States of Matter : Gases and Liquids

OBJECTIVE TYPE QUESTIONS

Multiple Choice Questions (MCQs)

1. Which one of the following statements is wrong for gases?
(a) Gases do not have a definite shape and volume.
(b) Volume of the gas is equal to volume of container confining the gas.
(c) Confined gas exerts uniform pressure on the walls of its container in all directions.
(d) Gases are not compressible.

2. Which of the following is true about gaseous state?
(a) Thermal energy = molecular attraction
(b) Thermal energy >> molecular attraction
(c) Thermal energy << molecular attraction
(d) Molecular forces >> those in liquids

3. Intermolecular forces in solid hydrogen are
(a) covalent forces
(b) van der Waals forces or London dispersion forces
(c) hydrogen bonds
(d) all of these.

4. For the type of interactions; (I) covalent bond, (II) van der Waals forces, (III) hydrogen bonding, (IV) dipole-dipole interaction, which represents the correct order of increasing stability?
(a) (I) < (III) < (II) < (IV)
(b) (II) < (III) < (IV) < (I)
(c) (II) < (IV) < (III) < (I)
(d) (IV) < (II) < (III) < (I)

5. The types of attractive forces between a polar molecule and a non-polar molecule are
(a) dipole-dipole forces
(b) hydrogen bonds
(c) dipole-induced dipole forces
(d) dispersion forces.

6. 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

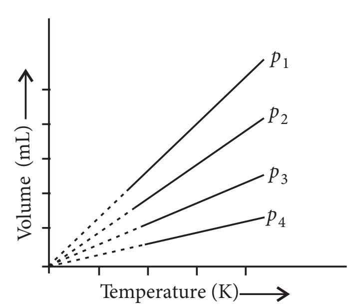
7. If P, V and T represent pressure, volume and temperature of the gas, the correct representation of Boyle's law is
(a) $V \propto 1/P$ (P is constant)
(b) $PV = RT$
(c) $V \propto 1/P$ (at constant T)
(d) $PV = nRT$

8. At 25°C and 380 mm pressure, 400 mL of dry oxygen was collected. If the temperature is constant, what volume will the oxygen occupy at 760 mm pressure?
(a) 200 mL (b) 400 mL
(c) 569 mL (d) 621 mL

9. Use of hot air balloons in sports and meteorological observations is an application of
(a) Boyle's law (b) Newton's law
(c) Kelvin's law (d) Charles' law.

10. Equal volumes of gases at the same temperature and pressure contain equal number of particles. This statement is direct consequence of
(a) perfect gas law
(b) partial law of volumes
(c) Charles' law
(d) ideal gas equation.

11. A plot of volume (V) versus temperature (T) for a gas at constant pressure is a straight line passing through the origin. The plots at different values of pressure are shown in the figure. Which of the following order of pressure is correct for this gas?



(a) $p_1 > p_2 > p_3 > p_4$ (b) $p_1 = p_2 = p_3 = p_4$
(c) $p_1 < p_2 < p_3 < p_4$ (d) $p_1 < p_2 = p_3 < p_4$

12. When the product of pressure and volume is plotted against pressure for a given amount of gas, the line obtained is
(a) parallel to x-axis
(b) parallel to y-axis
(c) linear with positive slope
(d) linear with negative slope.

13. Containers A and B have same gases. Pressure, volume and temperature of A are all twice as that of B, then the ratio of number of molecules of A and B are
(a) 1 : 2 (b) 2 : 1
(c) 1 : 4 (d) 4 : 1

14. A gas cylinder can withstand a pressure of 15 atm. The pressure of cylinder is measured 12 atm at 27°C. Upto which temperature limit the cylinder will not burst?
(a) 375°C (b) 102°C
(c) 33.75°C (d) 240°C

15. Select the correct statement. In the gas equation, $PV = nRT$
(a) n is the number of molecules of a gas
(b) n moles of the gas have a volume V
(c) V denotes volume of one mole of the gas
(d) P is the pressure of the gas when only one mole of gas is present.

16. Molar volume of CO_2 is maximum at
(a) NTP (b) 0°C and 2.0 atm
(c) 127°C and 1 atm (d) 273°C and 2.0 atm.

17. 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

18. Dimension of universal gas constant (R) is
(a) $[VPT^{-1}n^{-1}]$ (b) $[VP^{-1}Tn^{-1}]$
(c) $[VPTn^{-1}]$ (d) $[VPT^{-1}n]$

19. The mole fraction of dioxygen in a neon-dioxygen mixture is 0.18. If the total pressure of the mixture is 25 bar, the partial pressure of neon in the mixture would be
(a) 25.18 bar (b) 25.82 bar
(c) 4.5 bar (d) 20.5 bar

20. 25 g of each of the following gases are taken at 27°C and 600 mm pressure. Which of these will have the least volume?
(a) HBr (b) HCl
(c) HF (d) HI

21. At high pressure, the compressibility factor Z is equal to
(a) unity (b) $1-\frac{Pb}{RT}$
(c) $1+\frac{Pb}{RT}$ (d) zero.

22. Gas deviates from ideal gas nature because molecules
(a) are colourless
(b) attract each other
(c) contain covalent bond
(d) show Brownian movement.

23. van der Waals equation of state is obeyed by real gases. For n moles of a real gas, the expression will be
(a) $\left(\frac{P}{n}+\frac{na}{V^2}\right)\left(\frac{V}{n-b}\right) = RT$
(b) $\left(P+\frac{a}{V^2}\right)(V-b) = nRT$
(c) $\left(P+\frac{na}{V^2}\right)(nV-b) = nRT$
(d) $\left(P+\frac{n^2a}{V^2}\right)(V-nb) = nRT$

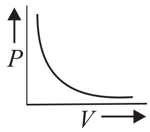
24. The van der Waals equation reduces itself to the ideal gas equation at
(a) high pressure and low temperature
(b) low pressure and low temperature
(c) low pressure and high temperature
(d) high pressure and high temperature.

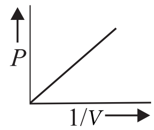
25. Maximum deviation from ideal gas is expected from
(a) $CH_{4(g)}$ (b) $NH_{3(g)}$
(c) $H_{2(g)}$ (d) $N_{2(g)}$

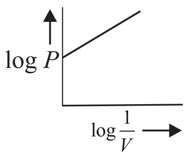
26. A gas such as carbon monoxide would be most likely to obey the ideal gas law at
(a) low temperatures and high pressures
(b) high temperatures and high pressures
(c) low temperatures and low pressures
(d) high temperatures and low pressures.

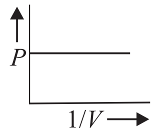
27. The correction factor 'a' to the ideal gas equation corresponds to
(a) density of the gas molecules
(b) volume of the gas molecules
(c) electric field present between the gas molecules
(d) forces of attraction between the gas molecules.

28. Which of the following plots is not according to Boyle's law?

(a) 

(b) 

(c) 

(d) 

29. Which of the following expressions does not represent Charles' law?

(a) $V_t = V_0\left[\frac{273.15 + t°C}{273.15}\right]$

(b) $V_t = a + bt$

(c) $V_t = \left[\frac{V_0}{273.15\text{ K}}\right]t$

(d) $V_t = V_0 t$

30. To raise the volume of a gas by four times, the following methods may be adopted. Which of the methods is wrong?

(a) T is doubled and P is also doubled.

(b) Keeping P constant, T is raised four times.

(c) Temperature is doubled and pressure is halved.

(d) Keeping temperature constant, pressure is reduced to 1/4 of its initial value.

Case Based MCQs

Case I : Read the passage given below and answer the following questions from 31 to 35.

Intermolecular forces are the forces of attraction and repulsion that exist between molecules of a compound. These cause the compound to exist in a certain state of matter – solid, liquid or gas and affect the melting and boiling points of compounds as well as the solubilities of one substance in another. Attractive intermolecular forces are also called van der Waals' forces. These are weak forces.

31. Dipole-dipole forces act between the molecules possessing permanent dipole. Ends of dipoles possess 'partial charges'. The partial charge is

(a) more than unit electronic charge

(b) equal to unit electronic charge

(c) less than unit electronic charge

(d) double the unit electronic charge.

32. The nature of inter-particle forces in benzene is

(a) dipole-dipole interaction

(b) dispersion force

(c) ion-dipole interaction

(d) H-bonding.

33. The interaction energy between two temporary dipoles is proportional to (where r is the distance between the two particles)

(a) $1/r^4$ (b) $1/r^2$

(c) $1/r^5$ (d) $1/r^6$

34. Attractive intermolecular forces known as van der Waals forces do not include which of the following types of interactions?

(a) London forces

(b) Dipole-dipole forces

(c) Ion-dipole forces

(d) Dipole-induced dipole forces

35. In which of the following molecules, the van der Waals forces are likely to be the most important in determining the m.pt. and b.pt?

(a) CO (b) H_2S

(c) Br_2 (d) HCl

Case II : Read the passage given below and answer the following questions from 36 to 40.

If a hydrogen atom is bonded to a highly electronegative element such as fluorine, oxygen, nitrogen, then the shared pair of electrons lies more towards the electronegative element. This leads to a polarity in the bond in such a way that a slight positive charge gets developed on H-atom, viz,

$$H^{\delta+} : O^{\delta-} \qquad H^{\delta+} : F^{\delta-} \qquad H^{\delta+} : N^{\delta-}$$

Such a bond between the hydrogen atom of one molecule and the more electronegative atom of the same or another molecule is called hydrogen bond.

36. Which of the following compounds can form hydrogen bond?

(a) CH_4 (b) H_2O

(c) NaCl (d) $CHCl_3$

37. The boiling point is not affected due to hydrogen bonding in

(a) water (b) ammonia

(c) methyl alcohol (d) hydrogen chloride.

38. Unusual high b.p. of water is result of

(a) intermolecular hydrogen bonding

(b) intramolecular hydrogen bonding

(c) both intra and intermolecular hydrogen bonding

(d) high specific heat.

39. Boiling point of hydrogen fluoride is highest amongst HF, HCl, HBr and HI. Which type of intermolecular forces are present in hydrogen fluoride?

(a) H—F has highest van der Waals forces and dipole moment.

(b) H—F has highest London forces.

(c) H—F has highest dipole moment hence has dipole-dipole, London forces and hydrogen bonding.

(d) H—F has strong intermolecular interactions like dipole-induced dipole.

40. Which of the following statements is not true?

(a) Intermolecular hydrogen bonds are formed between two different molecules of compounds.

(b) Intramolecular hydrogen bonds are formed between two different molecules of the same compound.

(c) Intramolecular hydrogen bonds are formed within the same molecule.

(d) Hydrogen bonds have strong influence on the physical properties of a compound.

Case III : Read the passage given below and answer the following questions from 41 to 45.

An ideal gas is a gas to which the laws of Boyle and Charles are strictly applicable under all conditions of temperatures and pressures.

From Boyle's law we get, $V \propto \frac{1}{P}$ (at constant n and T)

From Charles' law we get, $V \propto T$ (at constant n and P)

From Avogadro's law we get, $V \propto n$ (at constant T and P)

Combining the above three equations we get

$V \propto \frac{nT}{P}$ or, $V = R\frac{nT}{P}$ [where R = ideal gas constant]

or $PV = nRT$

Ideal gas equation is a relation between four variables and it describes the state of any gas. For this reason, it is also called equation of state.

41. 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) 365 mL

(b) 449 mL

(c) 569 mL

(d) 621 mL

42. 7.0 g of a gas at 300 K and 1 atm occupies a volume of 4.1 litre. What is the molecular mass of the gas?

(a) 42 (b) 38.24

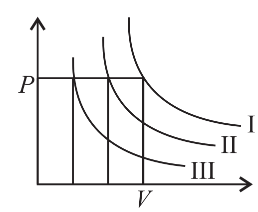
(c) 14.5 (d) 46.5

43. If P is the pressure and ρ is the density of a gas, then P and ρ are related as

(a) $P \propto \rho$ (b) $P \propto \rho^2$

(c) $P \propto 1/\rho$ (d) $P \propto 1/\rho^2$

44. I, II, III are three isotherms respectively at T_1, T_2 and T_3. Temperature will be in order



(a) $T_1 = T_2 = T_3$

(b) $T_1 < T_2 < T_3$

(c) $T_1 > T_2 > T_3$

(d) $T_1 > T_2 = T_3$

45. If volume of 2 moles of an ideal gas at 540 K is 44.8 litre, then its pressure will be

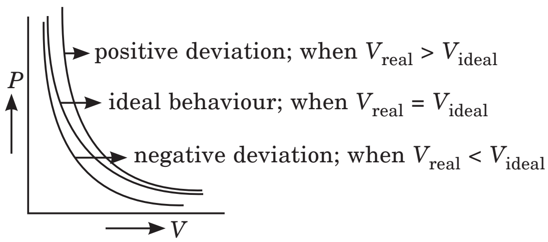
(a) 1 atm (b) 3 atm

(c) 2 atm (d) 4 atm

Case IV: Read the passage given below and answer the following questions from 46 to 48.

Real gases do not obey ideal gas equation under all conditions. They nearly obey ideal gas equation at higher temperatures and very low pressures. However, they show deviations from ideality at low temperatures and high pressures.

The isotherms obtained by plotting pressure, P against volume, V for real gases do not coincide with that of ideal gas, as shown :



V_{real} = Volume of the real gas at given pressure.
V_{ideal} = Volume of the gas calculated by ideal gas equation at given pressure.

The deviation from ideal gas behaviour can also be expressed by compressibility factor, Z.

46. The gas equation $PV = nZRT$ becomes ideal gas equation when
(a) $Z = 0$
(b) $Z = 0.5$
(c) $Z = 1$
(d) $Z = 2$

47. The units of van der Waals' constants *a* and *b* respectively are
(a) L atm^2 mol^{-1} and mol L^{-1}
(b) L atm mol^2 and mol L
(c) L^2 atm mol^{-2} and mol^{-1} L
(d) L^{-2} atm^{-1} mol^{-1} and L mol^{-2}

48. The correction factor 'b' to the ideal gas equation corresponds to
(a) density of the gas molecules
(b) excluded volume or covolume
(c) electric field present between the gas molecules
(d) forces of attraction between the gas molecules.

Assertion & Reasoning Based MCQs

For question numbers 49-55, a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

(a) Assertion and reason both are correct statements and reason is correct explanation for assertion.
(b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
(c) Assertion is correct statement but reason is wrong statement.
(d) Assertion is wrong statement but reason is correct statement.

49. Assertion : Three states of matter are the result of balance between intermolecular forces and thermal energy of the molecules.
Reason : Intermolecular forces tend to keep the molecules together but thermal energy of molecules tends to keep them apart.

50. Assertion : At constant temperature, PV *vs* V plot for real gases is not a straight line.
Reason : At high pressure all gases have $Z > 1$ but at intermediate pressure most gases have $Z < 1$.

51. Assertion : The plot of volume (V) versus pressure (P) at constant temperature is a hyperbola in the first quadrant.
Reason : $V \propto 1/P$ at constant temperature.

52. Assertion : Compressibility factor (Z) for for non ideal gases is always greater than 1.
Reason : The gases which lave $Z > 1$ are difficult to compress.

53. Assertion : van der Waals equation is applicable only to non-ideal gases.
Reason : Ideal gases obey the equation $PV = nRT$.

54. Assertion : Vapour pressure of NH_3 is higher than C_2H_5OH.
Reason : H-bonding is observed in both the molecules.

55. Assertion : The graph between P *v/s* $1/V$ is a straight line.
Reason : At constant temperature, $P \propto 1/V$.

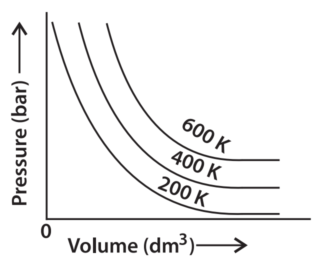
SUBJECTIVE TYPE QUESTIONS

Very Short Answer Type Questions (VSA)

1. Calculate the volume occupied by 4.0 mole of an ideal gas under NTP condition.

2. How is the partial pressure of a gas in a mixture related to the total pressure of the gaseous mixture?

3. The variation of pressure with volume of the gas at different temperatures can be graphically represented as shown in figure.



On the basis of this graph given, how will the volume of a gas change if its pressure is increased at constant temperature?

4. Under which of the following two conditions applied together, a gas deviates most from the ideal behaviour?

5. Physical properties of ice, water and steam are very different. What is the chemical composition of water in all the three states.

6. Name two intermolecular forces that exist between HF molecules in liquid state.

7. Define the partial pressure of gas.

8. Define compressibility factor.

9. Using the equation of state $PV = nRT$; show that at a given temperature, density of a gas is proportional to gas pressure P.

10. Why is boiling point of hydrogen fluoride higher than that of hydrogen chloride?

Short Answer Type Questions (SA-I)

11. 40 mL of O_2 was collected at 100 °C and 1 bar pressure. Calculate its volume (in mL) at 273 K and 1.013 bar.

12. 0.068 dm^3 of a sample of nitrogen is collected over water at 20 °C and 0.92 bar. What will be the volume of dry nitrogen at STP (in mL) (Aqueous tension of water at 20 °C = 0.023 bar)?

13. Calculate the volume (in m^3) occupied by 2 moles of an ideal gas at 25×10^5 Nm^{-2} pressure and 300 K temperature.

14. At 0°C, the density of a certain oxide of a gas at 2 bar is same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?

15. How much time (in years) would it take to distribute one Avogadro number of wheat grains, if 10^{10} grains are distributed each second?

16. Calculate the volume (in litres) occupied by 8.8 g of CO_2 at 31.1°C and 1 bar pressure.
(R = 0.083 bar L K^{-1} mol^{-1})

17. Which type of intermolecular forces exist among the following molecules?

(i) He atoms and HCl molecules

(ii) HF molecules

(iii) N_2 molecules

(iv) HCl molecules

18. Explain the physical significance of van der Waals' parameters.

19. Calculate the total pressure (in bar) in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of 1 dm^3 at 27°C.
R = 0.083 bar dm^3 K^{-1} mol^{-1}

20. 1 mole of sulphur dioxide occupies a volume of 350 mL at 27 °C and 5×10^6 Pa pressure. Calculate the compressibility factor of the gas. Is it less or more compressible than an ideal gas?

Short Answer Type Questions (SA-II)

21. Describe London forces or dispersion forces with example.

22. The drain cleaner, Drainex contains small bits of aluminium which react with caustic soda to produce dihydrogen. What volume of dihydrogen (in mL) at 20°C and one bar will be released when 0.15 g of aluminium reacts?

23. On the basis of intemolecular forces and thermal energy explain why substances exist in three different states of matter?

24. Two moles of ammonia gas are enclosed in a vessel of 5 litre capacity at 27ºC. Calculate the pressure exerted by the gas, assuming that

(i) the gas behaves like an ideal gas (using ideal gas equation)

(ii) the gas behaves like a real gas (using van der Waals equation)

Given that for ammonia, $a = 4.17$ atm litre^2 mol^{-2} and $b = 0.037$ litre mol^{-1}.

25. 2.9 g of a gas at 95 °C occupied the same volume as 0.184 g of dihydrogen at 17 °C at the same pressure. What is the molar mass of the gas?

26. On the basis of their interaction energy and thermal energy explain why

(i) a solid has high rigidity?

(ii) In gas, molecules are sufficiently apart from one another?

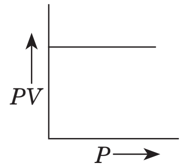
(iii) liquid has no definite shape?

27. Explain the following :

(i) Boyle's law

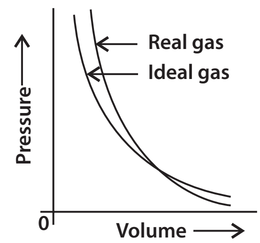
(ii) Avogadro's law

28. (a) Which gas law is shown by the following graph?



(b) At 25°C and 760 mm Hg pressure, a gas occupies 600 mL volume. What will be its pressure (in mmHg) at a height where temperature is 10°C and volume of the gas is 640 mL?

29. Pressure versus volume graph for a real gas and an ideal gas are shown in the figure.



Answer the following questions on the basis of this graph.

(i) Interpret the behaviour of real gas with respect to ideal gas at low pressure.

(ii) Interpret the behaviour of real gas with respect to ideal gas at high pressure.

(iii) Mark the pressure and volume by drawing a line at the point where real gas behaves as an ideal gas.

30. A perfectly elastic spherical balloon of 0.2 m diameter was filled with hydrogen at sea level. What will be its diameter (in m) when it has risen to an altitude where the pressure is 0.65 atm? (Assume no change in temperature and atmospheric pressure at sea level).

31. (i) Out of CO_2 and He, which gas have higher value of van der Waals' constant 'b'?

(ii) At 27°C density of a gaseous substance at 3 bar is same as that of hydrogen at 9 bar. What is molar mass of the substance?

32. What will be the pressure (in bar) of a gas mixture when 0.5 L of H_2 at 0.8 bar and at 2.0 L of oxygen at 0.7 bar are introduced in a 1 L vessel at 27 °C?

33. (a) 22 g of dry ice is placed in an evacuated bottle of 1 litre capacity and tightly stoppered. What would be the pressure inside the bottle in atm, when it is heated to 37°C?

(b) 3.12 g of sulphur is vapourised at 427°C and 760 mm pressure, when the vapours occupy a volume of 700 mL. Find the molecular formula of sulphur. (atomic mass of sulphur = 32).

34. An open beaker at 27°C is heated to 477°C. What percentage of air would have been expelled out?

35. Explain the difficulties faced by the mountaineers with respect to the air present around them. How is this difficulty solved?

Long Answer Type Questions (LA)

36. Payload is defined as the difference between the mass of displaced air and the mass of the balloon. Calculate the payload (in kilograms) when a balloon of radius 10 m, mass 100 kg is filled with helium at 1.66 bar at 27°C. (Density of air =1.2 kg m^{-3} and R = 0.083 bar dm^3 K^{-1} mol^{-1})

37. A liquefied petroleum gas (LPG) cylinder weighs 14.8 kg when empty. When full it weighs 29.0 kg and shows a pressure of 2.5 atm. In the course of use at 27 °C, the mass of the full cylinder is reduced to 23.2 kg. Find out the volume of the gas in cubic metres used up at the normal usage conditions and the final pressure inside the cylinder. Assume LPG to be n-butane with normal boiling point of 0 °C.

38. At temperature, T and pressure P, two ideal gases A and B are mixed. Show that the density d of the mixture is given by $d = (X_A M_A + X_B M_B) \times \frac{P}{RT}$ where X_A and X_B are the mole fractions and M_A and M_B are the molecular weights of the gases A and B respectively.

39. The volume of a gas is to be increased by 20% without changing the pressure. To what temperature (in °C) the gas must be heated if the initial temperature of the gas is 27 °C?

40. Answer the following :

(a) How are the van der Waals' constants 'a' and 'b' related to the molecular size?

(b) Using van der Waals' equation, calculate the constant 'a' when two moles of a gas confined in a 4 L flask exerts a pressure of 11.0 atmospheres at a temperature of 300 K. The value of 'b' is 0.05 L mol^{-1}.

ANSWERS

OBJECTIVE TYPE QUESTIONS

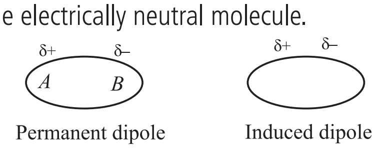
1. **(d) :** Gases are highly compressible.

2. **(b) :** For gaseous state, thermal energy >> molecular attraction.

3. **(b) :** Solid hydrogen, H_2 is non-polar compound and possesses London dispersion forces. Infact these are the only attractive forces which a non-polar compound can have.

4. **(c)**

5. **(c) :** It is the type of force between the polar molecule and a non-polar molecule. Dipole of polar molecule induces dipole on the electrically neutral molecule.



6. **(a) :** HCl is polar ($\mu \neq 0$) and He is non-polar ($\mu = 0$). Therefore, HCl and He atoms will posses dipole-induced dipole interaction.

7. **(c) :** Boyle's law relates pressure and volume of a gas at constant temperature *i.e.*, $V \propto \frac{1}{P}$ (at constant T).

8. **(a) :** Applying $P_1V_1 = P_2V_2$

P_1 = 380 mm, V_1 = 400 mL, P_2 = 760 mm, V_2 = ?

$$V_2 = \frac{P_1V_1}{P_2} = \frac{380 \times 400}{760} = 200 \text{ mL}$$

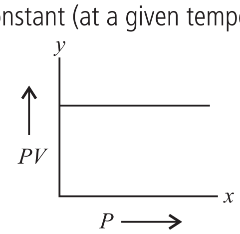
9. **(d) :** According to Charles' law $V \propto T$ *i.e.*, air expands on heating, its density decreases. Hence hot air is lighter.

10. **(d) :** Ideal gas equation $PV = nRT$

If P, V and T are same n will also be same.

11. **(c) :** Since $pV \propto T$ and $p \propto \frac{1}{V}$ hence, volume decreases with pressure. The values of p_1, p_2, p_3, p_4 show decrease in volume. Hence, the order of pressure is $p_1 < p_2 < p_3 < p_4$.

12. **(a) :** PV = constant (at a given temperature)



13. **(b) :** $P_AV_A = n_ART_A$ and $P_BV_B = n_BRT_B$

$$\frac{n_A}{n_B} = \frac{\frac{P_AV_A}{RT_A}}{\frac{P_BV_B}{RT_B}} = \frac{\frac{2P_B \times 2V_B}{2T_B}}{\frac{P_B \times V_B}{T_B}} = 2$$

Thus number of molecules are also in the ratio 2 : 1.

14. (b) : Cylinder will burst at that temperature when it attains the pressure of 15 atm

$P_1 = 12$ atm; $T_1 = 27°C = 27 + 273 = 300$ K; $P_2 = 15$ atm; $T_2 = ?$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\Rightarrow \; T_2 = \frac{15 \times 300}{12} = 375 \text{ K} = (375 - 273)°C = 102°C$$

15. (b) : From ideal gas equation, $PV = nRT$, it can be concluded that n moles of a gas occupy the volume V at pressure P and temperature T.

16. (c) : We know $V \propto \frac{1}{P}$ and $V \propto T$

$\therefore$ CO_2 has maximum volume at minimum pressure and maximum temperature.

17. (d) : $m = 1.8$ g

$$n = \frac{m}{M} = \frac{1.8}{18} = 0.1 \text{ mol}$$

$T = 374°C = 647$ K, $P = 1$ bar

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

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

18. (a) : From the gas equation, $PV = nRT$

$$R = \frac{P \times V}{n \times T} = [VPT^{-1}n^{-1}]$$

19. (d) : Given $P_{total} = 25$ bar, $x_{O_2} = 0.18$

$\therefore \quad x_{Ne} = 1 - 0.18 = 0.82$

$p_{Ne} = x_{Ne} \times P_{total} = 0.82 \times 25 = 20.5$ bar

20. (d) : Ideal gas equation is $PV = nRT$

If pressure and temperature are same for all the gases then

$V \propto n$ (from above equation)

$n =$ no. of moles of the gas $= \dfrac{\text{wt.}}{\text{Mol. wt.}}$

(here weight of all gases are equal)

$$\therefore \; V \propto \frac{1}{\text{Mol. wt.}}$$

From the options, HI has more molecular weight than remaining gases, so it has least volume.

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

At high pressure b cannot be neglected in comparison to V. Further though V becomes small, a/V^2 is large but as P is very high, a/V^2 can be neglected in comparison to P. Hence

$P(V - b) = RT$ or $PV = RT + Pb$

or $\quad \dfrac{PV}{RT} = 1 + \dfrac{Pb}{RT}$ *i.e.* $Z = 1 + \dfrac{Pb}{RT}$

22. (b) : Unlike postulates of kinetic molecular theory, intermolecular forces in a gas are not negligible.

23. (d) : van der Waal's equation is

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

24. (c)

25. (b) : It is a polar molecule, thus more attractive forces between its molecule.

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

27. (d)

28. (d) : $P \propto \frac{1}{V}$, the graph will be a straight line passing through origin.

29. (d) : $V_t = \dfrac{V_0}{t_0} \times t$

30. (a) : $\dfrac{PV}{T}$ = constant, only (a) is wrong.

31. (c) : Partial charge is a small charge developed by displacement of electrons. It is less than unit electronic charge and is represented as δ^+ or δ^-.

32. (b) : Benzene is non-polar compound and exhibits London or dispersion forces.

33. (d) **34. (c)**

35. (c) : In the molecules the van der Waals force are likely to determine the m.pt. and b.pt. Greater the mass of the molecule greater will be its van der Waals force and higher will be its m.pt. and b.pt. Br_2 has highest m.pt.

36. (b) : Oxygen has high electronegativity and small size, thus forms H-bond.

37. (d) : HCl does not undergo H-bonding and its boiling point is not affected by H-bonding.

38. (a) : Due to intermolecular hydrogen bond in H_2O, its molecules are associated with each other which is responsible for unusual high b.p. of water.

39. (c) : H—F has dipole-dipole interaction, London forces and hydrogen bonding due to highest electronegativity of F. Hence, boiling point of H—F is highest.

40. (b) : Intramolecular hydrogen bonds are formed within the same molecule.

41. (a) : Applying $P_1V_1 = P_2V_2$

$P_1 = 730$ mm, $V_1 = 380$ mL, $P_2 = 760$ mm, $V_2 = ?$

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

42. (a) : Applying, $PV = \frac{m}{M}RT$

$P = 1$ atm; $V = 4.1$ L; $m = 7.0$ g ;

$R = 0.0821$ L atm K^{-1} mol^{-1}; $T = 300$ K

$$M = \frac{RT}{PV} \times m = \frac{0.0821 \times 300}{1 \times 4.1} \times 7.0 = 42.05 \approx 42\, g\, mol^{-1}$$

43. (a) : $P = \frac{\rho RT}{M}$ *i.e.*, $P \propto \rho$

44. (c) : Draw a line at constant pressure parallel to volume-axis. Take volume corresponding to each temperature.

From volume axis, $V_1 > V_2 > V_3$

Hence $T_1 > T_2 > T_3$.

45. (c) : Number of moles, $n = 2$;

Temperature, $T = 540$ K

Volume, $V = 44.8$ L; $P = ?$

[R, gas constant = 0.0821 L atm K^{-1} mol^{-1}]

According to ideal gas equation $PV = nRT$

$$P = \frac{nRT}{V} = \frac{2 \times 0.0821 \times 540}{44.8} = 2\,\text{atm}$$

46. (c) : The ideal gas equation is $PV = nRT$

When Z (the compressibility factor) is one, the given equation $PV = nZRT$ becomes ideal gas equation.

47. (c) : a and b are expressed in terms of the units of P and V.

$$\text{Pressure correction} = P' = \frac{n^2 a}{V^2}$$

$$a = \frac{P'V^2}{n^2} = \frac{\text{pressure} \times (\text{volume})^2}{(\text{mole})^2}$$

Unit of a = atm × $(L)^2 mol^{-2}$

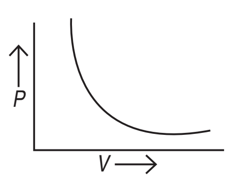
Unit of b is the same as for the volume. *i.e.*, L mol^{-1}

48. (b) : The correction factor 'b' represents excluded volume or covolume.

49. (a)

50. (b) : At constant temperature, plot of PV *vs* V for real gases is not linear because real gases have intermolecular forces of attraction.

51. (a) : According to Charles's law $V \propto \frac{1}{P}$.

∴ 

52. (d) : Compressibility factor (Z) can be greater, equal or lesser than 1.

53. (b) : The van der Waals equation is applicable to real gases only, while $PV = nRT$ is applicable to ideal gases.

54. (b) : H - bonding in liquid NH_3 is weaker than C_2H_5OH. Thus, escaping tendency is higher and hence, the vapour pressure of NH_3 is also higher.

55. (a)

SUBJECTIVE TYPE QUESTIONS

1. $PV = nRT$

or $V = \frac{nRT}{P} = \frac{4 \times 0.082 \times 273}{1} = 89.54$ L

2. Partial pressure of a gas

= Mole fraction of that gas × Total pressure

3. The volume of a gas decreases if the pressure on the gas is increased keeping the temperature constant.

4. At high pressure and low temperature, the gases deviate from ideal behaviour due to significant intermolecular forces and more than negligible size of the molecules.

5. The chemical composition of water remains same in all the physical states *i.e.*, solid, liquid and gas.

6. Dipole-dipole interactions and hydrogen bonding exist between HF molecules in liquid state.

7. In a mixture of gases, the pressure exerted by the individual gas is called its partial pressure.

8. The extent to which a real gas deviates from ideal behaviour can be conveniently studied in terms of quantity 'Z' called the compressibility factor, which is defined as $Z = \frac{PV}{nRT}$

For an ideal gas, as $PV = nRT$, $Z = 1$

9. $PV = nRT$

or $P = \frac{wRT}{MV}$ $\left(\text{Since } n = \frac{w}{M}\right)$

or $P = \frac{dRT}{M}$ $\left[\text{Since } d = \frac{w}{V}\right]$

or, $P \propto d$

Hence, density (d) of a gas ∝ P, because R, T and M are constants.

10. Boiling point of HF is higher than HCl due to extensive hydrogen bonding between H — F molecules.

H — F ········ H — F ········ H — F ········ H — F

11. Given, $P_1 = 1$ bar, $V_1 = 40$ mL = 0.04 L = 0.04 dm^3,

$T_1 = 373$ K, $P_2 = 1.013$ bar, $T_2 = 273$ K, $V_2 = ?$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

or $V_2 = \frac{P_1V_1T_2}{T_1P_2} = \frac{1 \times 0.04 \times 273}{373 \times 1.013}$ = 0.029 dm^3 or 29 mL

12. Given P_1 = 0.92 – 0.023 = 0.897 bar,
V_1 = 0.068 dm^3, T_1 = 293 K, P_2 = 1 bar (Pressure of dry gas), V_2 = ?, T_2 = 273 K

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{or} \quad V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

$$= \frac{0.897 \times 0.068 \times 273}{1 \times 293} = 0.057 \text{dm}^3 \text{ or } 57 \text{ mL}$$

13. According to ideal gas equation, $PV = nRT$

or $V = \frac{nRT}{P} = \frac{2 \times 8.314 \times 300}{25 \times 10^5} \approx 2 \times 10^{-3}$ m^3

14. $M_{oxide} = \frac{d_{oxide}RT}{P_{oxide}}$ ...(1)

$M_{N_2} = \frac{d_{N_2}RT}{P_{N_2}}$ ...(2)

Dividing equation (1) by equation (2) gives

$$\frac{M_{oxide}}{M_{N_2}} = \frac{d_{oxide} \times P_{N_2}}{d_{N_2} \times P_{oxide}}$$

But $d_{oxide} = d_{N_2}$

$\therefore$ $M_{oxide} = \frac{5\,\text{bar}}{2\,\text{bar}} \times 28 \text{ g mol}^{-1}$

M_{oxide} = 70 g mol^{-1}

15. Time taken to distribute 10^{10} grains = 1 sec.
Time taken to distribute 6.023 × 10^{23} grains

$$= \frac{1 \times 6.022 \times 10^{23}}{10^{10}} = 6.022 \times 10^{13} \text{ sec}$$

$$= \frac{6.022 \times 10^{13}}{60 \times 60 \times 24 \times 365} = 1.90956 \times 10^6 \text{ years}$$

16. According to ideal gas equation, $PV = nRT$

$n = \frac{8.8}{44}$ moles, $P = 1$bar,

T = 273 + 31.1 = 304.1 K,

V = ?

R = 0.083 bar L K^{-1} mol^{-1}

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

$$= \frac{\left(\frac{8.8}{44}\text{ moles}\right) \times (0.083 \text{ bar L K}^{-1} \text{ mol}^{-1}) \times (304.1\text{K})}{1\text{bar}}$$

= 5.05 L

17. (i) Dipole-induced dipole forces
(ii) Hydrogen bonding
(iii) Dispersion forces
(iv) Dipole-dipole forces

18. '*a*' measures the intermolecular forces of attraction. The greater the value of '*a*', the more will be intermolecular forces of attraction. '*b*' measures volume occupied by molecules of a gas.

19. Partial pressure of oxygen gas,

$P = \frac{nRT}{V}$, $n = \frac{8}{32}$ mol, V=1 dm^3, T = 300 K

$$P_{(O_2)} = \frac{8 \times 0.083 \times 300}{32 \times 1} = 6.225 \text{ bar}$$

Partial pressure of hydrogen gas,

$P = \frac{nRT}{V}$, $n = \frac{4}{2} = 2$ mol

$$P_{(H_2)} = \frac{2 \times 0.083 \times 300}{1} = 49.8 \text{ bar}$$

Total pressure $= P_{(O_2)} + P_{(H_2)}$ = 6.225 + 49.8
= 56.025 bar

20. Compressibility factor, $Z = \frac{PV}{nRT}$

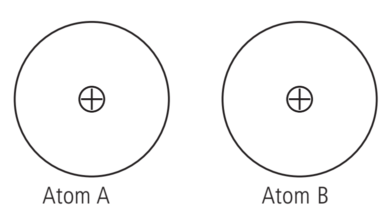
n = 1 mol, P = 5 × 10^6 Pa, V = 350 mL = 0.350 × 10^{-3} m^3
R = 8.314 N m K^{-1} mol^{-1}, T = 27 + 273 = 300 K

$$\therefore \quad Z = \frac{5 \times 10^6 \times 0.350 \times 10^{-3}}{1.0 \times 8.314 \times 300} = 0.702$$

Thus, SO_2 is more compressible than an ideal gas (which has Z = 1).

21. London or dispersion forces : This is the weakest intermolecular force. It is a temporary attractive force that results when the electrons in two adjacent atoms occupy positions that make the atoms to form temporary dipoles. This force is sometimes called an *dipole-induced dipole attraction.*

Because of the constant motion of the electrons, an atom or molecule can develop a temporary (instantaneous) dipole when its electrons are distributed unsymmetrically about the nucleus.



symmetrical distribution of electronic charge cloud
(a)

Atom 'A' with instantaneous dipole, more electron density on the right hand side

Atom 'B' with induced dipole,

(b)

Atom 'A' more electron density on the left hand side

Atom 'B' with induced dipole

(c)

Dispersion forces or London forces between atoms.

22. The reaction between aluminium and caustic soda is

$$\underset{\substack{2\times 27\\ =54\text{ g}}}{2Al} + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + \underset{\substack{3\times 22.4\text{ L}\\ \text{at STP}}}{3H_2}$$

$\therefore$ 54 g of Al produces H_2 at S.T.P. = 3 × 22.4 L

0.15 g of Al will produce H_2 at S.T.P.

$$= \frac{3\times 22.4}{54}\times 0.15 = 0.186\text{ L}$$

At STP	Given conditions
$P_1 = 1$ atm	$P_2 = 1$ bar $= 0.987$ atm
$V_1 = 0.186$ L	$V_2 = ?$
$T_1 = 273$ K	$T_2 = 273 + 20 = 293$ K

Applying ideal gas equation, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$$\frac{1\times 0.186}{273} = \frac{0.987\times V_2}{293}$$

$$V_2 = \frac{293}{0.987}\times\frac{1\times 0.186}{273} = 0.2030\text{ L} = 203\text{ mL}$$

23. The intermolecular forces tend to keep the molecules together but thermal energy tend to keep them apart. Thus, these two compete and the competition between these two (i.e., intermolecular forces and thermal energy) results in three states of matter.

(i) In a solid, the intermolecular forces predominate over the thermal energy and hence, the particles are held together in rigid, highly-oriented and close-packed structure.

(ii) In liquids, the intermolecular forces are no longer strong enough, however, these are still sufficient so that particles remain in each other's environment, hence, liquids have sufficient mobility.

(iii) In gases, the thermal energy dominates the effect of intermolecular forces, thus, the gas molecules acquire the unrestricted and independent mobility in the vapour state.

Predominance of thermal energy and the intermolecular forces in the three state of matter is as follows :

Gas $\longrightarrow$ Liquid $\longrightarrow$ Solid

Predominance of intermolecular interactions

Gas $\longleftarrow$ Liquid $\longleftarrow$ Solid

Predominance of thermal energy

24. Given, $n = 2$ moles, $V = 5$ litres, $T = 27°C = (27 + 273)$ K = 300 K

$a = 4.17$ atm litre^2 mol^{-2}, b = 0.037 litre mol^{-1}

Also, we know that R = 0.0821 litre atm K^{-1} mol^{-1}

(i) If the gas behaves like an ideal gas, we have $PV = nRT$

$$\therefore \quad P = \frac{nRT}{V} = \frac{2\times 0.0821\times 300}{5} = 9.85\text{ atm}$$

(ii) If the gas behaves like a real gas, we apply van der Waals' equation *i.e.*,

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT \text{ or } P = \frac{nRT}{V - nb} - \frac{an^2}{V^2}$$

$$= \frac{2\times 0.0821\times 300}{5 - 2\times 0.037} - \frac{4.17\times(2)^2}{(5)^2} = 9.33\text{ atm.}$$

25. Case I : Let molar mass of gas be M g mol^{-1}

Weight of gas = 2.9 g

$$\text{No. of moles} = \frac{\text{Weight}}{\text{Molar mass}} = \frac{2.9}{M}$$

$T = 273 + 95 = 368$ K

$$PV = \frac{2.9}{M}\times R\times 368 \quad \text{...(i)}$$

Case II : Mass of dihydrogen = 0.184 g

$$\text{No. of moles of } H_2 = \frac{0.184}{2}$$

$T = 273 + 17 = 290$ K

$$PV = \frac{0.184}{2}\times R\times 290 \quad \text{...(ii)}$$

From equations (i) and (ii),

$$\frac{2.9}{M}\times 368 = \frac{0.184}{2}\times 290$$

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

26. (i) A solid has high rigidity because thermal motion is too weak to overcome the strong intermolecular forces of attraction.

(ii) In a gas, thermal energy is so high that the molecules cannot come close together. Hence, there are large empty spaces between them.

(iii) In a liquid, there is a reasonable balance between the attractive intermolecular forces and thermal energy. Hence, molecules in a liquid exist together, *i.e.*, it is a condensed state of matter but there is no rigidity. That is why they have no definite shape.

27. (i) Boyle's law : According to Boyle's law, at constant temperature, the volume of a fixed amount of a gas is inversely proportional to its pressure, *i.e.*, if volume increases, the pressure would decrease. This is because as volume increases, the number of molecules striking the walls of a container in a given time decreases leading to decrease in pressure.

$$P \propto \frac{1}{V}\ ; P = \frac{\text{Constant}}{V}$$

$\Rightarrow \quad PV = \text{Constant} \Rightarrow P_1V_1 = P_2V_2$

(ii) Avogadro's law : This law states that equal volume of all gases under similar conditions of temperature and pressure contain equal number of molecules.

$$V \propto n\ (T, P \text{ constant}) = \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

28. (a) Boyle's law

(b) Applying gas equation (combined gas law),

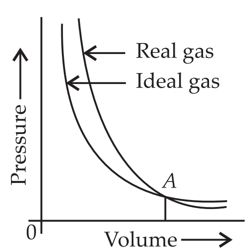
$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \Rightarrow \frac{760 \times 600}{298} = \frac{P_2 \times 640}{283}$$

or $\quad P_2 = 676.6$ mm of Hg

29. (i) At low pressure real gas starts behaving like an ideal gas.

(ii) At high pressure gases deviate from ideal behaviour.

(iii)



At point *A* real gas behaves as an ideal gas.

30. If r_1 is the radius of the balloon at sea level, then, volume of the balloon at sea level $= \frac{4}{3}\pi r_1^3 = \frac{4}{3}\pi(0.1\,\text{m})^3$

i.e., volume of the gas at sea level $(V_1) = \frac{4}{3}\pi(0.1\,\text{m})^3$

Pressure at the sea level $(P_1) = 1$ atm

Suppose the radius of the balloon at altitude $= r_2$

Then, volume of the balloon at altitude $(V_2) = \frac{4}{3}\pi r_2^3$

Pressure at the altitude $(P_2) = 0.65$ atm (Given)

As temperature remains constant, applying Boyle's law,

$$\underset{\text{(At sea level)}}{P_1V_1} = \underset{\text{(At altitude)}}{P_2V_2}$$

$$1\,\text{atm} \times \frac{4}{3}\pi(0.1\,\text{m})^3 = 0.65\,\text{atm} \times \frac{4}{3}\pi r_2^3$$

or $\quad r_2^3 = \dfrac{(0.1\,\text{m})^3}{0.65} = 1.54 \times 10^{-3}\ \text{m}^3$

$\therefore \quad r_2 = (1.54 \times 10^{-3})^{1/3}$ m $= 0.1154$ m

$\therefore \quad$ Diameter of the balloon at altitude

$= 2 \times 0.1154$ m $= 0.2308$ m

31. (i) Since, CO_2 molecules have larger size than that of He molecules, hence, CO_2 has larger value of van der Waals' constant 'b'.

(ii) $\quad M_1 = \dfrac{d_1RT}{P_1}$ ...(i)

[where,

M_1 = Molecular weight of the substance

d_1 = Density of the substance

P_1 = Pressure of the substance]

$M_{H_2} = \dfrac{d_{H_2}RT}{P_{H_2}}$ ...(ii)

Dividing equation (i) by equation (ii)

$$\frac{M_1}{M_{H_2}} = \frac{d_1 \times P_{H_2}}{d_{H_2} \times P_1} \text{ or } \frac{M_1}{M_{H_2}} = \frac{P_{H_2}}{P_1} \text{ [Given, } d_1 = d_{H_2}]$$

or $\quad M_1 = \dfrac{M_{H_2} \times P_{H_2}}{P_1} = \dfrac{2 \times 9}{3} = 6$

32. Partial pressure of H_2 :

$V_1 = 0.5$ L, $V_2 = 1$ L,

$P_1 = 0.8$ bar, $P_2 = ?$

By Boyle's law, $P_1V_1 = P_2V_2$

$$P_2 = \frac{0.8 \times 0.5}{1} = 0.4 \text{ bar}$$

Partial pressure of O_2 :

$V_1 = 2.0$ L, $V_2 = 1$ L, $P_1 = 0.7$ bar, $P_2 = ?$

By Boyle's law, $P_1V_1 = P_2V_2$

$$P_2 = \frac{0.7 \times 2.0}{1} = 1.4 \text{ bar}$$

Pressure of the gas mixture,

$P_{mix} = p_{H_2} + p_{O_2} = 0.4 + 1.4 = 1.8$ bar

33. (a) $W = 22$ g CO_2, $V = 1$ L, $M = 44$, $T = 37 + 273 = 310$ K, $P = ?$

Dry ice is solid CO_2, which when heated in an evacuated bottle it is converted into gaseous CO_2.

From ideal gas equation,

$$PV = \frac{WRT}{M} \Rightarrow P = \frac{22 \times 0.082 \times 310}{44 \times 1} = 12.71 \text{ atm.}$$

Now, pressure inside the bottle is 12.71 atm.

(b) For sulphur,

$W = 3.12$ g, $T = 427 + 273 = 700$ K

$P = 760$ mm $= 1$ atm.

$V = 700$ ml $= 0.7$ L.

Now molecular mass of sulphur,

$$M = \frac{WRT}{PV} = \frac{3.12 \times 0.082 \times 700}{1 \times 0.7} = 255.84$$

As atomic mass = 32, so no. of atoms in one molecule of sulphur $= \frac{255.84}{32} = 8$

Hence, molecular formula of sulphur is S_8.

34. Suppose the number of moles of gas present at 27°C in flask of volume V at pressure P is n_1, then assuming ideal gas behaviour,

$PV = n_1 R \times 300$ ...(i)

Suppose n_2 = number of moles at 477°C, then

$PV = n_2 R \times 750$ ...(ii)

From equation (i) and equation (ii), we get

$$n_2 = \frac{300}{750} \times n_1 = 0.4\, n_1$$

So, (1 – 0.4) × 100% *i.e.* 60% is expelled out.

35. At altitude, the atmospheric pressure is low. Hence, air is less dense. As a result, less oxygen is available for breathing. The person feels uneasiness, headache, etc. This is called altitude sickness. This difficulty is solved by carrying oxygen cylinders with them.

36. Volume of balloon $= \frac{4}{3}\pi r^3$

Radius of balloon, $r = 10$ m

$$V = \frac{4}{3} \times 3.14 \times (10)^3 = 4186.7 \text{ m}^3$$

Mass of displaced air = 4186.7 m^3 × 1.2 kg m^{-3}

= 5024.04 kg

$$\text{Moles of gas present} = \frac{PV}{RT}$$

$$= \frac{1.66 \times 4186.7 \times 10^3}{0.083 \times 300} = 279.11 \times 10^3 \text{ moles}$$

Mass of helium present = $279.11 \times 10^3 \times 4$

= 1116.44×10^3 g

= 1116.44 kg

Mass of filled balloon = 100 + 1116.44

= 1216.44 kg

Payload = Mass of displaced air – Mass of balloon

= 5024.04 – 1216.44 = 3807.6 kg

37. Weight of LPG originally present

= 29.0 – 14.8 = 14.2 kg

Pressure = 2.5 atm

Weight of LPG present after use = 23.2 – 14.8

= 8.4 kg

Since volume of the cylinder is constant, applying

$$PV = nRT \Rightarrow \frac{P_1}{P_2} = \frac{n_1}{n_2} = \frac{w_1/M}{w_2/M} = \frac{w_1}{w_2}$$

$$\Rightarrow \quad \frac{2.5}{P_2} = \frac{14.2}{8.4} \text{ or } P_2 = \frac{2.5 \times 8.4}{14.2} = 1.48 \text{ atm}$$

∵ Weight of used gas = 14.2 –8.4 = 5.8 kg

$$\text{Moles of gas} = \frac{5.8 \times 10^3}{58} = 100 \text{ mol}$$

Normal conditions : $P = 1$ atm, $T = 273 + 27 = 300$ K

Volume of 100 mol of LPG at 1 atm and 300 K

$$V = \frac{nRT}{P} = \frac{100 \times 0.082 \times 300}{1} = 2460 \text{ L} = 2.460 \text{ m}^3$$

38. At temperature T and pressure P, two ideal gases, A and B are mixed.

n_A and n_B are the number of moles of A and B respectively.

n is total moles of A and B present in the mixture.

$n = n_A + n_B$.

Let the gas mixture has a volume V.

V_A and V_B are volume of A and B respectively.

From ideal gas equation, $PV = nRT$.

For gas A, $PV_A = n_A RT$. ∴ $V_A = \frac{n_A RT}{P}$

For gas B, $PV_B = n_B RT$. ∴ $V_B = \frac{n_B RT}{P}$

M_A and M_B are molecular weights of the gases and X_A and X_B are the mole fractions of the gases A and B respectively.

Density of a gas, $d = \frac{\text{mass}}{\text{volume}}$

$$d = \frac{M}{V} = \frac{n_A M_A + n_B M_B}{\frac{n_A RT}{P} + \frac{n_B RT}{P}} = \frac{P}{RT}\left\{\frac{n_A M_A + n_B M_B}{n_A + n_B}\right\}$$

$$= \frac{P}{RT}\left\{\frac{n_A M_A + n_B M_B}{n}\right\} = \frac{P}{RT}\left\{\frac{n_A}{n}M_A + \frac{n_B}{n}M_B\right\}$$

$$d = \frac{P}{RT}\{X_A M_A + X_B M_B\}$$

39. Suppose volume of gas at 27°C = V cm^3

Increase in volume desired = 20% of $V = \frac{20}{100} \times V = 0.2$ V

$\therefore$ Final volume = V + 0.2 V = 1.2 V

Now, $V_1 = V$ cm^3, T_1 = 300 K,

V_2 = 1.2 V, T_2 = ?

At constant P,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{V}{300} = \frac{1.2\text{ V}}{T_2}$$

$\Rightarrow$ T_2 = 360 K = 360 – 273 = 87°C

40. (a) The van der Waals' constant 'a' is measure of intermolecular attractions. Therefore, the value of 'a' reflects the tendency of the gas to liquefy. The gas having larger value of 'a', will liquefy more easily. The van der Waals' constant 'b' is a measure of the close-packed molecular volume. Thus the molecule of a gas having greater value of 'b' has bigger size.

(b) The van der Waals' equation is

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

Given, n = 2 mol, V = 4 L,

P = 11.0 atm, T = 300 K

Substituting the values in the above equation,

$$\left(11.0\text{ atm} + \left(\frac{2\text{ mol}}{4\text{ L}}\right)^2 \cdot a\right)(4\text{ L} - 2\text{ mol} \times 0.05\text{ L mol}^{-1})$$

$$= 2\text{ mol} \times 0.082\text{ L atm K}^{-1}\text{ mol}^{-1} \times 300\text{ K}$$

$$\left(11.0\text{ atm} + \frac{a}{4}\text{ mol}^2\text{ L}^{-2}\right)(4\text{ L} - 0.1\text{ L}) = 49.2\text{ L atm}$$

$$\left(11.0\text{ atm} + \frac{a}{4}\text{ mol}^2\text{ L}^{-2}\right)(3.9\text{ L}) = 49.2\text{ L atm}$$

$$42.9\text{ L atm} + \frac{3.9a}{4}\text{ mol}^2\text{ L}^{-1} = 49.2\text{ L atm}$$

$$\frac{3.9}{4}a\text{ mol}^2\text{ L}^{-1} = (49.2\text{ L atm} - 42.9\text{ L atm})$$

$$= 6.3\text{ L atm}$$

$$a = \frac{6.3\text{ L atm} \times 4}{3.9\text{ mol}^2\text{ L}^{-1}} = 6.46\text{ atm L}^2\text{ mol}^{-2}$$