{CH_4}}{r_x} = 2 = \sqrt{\frac{M_x}{M_{CH_4}}} = \sqrt{\frac{M_x}{16}}$, or $M_x = 64$

13. **(d)** Pressure exerted by the gas, $P = \frac{1}{3}\frac{mnu^2}{V}$ ...(1)

Here, u = root mean square velocity
m = mass of a molecule, n = No. of molecules of the gas
Hence (a) & (b) are clearly wrong.

Again $u^2 = \frac{3RT}{M}$ [explained from (1)]

Here, M = Molecular wt. of the gas;
Hence (c) is wrong

Further, Average K.E. = $\frac{3}{2}$ KT; Hence (d) is true.

14. **(a)** Due to increase in the temperature, the kinetic energy of the gas molecules increases resulting in an increase in average molecular speed. The molecules are bombarded to the walls of the container with a greater velocity resulting in an increase in pressure.

15. **(a)** The mean free path, $\lambda = \frac{1}{\sqrt{2}\pi a^2 N}$

or $\lambda \alpha \frac{1}{a^2}$, where a = molecular diameter

$\therefore$ Smaller the molecular diameter, longer the mean free path. Hence H_2 is the answer.

16. **(c)** **NOTE :** The value of '*a*' indicates the magnitude of attractive forces between gas molecules.
Value of '*a*' $\propto$ size of molecule.
$\therefore$ inert gas will have minimum value of '*a*' followed by H_2O, C_6H_6 and $C_6H_5CH_3$

17. **(c)** The expression of root mean square speed is

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

Hence,

$$\frac{U_{rms}(H_2)}{U_{rms}(O_2)} = \left[\frac{3R(50K)/(2\,g\,mol^{-1})}{3R(800K)/(32\,g\,mol^{-1})}\right]^{1/2} = 1$$

18. (b) Under identical conditions, $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

As rate of diffusion is also inversely proportional to time, we will have, $\frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$

(a) Thus, For He, $t_2 = \sqrt{\frac{4}{2}}(5s) = 5\sqrt{2}s \neq 10s$;

(b) For O_2, $t_2 = \sqrt{\frac{32}{2}}(5s) = 20s$

(c) For CO, $t_2 = \sqrt{\frac{28}{2}}(5s) \neq 25s$;

(d) For CO_2, $t_2 = \sqrt{\frac{44}{2}}(5s) \neq 55s$

19. (b)

	$N_2O_4(g) \rightleftharpoons$	$2NO_2(g)$
At start	100/92 mol = 1.08 mol	0
At equilibrium	80/92 mol = 0.86 mol	20/46 mol =0.43 mol

According to ideal gas equation, at two conditions

At 300 K; $P_0V = n_0RT_0$

$1 \times V = 1.08 \times R \times 300$ ...(i)

At 600 K; $P_1V = n_1RT_1$

$P_1 \times V = (0.86 + 0.43) \times R \times 600$ ...(ii)

Divide (ii) by (i),

$$\frac{P_1}{1} = \frac{1.29 \times 600}{1.08 \times 300};$$

$$P_1 = \frac{1.29 \times 2}{1.08} = 2.38 \text{ atm.} \simeq 2.4 \text{ atm.}$$

20. (b) The compressibility factor of a gas is defined as

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

For an ideal gas, $pV_m = RT$. Hence $Z = 1$

21. (c) For an ideal-gas behaviour, the molecules of a gas should be far apart. The factors favouring this condition are high temperature and low pressure.

22. (c) **TIPS/Formulae :**

$$U_{rms} = \sqrt{\frac{3RT}{M}} \Rightarrow \sqrt{\frac{3RT_{H_2}}{2}} = \sqrt{7}\sqrt{\frac{3RT_{N_2}}{28}};$$

$$\therefore\ T_{N_2} = 2T_{H_2} \text{ or } T_{N_2} > T_{H_2}$$

23. (b) $(PV)_{Observed}/(PV)_{Ideal} < 1$

$\Rightarrow V_{obs} < V_{ideal}$, $V_{obs} < 22.4$ litre.

24. (c) Mass of 1 L of vapour = volume × density

$= 1000 \times 0.0006 = 0.6$ g

$$V \text{ of liquid water} = \frac{\text{mass}}{\text{density}} = \frac{0.6}{1} = 0.6 \text{ cm}^3$$

25. (d) $U_{RMS} = \sqrt{\frac{3RT}{M}}$ Using ideal gas equation,

$PV = nRT = \frac{w}{M}RT$; $\frac{RT}{M} = \frac{PV}{w} = \frac{p}{d}$ where d is the density of the gas

$\therefore\ U_{RMS} = \sqrt{\frac{3P}{d}}$ at constant pressure, $U_{RMS} \propto \frac{1}{\sqrt{d}}$

26. (c) **TIPS/Formulae :**

Find the volume by either

$V = RT/P$ ($PV = RT$) or $P_1V_1 = P_2V_2$ and and match it with the values given in graph to find correct answer.

Volume of 1 mole of an ideal gas at 273 K and 1 atm is 22.4 L and that at 373 K and 1 atm pressure is calculated as ;

$$V = \frac{RT}{P} = \frac{0.082 \times 373}{1} = 30.58L \simeq 30.6 \text{ L}$$

27. (b) Upon increase of temperature the internal energy of water or any system increases resulting in decrease in intermolecular force and hence decrease in surface tension. Surface tension decreases with increase in mobility due to increase in temperature.

28. (c) For positive deviation: $PV = nRT + nPb$

$$\Rightarrow \frac{PV}{nRT} = 1 + \frac{Pb}{RT}$$

Thus, the factor nPb is responsible for increasing the PV value, above ideal value. 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.

29. (c) Average KE $= E = \frac{1}{2}Mu_{rms}^2$

$$\therefore\ u_{rms}^2 = \frac{2E}{M} \text{ or } u_{rms} = \sqrt{\frac{2E}{M}}$$

30. (b) **TIPS/Formulae :**

Use Grahams' law of diffusion

$$\frac{r_{He}}{r_{CH_4}} = \sqrt{\frac{M_{CH_4}}{M_{He}}} = \sqrt{\frac{16}{4}} = 2$$

31. **(a)** $TV^{\gamma-1}$ = Constant (∵ change is adiabatic)

$$T_1V_1^{\gamma-1} = T_2V_2^{\gamma-1}$$

For monoatomic gas $\gamma = \frac{5}{3}$

$$\therefore\ T_1V_1^{2/3} = T_2V_2^{2/3} \Rightarrow T(1)^{2/3} = T_2(2)^{2/3}$$

$$T_2 = \frac{T}{2^{(2/3)}}$$

32. **(b)** In general, the molar heat capacity for any process is given by

$$C = C_v + \frac{R}{1-n},\ \text{when } PV^n = \text{constant}$$

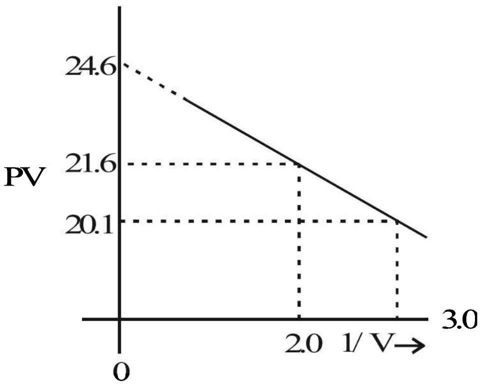
Here $\frac{P}{V} = 1$, i.e. PV^{-1} = constant

For monoatomic gas, $C_v = \frac{3}{2}R$

$$\therefore C = \frac{3}{2}R + \frac{R}{1-(-1)} = \frac{3}{2}R + \frac{R}{2} = \frac{4R}{2} = 2R.$$

33. **(b)** Correction factor for attractive force for n moles of real gas is given by the term mentioned in (b).

34. **(c)**



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

$$PV + a/V = RT;\ PV = RT - a(V)$$

$$y = RT - a(x)$$

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

35. **(d)**

- A solution of CH_3OH and water shows positive deviation from Raoult's law, it means by adding CH_3OH intermolecular force of attraction decreases and hence surface tension decreases.
- By adding KCl in water, intermolecular force of attraction bit increases, so surface tension increases by small value.
- By adding surfactant like $CH_3(CH_2)_{11}OSO_3^-Na^+$, surface tension decreases rapidly and after forming micelle it slightly increases.

D. MCQs with One or More Than One Correct

1. **(b)** No work is required to tear apart the molecules due to the absence of attractive forces in an ideal gas.

2. **(a,b)** At constant temp., when gas expands the K.E. of the molecules remains the same, but the pressure decreases. $\left(\because P \propto \frac{1}{V}\right)$

3. **(d)** Pressure exerted by H_2 is proportional to its mole fraction.

$$\text{Mole fraction of } H_2 = \frac{\frac{W}{2}}{\frac{W}{2}+\frac{W}{30}} = \frac{30}{32} = \frac{15}{16}$$

4. **(c)** According to Graham's law of diffusion for two gases undergoing diffusion at different pressures through same hole

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}} \times \frac{P_A}{P_B}$$

$$\left(r \propto P \times \sqrt{\frac{1}{M}} \quad \text{At constant temperature}\right)$$

5. **(b)** For gas A, $a = 0, Z = 1 + \frac{Pb}{RT}$ implies Z varies linearly with pressure.

For gas B, $b = 0, Z = 1 - \frac{a}{VRT}$. Hence, Z does not vary linearly with pressure.

Given the intersection data for gas C, it is possible to find the values of 'a' and 'b'. All vander Waal gases, like gas C, give positive slope at high pressures.

6. **(a,c)** Vander Waals equation is

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT \quad \text{[For n moles of a gas)}$$

a, b are vander Waals constants

The ideal gas equation is $PV = nRT$ [For n moles of a gas]

where P is pressure excerted by ideal gas and V is volume occupied by ideal gas.

In vander Waals equation the term $\left(P + \frac{n^2a}{V^2}\right)$ represents the pressure exerted by the gas and $(V - nb)$ the volume occupied by the gas. At low pressure, when the gas occoupies large volume the intermolecular distance between gaseous moleculas is quite large and in such case there is no significant role played by intermolecular forces and thus the gas behaves like an ideal gas thus (a) is correct

NOTE : Under high pressure the intermolecular distance decreases and the intermolecular forces play a significant role and the gas shows a devation from ideal behaviour.

Thus (b) is *not* correct.

a, b i.e. the vander Waals coefficients defined on the nature of gas and are independent of temperature so (c) is correct.

The pressure $\left(P+\frac{n^2a}{V}\right)$ is not lower than P so (d) is not correct.

Hence the correct anser is (a, c).

7. **(a, b, c, d)**
8. **(c)** P(V–b) = RT

$\Rightarrow \quad PV - Pb = RT$

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

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

Hence Z > 1 at all pressures.

This means , repulsive tendencies will be dominant when interatomic distance are small.

This means, interatomic potential is never negative but becomes positive at small interatomic distances.

E. Subjective Problems

1. **TIPS/Formulae :**

$$PV = \frac{m}{M}RT \Rightarrow P = \frac{m}{V}\times\frac{RT}{M} = d\frac{RT}{M} \therefore d = \frac{MP}{RT}$$

Substituting the value, we get

$$d = \frac{17\times 5}{0.082\times 303} = \mathbf{3.42\ g/litre}$$

2. Given, moles $= \frac{\text{mass}}{\text{mol.wt}}$

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

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

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

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

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

For gas $P_1V_1 = n_2RT_2$ ....(ii)

($\because$ Pressure and volume of gas are same)

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

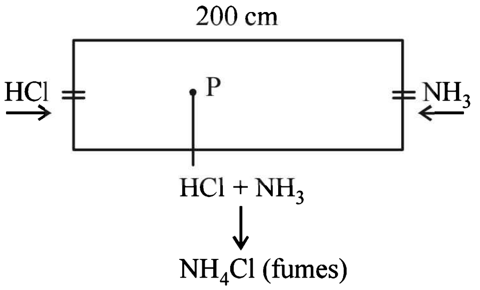
$$\frac{P_1V_1}{P_1V_1} = \frac{n_1RT_1}{n_2RT_2} \text{ or } 1 = \frac{0.092\times 298}{n_2\times 290}$$

$$\text{or } n_2 = \frac{0.092\times 298}{290} \text{ or } \frac{3.7}{\text{M}} = \frac{0.092\times 298}{290}$$

$$\text{or } \frac{3.7}{\text{M}} = 0.0945 \therefore \text{M} = \frac{3.7}{0.0945} = \mathbf{39.15}$$

3. Let NH_3 diffuse through = x cm

HCl diffuses through = y cm



According to Graham's law of diffusion

$$\frac{x}{y} = \sqrt{\frac{\text{Mol. wt HCl}}{\text{Mol. wt of } NH_3}} = \sqrt{\frac{36.5}{17}} = \sqrt{2.14} = 1.465$$

$x = 1.465y$ ... (1)

$x + y = 200$ cm ... (2)

From these equations; y = 85.2 cm

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

4. Following reaction takes places in tube

$C + CO_2 \longrightarrow 2CO$

Volume of mixture of CO and CO_2 = 1L

Let volume of CO_2 in mixture = x

$\therefore$ Volume of CO in mixture = $2x$

$\therefore$ Original volume of CO in mixture = $1 - x$

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

$1 + x = 1.6$ ($\because$ It is given total volume after reaction = 1.6L)

$\therefore x = 0.6$ L $\quad \therefore$ Volume of CO_2 = 0.6 L

Volume of CO = 0.4 L

$CO_2 : CO = \mathbf{3 : 2}$

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

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

$$\text{and } r_2 \text{ (of } NH_3) \text{ at 1 atm. pressure } P = 40 = \frac{k\times 1}{\sqrt{17}} \quad .....(ii)$$

From (i) and (ii)

$$\frac{r_1}{r_2} = \frac{60}{40} = \frac{kP}{\sqrt{36.5}} \times \frac{\sqrt{17}}{k \times 1}$$

$$P = \frac{60}{40} \times \frac{\sqrt{36.5}}{\sqrt{17}} = \mathbf{2.197\ atm}$$

6. **TIPS/Formulae :**

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

where n = Number of moles of the gas

R = Gas constant

T = Absolute temperature

Molecular weight of methane,

$CH_4 = 12 + 4 \times 1 = 16$

$\therefore$ Number of moles of methane in 8.0 gm of methane

$$= \frac{8.0}{16.0} = 0.5$$

R = 8.314 joules/K/mole, $T = 27 + 273 = 300$ K

$\therefore$ Total kinetic energy of the molecules in 8.0 gm of methane at 27°C $= n \times 3/2\ RT = 0.5 \times 3/2 \times 8.314 \times 300 = 1870.65$ joules

$$\therefore \text{Average kinetic energy} = \frac{1870.65}{6.023 \times 10^{23} \times 0.5}$$

$$= \mathbf{6.21 \times 10^{-21}\ joules/molecule}$$

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

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

$$= \frac{7.6 \times 10^{-10}}{760}\ \text{atm.} = 1 \times 10^{-12}\ \text{atm.}$$

V = 1 litre, $T = 273 + 0 = 273$K, R = 0.082 litre atm./K/mol

Putting the values in equation

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

Now since 1 mole = 6.023×10^{23} molecules

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

$$= \mathbf{2.7 \times 10^{10}\ molecules}$$

8. From ideal gas equation,

$$PV = nRT \Rightarrow PV = \left(\frac{m}{M}\right) RT \text{ or } M = m\frac{RT}{PV}$$

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

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

$$\therefore \frac{M_A}{M_B} = \frac{2RT}{V} \times \frac{0.5V}{3RT} = \frac{2 \times 0.5}{3} = \frac{1}{3}$$

Therefore, the ratio $M_A : M_B = \mathbf{1 : 3}$

9. $$U_{rms} = \sqrt{\frac{3RT}{M}}$$

Given $T = 20°C = 20 + 273 = 293$K

$R = 8.314 \times 10^7$ erg per degree per mol

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

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

10. Volume of ballon = 4.851 L (as calculated above)

Let no. of balloons to be filled = n

$\therefore$ Total volume occupied by n balloons = $4.851 \times n$

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

$\therefore$ Total volume of H_2 at NTP = (4.851n + 2.82)L

P_1 = 1 atm	P_2 = 20 atm
$V_1 = 4.85 \times n + 2.82$ L	V_2 = 2.82 L
T_1 = 273 K	T_2 = 300 K

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \text{ or } \frac{1 \times (4.85 \times n + 2.82)}{273} = \frac{20 \times 2.82}{300}$$

$$\therefore n = \frac{48.504}{4.851} \simeq 10$$

11. **TIPS/Formulae :**

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

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

Given -For CO_2

Average velocity at T_1 = Most probable velocity at T_2

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

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

$$\therefore 9 \times 10^2 = \sqrt{\frac{8 \times 8.314 \times T_1}{3.14 \times 44 \times 10^{-3}}} \quad \text{...(A)}$$

[Average velocity at T_1K]

$$\text{and } 9 \times 10^2 = \sqrt{\frac{2 \times 8.314 \times T_2}{44 \times 10^{-3}}} \quad \text{...(B)}$$

[Most probable velocity at T_2K]

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

12. Applying the general gas equation

$$PV = nRT = \frac{m}{M}RT$$

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

$T = 50°C = 50 + 273 = 323$ K

$$\therefore V = \frac{mRT}{MP} \text{ or } V = \frac{5 \times 0.082 \times 323 \times 760}{26 \times 740} = \mathbf{5.23\ L}$$

13. Using gas equation; $PV = nRT$

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

$$= \frac{PV}{RT} = \frac{6 \times 3}{0.0821 \times 300} = 0.7308 \text{ mol.}$$

Thus no. of moles of unknown gas $= 0.7308 - 0.7$
$= 0.0308$ mol.

Now we know that

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

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

$$\therefore \quad M_2 = \left(\frac{r_1}{r_2}\right)^2 M_1 \text{ or } M_2 = \left(\frac{0.7}{0.0308}\right)^2 \times 2 = \mathbf{1033}$$

14. $2NO + O_2 \rightarrow 2NO_2 \rightarrow N_2O_4$

Calculating the number of moles of NO and O_2 by applying the formula, $n = \frac{PV}{RT}$

$$\text{Moles of NO in the larger flask} = \frac{1.053 \times 0.250}{0.082 \times 300} = 0.0107$$

[250 mL = 0.250 L]

$$\text{Moles of } O_2 \text{ in the smaller flask} = \frac{0.789 \times 0.100}{0.082 \times 300} = 0.0032$$

[100 mL = 0.100 L]

The reaction takes place as follows.

	$2NO$	+ O_2	$\longrightarrow N_2O_4$
Mols before reaction	0.0107	0.0032	0
Mole after reaction	(0.0107 – 2 × .0032)	0	0.0032

Hence moles of NO reacting completely with 0.0032 moles of $O_2 = 2 \times 0.0032 = 0.0064$

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

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

$$= \frac{nRT}{V} = \frac{0.0043 \times 0.082 \times 220}{0.350} = \mathbf{0.221 \text{ atm}}$$

[Total volume = 0.250 + 0.100 = 0.350 L]

15. Given $V = 1\text{L} = 10^{-3} \text{m}^3$, $P = 7.57 \times 10^{-3} \text{Nm}^{-2}$, $R = 8.314\text{J}$, $n = 2 \times 10^{21}/6.023 \times 10^{23}$ moles

$$PV = nRT \text{ or } T = \frac{PV}{nR}$$

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

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

(Given U)

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

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

16. TIPS/Formulae :

Partial pressure = Mole fraction × Total pressure

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

$$\left[\text{mole fraction of He} = \frac{4}{5}\right]$$

$$\therefore \quad p_{CH_4} = 20 - 16 = 4 \text{ bar}$$

Now applying the formula

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

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

$\therefore$ Composition of the mixture (He : CH_4) effusing out = **8 : 1**

17. *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

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

Applying gas equation,

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

[27°C = 273 + 27 = 300]

Calculation of pressure in cylinder after use.

Weight of cylinder after use = 23.2 kg

Weight of empty cylinder = 14.8 kg

$$\therefore \text{Wt. of unused gas} = 8.4 \text{ kg} = \frac{8.4 \times 10^3}{58} \text{ moles of butane}$$

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

[V = 2412 L]

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

Weight of used gas = 14.2 – 8.4 = 5.8 kg

Pressure under normal usage conditions = 1 atm

$$V=\frac{nRT}{P}=\frac{5.8\times10^3}{58}\times\frac{0.0821\times300}{1}\qquad\left[\because n=\frac{5.8}{58}\right]$$

$= 2463$ litres $=$ **2.463 m³**

18. Let the volume of ethane in mixture = x litre

$\therefore$ Volume of ethene = $(40 - x)$ litre

Combustion reactions of ethane and ethene are :

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

or $2C_2H_6(g)+7O_2(g)\longrightarrow 4CO_2(g)+6H_2O(l)$

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

Volume of O_2 required for complete combustion of ethane

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

Volume of O_2 required for complete combustion of ethene

$=(40–x)\times3$ [For $(40-x)$ L]

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

Calculation of number of moles (n)

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

$T = 400$ K

$$\text{Since } n=\frac{PV}{RT}=\frac{1\times\left[\frac{7x}{2}+(40-x)3\right]}{0.082\times400}=\frac{7x+(40-x)6}{2\times0.082\times400}$$

$$\text{Mass of } n \text{ moles of } O_2=\left[\frac{7x+(40-x)6}{2\times0.082\times400}\right]\times32=130$$

$$\text{or } 130=\left[\frac{7x+240-6x}{65.6}\right]\times32$$

$$\Rightarrow 8528=32x+240\times32\Rightarrow 32x=848\Rightarrow \text{or } x=\frac{848}{32}=26.5$$

Hence mole fraction (%) of ethane $=\frac{26.5}{40}\times100=$ **66.25%**

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

19.

Mixture	Krypton
$r_{mix}=1.16$	$r_{Kr}=1$
$M_{mix}=?$	$M_{Kr}=84$

We know that

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

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

Determination of the composition of the equilibrium mixture

Let the fraction of Cl_2 molecules dissociated at equlibrium $= x$

	Cl_2	$\rightleftharpoons$ 2Cl	Total
Initially	1	0	1
At equilibrium	$1-x$	$2x$	$1-x+2x=1+x$

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

$$\because\ \frac{\text{Normal molecular mass}}{\text{Experimental molecular mass}}=1+x$$

$$\therefore\ \frac{71}{64.426}=1+\alpha$$

$\therefore \alpha = 0.137 =$ **13.7%.**

20. **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 remains as such and the other reactions involved are :

$$CO(g)+\frac{1}{2}O_2(g)\rightarrow CO_2(g)\qquad ...(i)$$

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

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,

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

$$\therefore\ [20-(a+b)]+\left(a+\frac{1}{2}a\right)+(b+2b)-13$$

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

(Since for gases, volume α no. of moles)

$$\therefore\ \frac{1}{2}a+2b=13 \text{ or } a+4b=26\qquad ...(iv)$$

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

$\therefore CH_4=\frac{4}{20}\times100=$ **20%**, $CO=\frac{10}{20}\times100=$ **50%**

& He = 100 – (20 + 50) = **30%**

21. Weight of liquid = 148 – 50 = 98 g

Volume of liquid $=\frac{98}{0.98}=100$ ml = volume of vessel

It means, vessel of 100 ml contains ideal gas at 760 mm Hg at 300 K

Weight of gas = 50.5 – 50 = 0.5g

using, $PV = nRT = \frac{w}{m}RT$

$$\frac{760}{760}\times\frac{100}{1000}=\frac{0.5}{m}\times 0.082\times 300 \qquad \left[n=\frac{0.5}{m}\right]$$

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

22.

	PCl_5	$\rightleftharpoons$	PCl_3	+	Cl_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 $\frac{\text{Normal mol. wt. of } PCl_5}{\text{Exp. mol. wt. of } PCl_5} = 1+\alpha = 1.4$

or $\frac{208.5}{\text{Exp. mol. wt. of } PCl_5} = 1.4$

∴ Exp. mol. wt. of PCl_5 or m. wt. of mixture $= \frac{208.5}{1.4}$

Now using, $PV = \frac{w}{m}RT$ for mixture

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

23. van der Waals equation for n moles of gas is

$$\left[P+\frac{n^2a}{V^2}\right][V-nb] = nRT$$

Given $V = 4$ litre; $P = 11.0$ atm, $T = 300$ K;
$b = 0.05$ litre mole^{-1}, $n = 2$

Thus, $\left[11+\frac{2^2 a}{4^2}\right][4-2\times 0.05] = 2\times 0.082\times 300$

∴ $a = \mathbf{6.46\ atm\ litre^2\ mol^{-2}}$

24.

	$N_2O_5(g)$	$\rightleftharpoons$	$2NO_2(g)$	+	$\frac{1}{2}O_2(g)$
Initial pressure	600		0		0
Final pressure	600-P		2P		P/2

$P \propto$ moles when V and T are constant
(where moles equivalent to pressure P are decomposed)
Total pressure = $600 - P + 2P + P/2 = 960$ mm of Hg
∴ $P = 240$ mm Hg

Thus moles of N_2O_5 decomposed $= \frac{240}{600} = \mathbf{0.4}$

25. We know that

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

or $\frac{1}{38}\times\frac{57}{1} = \sqrt{\frac{M}{28}}\times\frac{0.8}{1.6}$ ∴ $M = 252$

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

Thus compound of xenon with fluorine is $\mathbf{XeF_6}$

26. (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,

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

Using gas equation, $PV = \frac{w}{m}RT$

Case I. $1\times V = \frac{12}{m}R(t+273)$ ...(1)

Case II. $1.1\times V = \frac{12}{m}R(t+283)$ ...(2)

From (1) and (2), t = –173°C or t = **100 K**
Also from (1), on substituting t and m (120), V = **0.82 litre**

27. van der Waals' equation for one mole of a gas is

$$\left[P+\frac{a}{V^2}\right](V-b) = RT \qquad \text{...(1)}$$

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

Hence, (1) becomes $\left[P+\frac{a}{V^2}\right]V = RT$ or $P = \frac{RT}{V}-\frac{a}{V^2}$

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

$$\therefore P = \frac{0.082\times 273}{22.4}-\frac{3.592}{(22.4)^2} = \mathbf{0.9922\ atm.}$$

28. We know that, Compressibility factor, $Z = \frac{PV}{RT}$

$0.5 = \frac{100\times V}{0.082\times 273}$ ∴ $V = 0.1119$L

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

$$\left(P+\frac{a}{V^2}\right)(V-0) = RT$$

or $PV = RT - \frac{a}{V}$ or $a = RTV - PV^2$

Substituting the values
$a = (0.082\times 0.1119\times 273) - (100\times 0.1119\times 0.1119)$
$= \mathbf{1.253\ atm\ L^2\ mol^{-2}}$

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

(i) From Graham's Law of diffusion

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

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

where $M_v = MW$ of the vapour

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

$\frac{0.36}{18.09}$ mol occupies 1 L volume, so 1 mol occupies

$\frac{18.09}{0.36}$L = 50.25L

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

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

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow V_2 = 22.4\times\frac{500}{273} = \mathbf{41.025L}$$

(iii) Compressibility factor (Z)

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

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

($\because$ actual density is less than given density)

(b) $E = \frac{3}{2}KT = \frac{3}{2} \times \frac{8.31}{6.02 \times 10^{23}} \times 100$

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

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

30. **TIPS/Formulae :**

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

$$\frac{C_{rms}}{C_{av}} = \sqrt{\frac{3RT}{M}} \times \sqrt{\frac{\pi M}{8RT}} = \sqrt{\frac{3\pi}{8}} = 1.085$$

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

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

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

$$PV_m = RT + Pb - \frac{a}{V_m} + \frac{ab}{V_m^2} \qquad \text{...(i)}$$

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

$\therefore PV_m = RT + Pb$

When $P = 0$, intercept = $\boldsymbol{RT}$

F. Match The Following

1.	**(A) : (p) and (s)**	Because 200 atm pressure is very large. For H_2 gas, at very high pressure $Z > 1$.
	(B) : (r)	Since $P \sim 0$, it means very low presure, so ideal behaviour is observed.
	(C) : (p) and (q)	Since P is 1 atm, Z for CO_2 would be less than 1.
	(D) : (r)	In real gas with very high molar volume, molecules will be very far apart from each other due to which van der Waal's forces as well as actual volume occupied by molecules will be negligible.

G. Comprehension Based Questions

1. **(c)** According to Graham's law of diffusion, if all conditions are identical,

$$r = \frac{1}{\sqrt{M}}$$

As in this question, all conditions are identical for X and Y, then

$$\frac{r_x}{r_y} = \sqrt{\frac{M_y}{M_x}}$$

$$\frac{d}{24 - d} = \sqrt{\frac{40}{10}} = 2$$

$\Rightarrow d = 48 - 2d \Rightarrow 3d = 48 \Rightarrow d = 16\,cm$

2. **(d)** The general formula of mean free path (λ) is

$$\lambda = \frac{RT}{\sqrt{2}\pi d^2 N_A p}$$

(d = diameter of molecule, p = pressure inside the vessel)

Since d and p are same for both gases, ideally their λ are same. Hence it must be the higher drift speed of X due to which it is facing more collisions per second with the inert gas in comparison to gas Y. Hence X faces more resistance from inert gas than Y and hence covers lesser distance than that predicted by Graham's law.

H. Assertion & Reason Type Questions

1. **(a)** 'a' indicates the magnitude of the attractive forces among the gas molecules, which increases in NH_3 due to H-bonding.

2. **(a)** At constant volume.

$P \propto T$ (from $PV = nRT$)

$$U_{rms} = \sqrt{\frac{3RT}{M}} \quad \therefore \quad U_{rms} \propto \sqrt{T} \qquad \text{......(i)}$$

Collision frequency is directly proportional to U_{rms}.

Thus collision frequency $\propto \sqrt{T}$

Hence, on increasing the collosion pressure, increases frequency.

I. Integer Value Correct Type

1. v_{rms} of $X = \sqrt{\frac{3RT_x}{M_x}}$; v_{mp} of $Y = \sqrt{\frac{2RT_y}{M_y}}$

Given $v_{rms} = v_{mp} \Rightarrow \sqrt{\frac{3RT_x}{M_x}} = \sqrt{\frac{2RT_y}{M_y}}$

$$\Rightarrow M_y = \frac{2RT_yM_x}{3RT_x} = \frac{2 \times 60 \times 40}{3 \times 400} = 4$$

2. **7**

$P_{He} = 1 - 0.68 = 0.32$ atm, $n = 0.1$

$$V = \frac{nRT}{P} = \frac{0.1 \times 0.0821 \times 273}{0.32} = 7$$

3. **(4)** Diffusion coefficient $\propto \lambda\mu$

Since $\lambda \propto \frac{T}{P}$ and $\mu \propto \sqrt{T}$

$\therefore$ Diffusion coefficient $\propto \frac{T\sqrt{T}}{P}$

$$\text{Thus } \frac{D_i}{D_f} = \frac{\frac{T\sqrt{T}}{P}}{\frac{4T\sqrt{4T}}{2P}} = \frac{1}{(4 \times 2)/2} = \frac{1}{4} \quad \text{or} \quad \frac{D_f}{D_i} = \frac{4}{1}$$

Section-B JEE Main/ AIEEE

1. **(c)** PV = nRT (number of moles = n/V) $\therefore$ n/V = P/RT.

2. **(c)** Value of gas constant

$$(R) = 0.0821 \text{ L atm } K^{-1} mol^{-1}$$
$$= 8.314 \times 10^7 \text{ ergs } K^{-1} mol^{-1}$$
$$= 8.314 JK^{-1} mol^{-1} = 1.987 \text{ cal } K^{-1} mol^{-1}$$

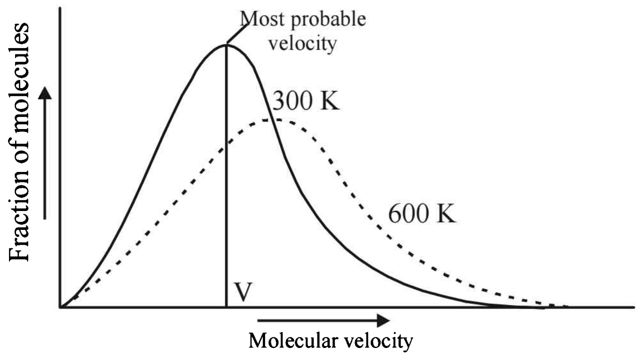
3. **(d)** Kinetic theory of gases proves all the given gas laws.

4. **(b)** According to kinetic theory the gas molecules are in a state of constant rapid motion in all possible directions colloiding in a random manner with one another and with the walls of the container and between two successive collisions molecules travel in a straight line path but show haphazard motion due to collisions.

5. **(a)** $$\frac{\text{K.E of neon at } 40°C}{\text{K.E of neon at } 20°C} = \frac{\frac{3}{2}K \times 313}{\frac{3}{2}K \times 293} = \frac{313}{293}$$

6. **(a)** In van der waals equation 'b' is for volume correction

7. **(c)** Distribution of molecular velocities at two different temperature is given shown below.



NOTE : At higher temperature more molecules have higher velocities and less molecules have lower velocities.

As evident from fig. thus it is clear that With the increase in temperature the most probable velocity increase butthe fraction of such molecules decreases.

8. **(d)** From the ideal gas equation :

$$PV = nRT \text{ or } n = \frac{PV}{RT} = \frac{3170 \times 10^{-3}}{8.314 \times 300} = 1.27 \times 10^{-3}$$

9. **(d)** The value of a is a measure of the magnitude of the attractive forces between the molecules of the gas. Greater the value of 'a', larger is the attractive intermolecular force between the gas molecules.

The value of b related to the effective size of the gas molecules. It is also termed as excluded volume.

The gases with higher value of a and lower value of b are more liquefiable, hence for Cl_2 "a" should be greater than for C_2H_6 but for it b should be less than for C_2H_6.

10. **(c)** $\left(P + \frac{a}{V^2}\right)(V - b) = RT$ at high pressure $\frac{a}{V^2}$ can be neglected

$$PV - Pb = RT \text{ and } PV = RT + Pb$$

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

$$Z = 1 + \frac{Pb}{RT}; \; Z > 1 \text{ at high pressure}$$

11. **(c)** Most probable speed (C*) = $\sqrt{\frac{2RT}{M}}$

Average Speed $(\bar{C}) = \sqrt{\frac{8RT}{\pi M}}$

Root mean square velocity (c) = $\sqrt{\frac{3RT}{M}}$

$$C^* : \bar{C} : C = \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}}$$

$$= 1 : \sqrt{\frac{4}{\pi}} : \sqrt{\frac{3}{2}} = 1 : 1.128 : 1.225$$

12. **(b)** Compressibility factor $(Z) = \frac{PV}{RT}$

(For one mole of real gas)

van der Waals equation

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

At low pressure, volume is very large and hence correction term b can be neglected in comparison to very large volume of V.

i.e. $V - b \approx V$

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

$$PV = RT - \frac{a}{V} \;;\; \frac{PV}{RT} = 1 - \frac{a}{VRT}$$

Hence, $\boxed{Z = 1 - \frac{a}{VRT}}$

13. **(b)** Number of moles of $O_2 = \frac{w}{32}$

Number of moles of $N_2 = \frac{4w}{28} = \frac{w}{7}$

$\therefore$ Ratio $= \frac{w}{32} : \frac{w}{7} = 7 : 32$

14. **(b)** Hydrogen bond is a type of strong electrostatic dipole-dipole interaction and dependent on the inverse cube of distance between the molecular ion-dipole interaction $\propto \frac{1}{r^3}$.

15. **(a)** For a given mass of an ideal gas, the volume and amount (moles) of the gas are directly proportional if the temperature and pressure are constant. i.e

$V \propto n$

Hence in the given case.

Initial moles and final moles are equal $(n_T)_i = (n_T)_f$

$$\frac{P_iV}{RT_1} + \frac{P_iV}{RT_1} = \frac{P_fV}{RT_1} + \frac{P_fV}{RT_2}$$

$$2\frac{P_i}{T_1} = \frac{P_f}{T_1} + \frac{P_f}{T_2} \Rightarrow P_f = \frac{2P_iT_2}{T_1 + T_2}$$