CHAPTER

5

States of Matter

5.1 Intermolecular Forces

1. Dipole-induced dipole interactions are present in which of the following pairs?

(a) HCl and He atoms (b) SiF_4 and He atoms

(c) H_2O and alcohol (d) Cl_2 and CCl_4

(NEET 2013)

2. Which one of the following is the correct order of interactions?

(a) Covalent < hydrogen bonding < van der Waals' < dipole-dipole

(b) van der Waals' < hydrogen bonding < dipole-dipole < covalent

(c) van der Waals' < dipole-dipole < hydrogen bonding < covalent

(d) Dipole-dipole < van der Waals' < hydrogen bonding < covalent. *(1993)*

5.4 The Gaseous State

3. Which of the following statements is wrong for gases?

(a) Confined gas exerts uniform pressure on the walls of its container in all directions.

(b) Volume of the gas is equal to volume of container confining the gas.

(c) Gases do not have a definite shape and volume.

(d) Mass of a gas cannot be determined by weighing a container in which it is enclosed. *(1999)*

5.5 The Gas Laws

4. At 25°C and 730 mm pressure, 380 mL of dry oxygen was collected. If the temperature is constant, what volume will the oxygen occupy at 760 mm pressure?

(a) 569 mL (b) 365 mL

(c) 265 mL (d) 621 mL *(1999)*

5. Pressure remaining the same, the volume of a given mass of an ideal gas increases for every degree centrigrade rise in temperature by definite fraction of its volume at

(a) 0°C (b) its critical temperature

(c) absolute zero (d) its Boyle temperature.

(1989)

5.6 Ideal Gas Equation

6. A mixture of N_2 and Ar gases in a cylinder contains 7 g of N_2 and 8 g of Ar. If the total pressure of the mixture of the gases in the cylinder is 27 bar, the partial pressure of N_2 is

[Use atomic masses (in g mol^{-1}) : N = 14, Ar = 40]

(a) 9 bar (b) 12 bar

(c) 15 bar (d) 18 bar. *(NEET 2020)*

7. The volume occupied by 1.8 g of water vapour at 374°C and 1 bar pressure will be

[Use $R = 0.083$ bar L K^{-1} mol^{-1}]

(a) 96.66 L (b) 55.87 L

(c) 3.10 L (d) 5.37 L

(Odisha NEET 2019)

8. Equal moles of hydrogen and oxygen gases are placed in a container with a pin-hole through which both can escape. What fraction of the oxygen escapes in the time required for one-half of the hydrogen to escape?

(a) 3/8 (b) 1/2

(c) 1/8 (d) 1/4 *(NEET-I 2016)*

9. What is the density of N_2 gas at 227°C and 5.00 atm pressure? ($R = 0.082$ L atm$K^{-1}$$mol^{-1}$)

(a) 1.40 g/mL (b) 2.81 g/mL

(c) 3.41 g/mL (d) 0.29 g/mL

(Karnataka NEET 2013)

10. 50 mL of each gas *A* and of gas *B* takes 150 and 200 seconds respectively for effusing through a pin hole under the similar conditions. If molecular mass of gas *B* is 36, the molecular mass of gas *A* will be

(c) 32 (d) 64 *(2012)*

11. A certain gas takes three times as long to effuse out as helium. Its molecular mass will be

(a) 27 u (b) 36 u
(c) 64 u (d) 9 u *(Mains 2012)*

12. Two gases *A* and *B* having the same volume diffuse through a porous partition in 20 and 10 seconds respectively. The molecular mass of *A* is 49 u. Molecular mass of *B* will be

(a) 50.00 u (b) 12.25 u
(c) 6.50 u (d) 25.00 u *(2011)*

13. A gaseous mixture was prepared by taking equal moles of CO and N_2. If the total pressure of the mixture was found 1 atmosphere, the partial pressure of the nitrogen (N_2) in the mixture is

(a) 0.5 atm (b) 0.8 atm
(c) 0.9 atm (d) 1 atm *(2011)*

14. A bubble of air is underwater at temperature 15°C and the pressure 1.5 bar. If the bubble rises to the surface where the temperature is 25°C and the pressure is 1.0 bar, what will happen to the volume of the bubble?

(a) Volume will become greater by a factor of 1.6.
(b) Volume will become greater by a factor of 1.1.
(c) Volume will become smaller by a factor of 0.70.
(d) Volume will become greater by a factor of 2.5.
(Mains 2011)

15. The pressure exerted by 6.0 g of methane gas in a 0.03 m^3 vessel at 129°C is (Atomic masses: C = 12.01, H = 1.01 and $R = 8.314\ J\ K^{-1}\ mol^{-1}$)

(a) 215216 Pa (b) 13409 Pa
(c) 41648 Pa (d) 31684 Pa *(Mains 2010)*

16. Which of the following mixtures of gases does not obey Dalton's law of partial pressure?

(a) Cl_2 and SO_2 (b) CO_2 and He
(c) O_2 and CO_2 (d) N_2 and O_2 *(1996)*

17. At what temperature, the rate of effusion of N_2 would be 1.625 times than the rate of SO_2 at 50°C?

(a) 373°C (b) 620°C
(c) 100°C (d) 173°C *(1996)*

18. 50 mL of hydrogen diffuses out through a small hole of a vessel, in 20 minutes. The time taken by 40 mL of oxygen to diffuse out is

(a) 32 minutes (b) 64 minutes
(c) 8 minutes (d) 12 minutes *(1994)*

19. Under what conditions will a pure sample of an ideal gas not only exhibit a pressure of 1 atm but also a concentration of 1 mole $litre^{-1}$?
($R = 0.082$ litre atm mol^{-1} deg^{-1})

(a) At STP
(b) When $V = 22.4$ litres
(c) When $T = 12$ K
(d) Impossible under any conditions *(1993)*

20. The correct value of the gas constant '*R*' is close to

(a) 0.082 litre-atmosphere K
(b) 0.082 litre-atmosphere K^{-1} mol^{-1}
(c) 0.082 litre-$atmosphere^{-1}$ K mol^{-1}
(d) 0.082 $litre^{-1}$ $atmosphere^{-1}$ K mol. *(1992)*

21. Select one correct statement. In the gas equation, $PV = nRT$

(a) *n* is the number of molecules of a gas
(b) *V* denotes volume of one mole of the gas
(c) *n* moles of the gas have a volume *V*
(d) *P* is the pressure of the gas when only one mole of gas is present. *(1992)*

22. At constant temperature, in a given mass of an ideal gas

(a) the ratio of pressure and volume always remains constant
(b) volume always remains constant
(c) pressure always remains constant
(d) the product of pressure and volume always remains constant. *(1991)*

23. If *P, V, M, T* and *R* are pressure, volume, molar mass, temperature and gas constant respectively, then for an ideal gas, the density is given by

(a) $\frac{RT}{PM}$ (b) $\frac{P}{RT}$ (c) $\frac{M}{V}$ (d) $\frac{PM}{RT}$
(1989)

24. Correct gas equation is

(a) $\frac{V_1T_2}{P_1} = \frac{V_2T_1}{P_2}$ (b) $\frac{P_1V_1}{P_2V_2} = \frac{T_1}{T_2}$
(c) $\frac{P_1T_1}{V_1} = \frac{P_2V_2}{T_2}$ *(d)* $\frac{V_1V_2}{T_1T_2} = P_1P_2$ *(1989)*

5.7 Kinetic Energy and Molecular Speeds

25. By what factor does the average velocity of a gaseous molecule increase when the temperature (in Kelvin) is doubled?

(a) 2.0 (b) 2.8
(c) 4.0 (d) 1.4 *(2011)*

26. The temperature of a gas is raised from 27°C to 927°C. The root mean square speed of the gas

(a) remains same (b) gets $\sqrt{\frac{927}{27}}$ times
(c) gets halved (d) gets doubled. *(1994)*

27. The ratio among most probable velocity, mean velocity and root mean square velocity is given by

(a) $1:2:3$ (b) $1:\sqrt{2}:\sqrt{3}$
(c) $\sqrt{2}:\sqrt{3}:\sqrt{8/\pi}$ (d) $\sqrt{2}:\sqrt{8/\pi}:\sqrt{3}$
(1993)

28. The root mean square velocity at STP for the gases H_2, N_2, O_2 and HBr are in the order

(a) $H_2 < N_2 < O_2 < HBr$ (b) $HBr < O_2 < N_2 < H_2$
(c) $H_2 < N_2 = O_2 < HBr$ (d) $HBr < O_2 < H_2 < N_2$
(1991)

29. Root mean square velocity of a gas molecule is proportional to
(a) $m^{1/2}$ (b) m^0
(c) $m^{-1/2}$ (d) m *(1990)*

5.8 Kinetic Molecular Theory of Gases

30. The energy absorbed by each molecule (A_2) of a substance is 4.4×10^{-19} J and bond energy per molecule is 4.0×10^{-19} J. The kinetic energy of the molecule per atom will be
(a) 2.2×10^{-19} J (b) 2.0×10^{-19} J
(c) 4.0×10^{-20} J (d) 2.0×10^{-20} J *(2009)*

31. If a gas expands at constant temperature, it indicates that
(a) kinetic energy of molecules remains the same
(b) number of the molecules of gas increases
(c) kinetic energy of molecules decreases
(d) pressure of the gas increases. *(2008)*

32. Average molar kinetic energy of CO and N_2 at same temperature is
(a) $KE_1 = KE_2$
(b) $KE_1 > KE_2$
(c) $KE_1 < KE_2$
(d) can't say anything. Both volumes are not given.
(2000)

33. The average kinetic energy of an ideal gas, per molecule in S.I. units, at 25°C will be
(a) 6.17×10^{-20} J (b) 7.16×10^{-20} J
(c) 61.7×10^{-20} J (d) 6.17×10^{-21} J *(1996)*

34. At STP, 0.50 mol H_2 gas and 1.0 mol He gas
(a) have equal average kinetic energies
(b) have equal molecular speeds
(c) occupy equal volumes
(d) have equal effusion rates. *(1993)*

35. Internal energy and pressure of a gas per unit volume are related as
(a) $P = \frac{2}{3}E$ (b) $P = \frac{3}{2}E$
(c) $P = \frac{1}{2}E$ (d) $P = 2E$ *(1993)*

36. A closed flask contains water in all its three states solid, liquid and vapour at 0°C. In this situation, the average kinetic energy of water molecules will be
(a) the greatest in all the three states
(b) the greatest in vapour state
(c) the greatest in the liquid state
(d) the greatest in the solid state. *(1992)*

37. Which is not true in case of an ideal gas?
(a) It cannot be converted into a liquid.
(b) There is no interaction between the molecules.
(c) All molecules of the gas move with same speed.
(d) At a given temperature, PV is proportional to the amount of the gas. *(1992)*

5.9 Behaviour of Real Gases - Deviation from Ideal Gas Behaviour

38. A gas at 350 K and 15 bar has molar volume 20 percent smaller than that for an ideal gas under the same conditions. The correct option about the gas and its compressibility factor (Z) is
(a) $Z < 1$ and repulsive forces are dominant
(b) $Z > 1$ and attractive forces are dominant
(c) $Z > 1$ and repulsive forces are dominant
(d) $Z < 1$ and attractive forces are dominant.
(NEET 2019)

39. A gas such as carbon monoxide would be most likely to obey the ideal gas law at
(a) low temperatures and high pressures
(b) high temperatures and high pressures
(c) low temperatures and low pressures
(d) high temperatures and low pressures. *(2015)*

40. Maximum deviation from ideal gas is expected from
(a) $CH_{4(g)}$ (b) $NH_{3(g)}$
(c) $H_{2(g)}$ (d) $N_{2(g)}$ *(NEET 2013)*

41. For real gases van der Waals' equation is written as $\left(p + \frac{an^2}{V^2}\right)(V - nb) = nRT$ where a and b are van der Waals' constants. Two sets of gases are
(I) O_2, CO_2, H_2 and He
(II) CH_4, O_2 and H_2
The gases given in set-I in increasing order of b and gases given in set-II in decreasing order of a, are arranged below. Select the correct order from the following :
(a) (I) $He < H_2 < CO_2 < O_2$ (II) $CH_4 > H_2 > O_2$
(b) (I) $O_2 < He < H_2 < CO_2$ (II) $H_2 > O_2 > CH_4$
(c) (I) $H_2 < He < O_2 < CO_2$ (II) $CH_4 > O_2 > H_2$
(d) (I) $H_2 < O_2 < He < CO_2$ (II) $O_2 > CH_4 > H_2$
(Mains 2012)

42. van der Waals' real gas, acts as an ideal gas, at which conditions?
(a) High temperature, low pressure
(b) Low temperature, high pressure
(c) High temperature, high pressure
(d) Low temperature, low pressure *(2002)*

43. When is deviation more in the behaviour of a gas from the ideal gas equation $PV = nRT$?

(a) At high temperature and low pressure
(b) At low temperature and high pressure
(c) At high temperature and high pressure
(d) At low temperature and low pressure *(1993)*

44. A gas is said to behave like an ideal gas when the relation PV/T = constant. When do you expect a real gas to behave like an ideal gas?

(a) When the temperature is low.
(b) When both the temperature and pressure are low.
(c) When both the temperature and pressure are high.
(d) When the temperature is high and pressure is low. *(1991)*

45. In van der Waals' equation of state for a non-ideal gas, the term that accounts for intermolecular forces is

(a) $(V - b)$ (b) $(RT)^{-1}$
(c) $\left(P + \frac{a}{V^2}\right)$ (d) RT *(1990)*

5.10 Liquefaction of Gases

46. Given van der Waals' constant for NH_3, H_2, O_2 and CO_2 are respectively 4.17, 0.244, 1.36 and 3.59, which one of the following gases is most easily liquefied?

(a) NH_3 (b) H_2
(c) O_2 (d) CO_2 *(NEET 2018)*

47. An ideal gas, obeying kinetic theory of gases cannot be liquefied, because

(a) it solidifies before becoming a liquid
(b) forces acting between its molecules are negligible
(c) its critical temperature is above 0°C
(d) its molecules are relatively small in size. *(1995)*

5.11 Liquid State

48. The beans are cooked earlier in pressure cooker because

(a) boiling point increases with increasing pressure
(b) boiling point decreases with increasing pressure
(c) extra pressure of pressure cooker softens the beans
(d) internal energy is not lost while cooking in pressure cooker. *(2011)*

ANSWER KEY

1.	(a)	**2.**	(b)	**3.**	(d)	**4.**	(b)	**5.**	(a)	**6.**	(c)	**7.**	(d)	**8.**	(c)	**9.**	(c)	**10.**	(None)
11.	(b)	**12.**	(b)	**13.**	(a)	**14.**	(a)	**15.**	(c)	**16.**	(a)	**17.**	(c)	**18.**	(b)	**19.**	(c)	**20.**	(b)
21.	(c)	**22.**	(d)	**23.**	(d)	**24.**	(b)	**25.**	(d)	**26.**	(d)	**27.**	(d)	**28.**	(b)	**29.**	(c)	**30.**	(d)
31.	(a)	**32.**	(a)	**33.**	(d)	**34.**	(a)	**35.**	(a)	**36.**	(b)	**37.**	(c)	**38.**	(d)	**39.**	(d)	**40.**	(b)
41.	(c)	**42.**	(a)	**43.**	(b)	**44.**	(d)	**45.**	(c)	**46.**	(a)	**47.**	(b)	**48.**	(a)				

Hints & Explanations

1. **(a) :** HCl is polar ($\mu \neq 0$) and He is non-polar ($\mu = 0$) gives dipole-induced dipole interactions.

2. **(b) :** The strength of interaction follows the order : van der Waals' < hydrogen-bonding < dipole-dipole < covalent. It is so because bond length of H-bond is larger than that of a covalent bond.

And also covalent bond is strongest because, the greater the extent of overlapping, the stronger is the bond formed.

3. **(d) :** Mass of the gas = Mass of the cylinder including gas – Mass of empty cylinder.

So, mass of a gas can be determined by weighing the container in which it is enclosed.

Thus, the statement (d) is wrong for gases.

4. **(b) :** $V_1 = 380$ mL, $P_1 = 730$ mm, $V_2 = ?$, $P_2 = 760$ mm.

From Boyle's law, $P_1V_1 = P_2V_2$

$$\Rightarrow V_2 = \frac{P_1V_1}{P_2} = \frac{730 \times 380}{760} = 365 \text{ mL}$$

5. **(a) :** According to Charles' law which states that "The volume of the given mass of a gas increases or decreases by $\frac{1}{273}$ of its volume at 0°C for each degree rise or fall of temperature at constant pressure."

$V_t = V_0\left(1+\frac{t}{273}\right)$ at constant P and n.

6. (c) : Number of moles of $N_2 = \frac{7}{28} = 0.25$ mol

Number of moles of Ar $= \frac{8}{40} = 0.2$ mol

Mole fraction of $N_2 = \frac{0.25}{0.25+0.2} = \frac{0.25}{0.45} = 0.55$

Partial pressure of N_2 gas = mole fraction × total pressure $= 0.55 \times 27 = 14.85 \approx 15$ bar

7. (d) : $m = 1.8\text{ g} \Rightarrow n = \frac{m}{M} = \frac{1.8}{18} = 0.1\text{ mol}$

$T = 374°\text{C} = 647\text{ K}$, $P = 1$ bar, $R = 0.083\text{ bar L K}^{-1}\text{ mol}^{-1}$

$$V = \frac{nRT}{P} = \frac{0.1\times 0.083\times 647}{1} = 5.37\text{ L}$$

8. (c) : Let the number of moles of each gas $= x$

Fraction of hydrogen escaped $= \frac{1}{2}x$

$$\frac{r_{O_2}}{r_{H_2}} = \sqrt{\frac{M_{H_2}}{M_{O_2}}} \Rightarrow \frac{n_{O_2}/t}{\frac{x}{2}/t} = \sqrt{\frac{2}{32}} = \sqrt{\frac{1}{16}} = \frac{1}{4}$$

$$\Rightarrow \frac{n_{O_2}/t}{\frac{x}{2}/t} = \frac{1}{4} \Rightarrow n_{O_2} = \frac{1}{8}x$$

Hence, fraction of oxygen escaped $= \frac{1}{8}$

9. (c) : $PV = nRT$

$$PV = \frac{w}{M}RT \quad \left[n = \frac{\text{Weight of the gas taken }(W)}{\text{Mol. mass of gas }(M)}\right]$$

$$P = \frac{w}{M}\times\frac{RT}{V}$$

$$P = \frac{dRT}{M} \qquad \left[\text{Density} = \frac{\text{Mass}}{\text{Volume}}\right]$$

$$d = \frac{PM}{RT} = \frac{\square 5\times 28}{0.0821\times 500} = 3.41\text{ g/mL}$$

10. (None) : According to Graham's law of diffusion,

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}},\ r_A = \frac{V_A}{T_A},\ r_B = \frac{V_B}{T_B}$$

$$\frac{V_A/T_A}{V_B/T_B} = \sqrt{\frac{M_B}{M_A}}$$

$V_A = V_B$, $T_A = 150$ sec, $T_B = 200$ sec, $M_B = 36$, $M_A = ?$

$$\frac{T_B}{T_A} = \sqrt{\frac{M_B}{M_A}} \Rightarrow \frac{200}{150} = \sqrt{\frac{36}{M_A}}$$

$$\frac{4}{3} = \sqrt{\frac{36}{M_A}} \text{ or } \frac{4\times 4}{3\times 3} = \frac{36}{M_A}$$

$$\text{or } M_A = \frac{36}{4\times 4}\times 3\times 3 = 20.25$$

11. (b) : According to Graham's law of diffusion,

$$r \propto \frac{1}{\sqrt{d}} \propto \frac{1}{\sqrt{M}} \Rightarrow \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Rate of diffusion $= \frac{\text{Volume of gas diffused }(V)}{\text{Time taken }(t)}$

$$\therefore \frac{V_1/t_1}{V_2/t_2} = \sqrt{\frac{M_2}{M_1}}$$

If same volume of two gases diffuse, then $V_1 = V_2$

$$\therefore \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

Here $t_2 = 3t_1$, $M_1 = 4$ u, $M_2 = ?$

$$\therefore \frac{3t_1}{t_1} = \sqrt{\frac{M_2}{4}} \Rightarrow 3 = \sqrt{\frac{M_2}{4}}$$

$$\Rightarrow 9 = \frac{M_2}{4} \Rightarrow M_2 = 36\text{ u}$$

12. (b) : We know that $\frac{r_A}{r_B} = \frac{V/t_A}{V/t_B} = \sqrt{\frac{M_B}{M_A}}$

$$\frac{t_B}{t_A} = \sqrt{\frac{M_B}{M_A}} \Rightarrow \frac{10}{20} = \sqrt{\frac{M_B}{49}}$$

$$\Rightarrow \left(\frac{10}{20}\right)^2 = \frac{M_B}{49} \Rightarrow \frac{100}{400} = \frac{M_B}{49}$$

$$\Rightarrow M_B = \frac{49\times 100}{400} = 12.25\text{ u}$$

13. (a) : $p_{CO} + p_{N_2} = 1$ atm

$2p_{N_2} = 1$ $\quad [\because n_{CO} = n_{N_2}]$

$$p_{N_2} = \frac{1}{2} = 0.5\text{ atm}$$

14. (a) : From ideal gas equation, $V \propto \frac{T}{P}$

Given $T_1 = 15 + 273 = 288$ K, $P_1 = 1.5$ bar

$T_2 = 25 + 273 = 298$ K, $P_2 = 1$ bar

$$V_1 \propto \frac{288}{1.5} \text{ } i.e., V_1 \propto 192 \text{ and } V_2 \propto \frac{298}{1}$$

$$\frac{V_2}{V_1} = \frac{298}{192} = 1.55 \approx 1.6$$

15. (c) : Given, mass of CH_4, $w = 6$ g

Volume of CH_4, $V = 0.03\text{ m}^3$

$T = 129°\text{C} = 129 + 273 = 402$ K

$R = 8.314\text{ J K}^{-1}\text{ mol}^{-1}$

Molecular mass of CH_4, $M = 12.01 + 4 \times 1.01 = 16.05$

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

$$\therefore P = \frac{w}{M}\frac{RT}{V} = \frac{6}{16.05}\times\frac{8.314\times 402}{0.03}$$

$= 41647.7\text{ Pa} \approx 41648\text{ Pa}$

16. (a) : $Cl_2 + SO_2 \xrightarrow{\text{Sunlight}} SO_2Cl_2$ (Sulphuryl chloride)

Dalton's law of partial pressure is applicable only in those cases where gases are non-reacting. As Cl_2 and SO_2 reacts to form SO_2Cl_2 so this law is not obeyed in given case.

17. (c) : $r_1 = 1.625 r_2$ and $T_2 = 50°C = 323$ K

We know that $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1} \times \frac{T_1}{T_2}}$

or $1.625 = \sqrt{\frac{64}{28} \times \frac{T_1}{323}}$

or $T_1 = \frac{(1.625)^2 \times 28 \times 323}{64} = 373.15\,K = 100.15°C$

18. (b) : Volume of hydrogen = 50 mL; Time for diffusion (t) = 20 min and volume of oxygen = 40 mL.

Rate of diffusion of hydrogen (r_1) = $\frac{50}{20}$ = 2.5 mL/min

Rate of diffusion of oxygen (r_2) = $\frac{40}{t}$ mL/min

Since the molecular mass of hydrogen (M_1) = 2 and that of oxygen (M_2) = 32, therefore

$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \Rightarrow \frac{2.5}{40/t} = \sqrt{\frac{32}{2}} \Rightarrow \frac{t}{16} = 4 \Rightarrow t = 64$ minutes

19. (c) : $PV = nRT$ or $P = \frac{n}{V}RT = CRT$

Hence, $1 = 1 \times 0.082 \times T \Rightarrow T = \frac{1}{0.082} = 12$ K

20. (b)

21. (c) : In ideal gas equation, $PV = nRT$

n moles of the gas have volume V.

22. (d) : According to Boyle's law at constant temperature, $P \propto \frac{1}{V}$ or $PV =$ constant

23. (d) : Ideal gas equation is

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

or $PM = \frac{m}{V}RT = dRT$ [here d = density]

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

24. (b) : $\frac{PV}{T}$ = constant or $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \Rightarrow \frac{P_1V_1}{P_2V_2} = \frac{T_1}{T_2}$

25. (d) : Average velocity $= \sqrt{\frac{8RT}{\pi M}}$

When T becomes $2T$ then

average velocity $= \sqrt{\frac{8R(2T)}{\pi M}}$

i.e., $\sqrt{2}$ or 1.41 times increase.

26. (d) : $T_1 = 27°C = 300$ K and $T_2 = 927°C = 1200$ K

We know that root mean square speed $(v) \propto \sqrt{T}$.

Therefore root mean square speed of the gas, when its temperature is raised $= \sqrt{\frac{T_2}{T_1}} = \sqrt{\frac{1200}{300}} = 2$ times

27. (d) : Most probable velocity, $(u_{mp}) = \sqrt{\frac{2RT}{M}}$

Mean velocity, $(\bar{v}) = \sqrt{\frac{8RT}{\pi M}}$

Root mean square velocity, $(u_{r.m.s}) = \sqrt{\frac{3RT}{M}}$

$\therefore u_{mp} : \bar{v} : u_{r.m.s} = \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}}$

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

28. (b) : We know, $PV = \frac{1}{3}mnu^2 = \frac{1}{3}Mu^2$

or $u = \sqrt{3PV/M}$

At STP, $u \propto \sqrt{\frac{1}{M}}$

and molecular masses of H_2, N_2, O_2 and HBr are 2, 28, 32 and 81.

29. (c) : $PV = \frac{1}{3}mNu^2$,

here u = root mean square velocity.

Now $u^2 = \frac{3PV}{mN}$ or $u \propto \frac{1}{\sqrt{m}}$

30. (d) : Energy absorbed by each molecule

$= 4.4 \times 10^{-19}$ J

Energy required to break the bond = 4.0×10^{-19} J

Remaining energy to get converted to kinetic energy $= (4.4 \times 10^{-19} - 4.0 \times 10^{-19})$ J $= 0.4 \times 10^{-19}$ J per molecule

$\therefore$ Kinetic energy per atom $= 0.2 \times 10^{-19}$ J or 2×10^{-20} J

31. (a) : The average translational *K.E.* of one molecule of an ideal gas will be given by

$E_t = \frac{K.E.}{N_A} = \frac{3/2\,RT}{N_A} = \frac{3}{2}kT$

where R/N_A = Boltzmann constant *i.e.* $E_t \propto T$

So, at constant temperature, *K.E.* of molecules remains the same.

32. (a) : $K.E. = \frac{3}{2}RT$ (for one mole of a gas)

As temperatures are same and *KE* is independent of molecular mass, so $KE_1 = KE_2$.

33. (d) : Temperature (T) = 25°C = 298 K.

Therefore, *K.E.* per molecule

$= \frac{3RT}{2N_A} = \frac{3 \times 8.314 \times 298}{2 \times (6.02 \times 10^{23})} = 6.17 \times 10^{-21}$ J

34. (a) : Because average kinetic energy depends only on temperature $K.E. = \frac{3}{2}kT$

35. (a) : $PV = \frac{1}{3}mnu^2 = \frac{1}{3}Mu^2$

$$= \frac{2}{3}\cdot\frac{1}{2}Mu^2 = \frac{2}{3}E \quad \left\{\because \frac{1}{2}Mu^2 = E\right\}$$

or $P = \frac{2}{3}E$ per unit volume.

36. (b) : Velocity and hence average *K.E.* of water molecules is maximum in the gaseous state.

37. (c) : Molecules in an ideal gas move with different speeds. Due to collision between the particles their speed changes.

38. (d) : $V_{\text{ideal}} = V$, $V_{\text{real}} = V - 0.2\text{ V} = 0.8\text{ V}$

$$Z = \frac{V_{\text{real}}}{V_{\text{ideal}}} = 0.8$$

if value of $Z < 1$ then attractive forces are dominant.

39. (d) : Real gases show ideal gas behaviour at high temperatures and low pressures.

40. (b) : NH_3 is a polar molecule, thus more attractive forces between NH_3 molecules.

41. (c) : van der Waals' gas constant '*a*' represent intermolecular forces of attraction of gaseous molecules and van der Waals' gas constant '*b*' represent effective size of molecules. Therefore order should be

(I) $H_2 < He < O_2 < CO_2$ (II) $CH_4 > O_2 > H_2$

42. (a) : At low pressure and high temperature van der Waals real gas acts as ideal gas and observed to obey $PV = nRT$ relation. At very low pressure when the gas-volume is quite large the space occupied by the molecules themselves becomes negligible comparatively and because the molecules are then far apart, the force of mutual attraction becomes too feeble, the real gas would satisfy the postulates of kinetic theory. As temperature is raised, the volume of the gas increases and we can consider $\left(P + \frac{n^2a}{V^2}\right)$ term as P and at low pressure $(V - nb)$ term as V.

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT \quad \text{(van der Waals' equation)}$$

This equation becomes $PV = nRT$

This is an ideal gas equation.

43. (b) : At low temperature and high pressure, there is a deviation from the ideal behaviour in gases.

44. (d) : At high temperature and low pressure the effect of a/V^2 and b is negligible.

As we know, $PV = nRT$ (Ideal gas equation)

$$PV = RT \text{ or } \frac{PV}{RT} = 1$$

$\therefore \quad Z = 1$ [*Z* is compressibility factor]

Hence gas shows ideal behaviour.

45. (c) : van der Waals' equation for 1 mole is

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

Here, $\left(P + \frac{a}{V^2}\right)$ represents the intermolecular forces and $(V - b)$ is the correct volume.

46. (a) : van der Waals' constant '*a*' signifies the intermolecular forces of attraction between the particle of gas. So, higher the value of '*a*', easier will be the liquefaction of gas.

47. (b) : A gas can only be liquefied, if some forces of attraction are acting in its molecules. According to kinetic theory, an ideal gas is devoid of force of attraction in its molecules, therefore it cannot be liquefied.

48. (a) : More is the pressure, greater will be the boiling point.