5 States of Matter

Facts that Matter

• Intermolecular Forces

Intermolecular forces are the forces of attraction and repulsion between interacting particles (atoms and molecules).

van der Waals Forces: Attractive intermolecular forces are known as van der Waals forces.

Dipole-Dipole Forces: It is the force of attraction which exists in dipolar molecules which have permanent dipole moments. This interaction is stronger than the London forces but is weaker than ion-ion interaction because only partial charges are involved.

The attractive forces decrease with the increase of distance between dipoles. The interaction energy is proportional to $\frac{1}{r^6}$ where r is the distance between polar molecules.

Ion-Dipole Interaction: This is the force of attraction which exists between the ions (cations or anions) and polar molecules. The ion is attracted towards the oppositely charged end of dipolar molecules.

The strength of attraction depends upon the charge and size of the ion and the dipole moment and the size of the polar molecule.

For example: Solubility of common salt (NaCl) in water.

• Ion-induced Dipolar Interactions

In this type of interaction permanent dipole of the polar molecule induces dipole on the electrically neutral molecule by deforming its electronic cloud. Interaction energy is proportional to $\frac{1}{r^6}$ where r is the distance between two molecules.

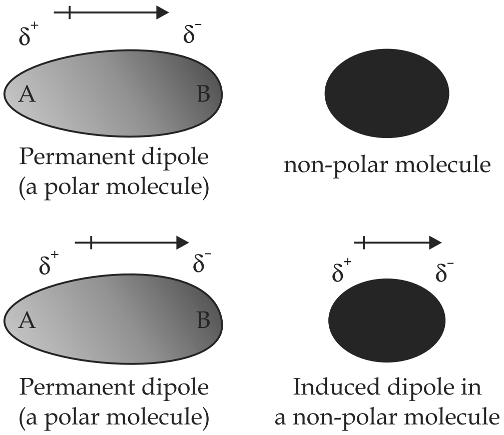
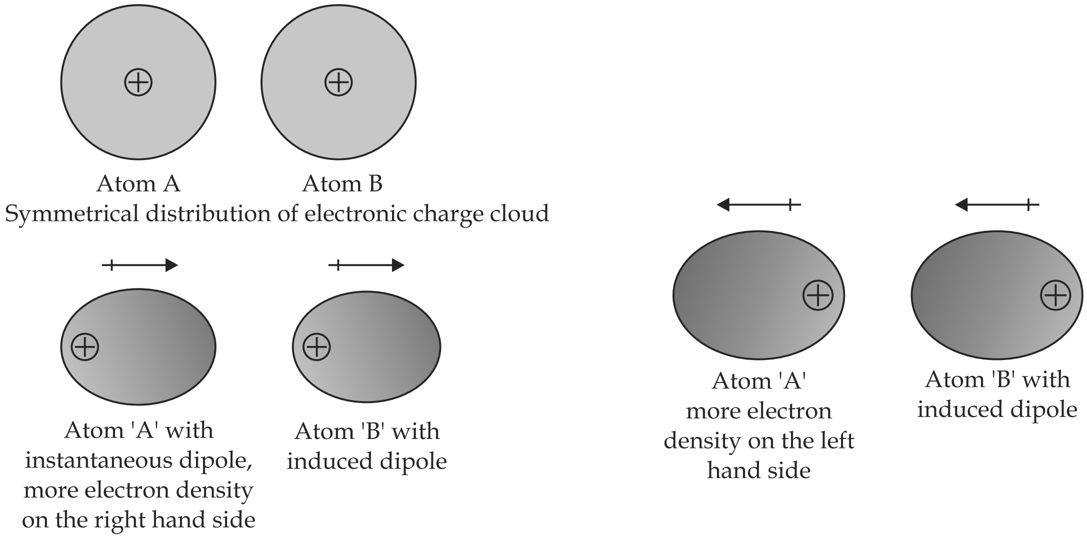


Fig. 5.1 *Dipole-induced dipole interaction between permanent dipole and induced dipole.*

• London Forces or Dispersion Forces

As we know that in non-polar molecules, there is no dipole moment because their electronic charge cloud is symmetrically distributed. But, it is believed that at any instant of time, the electron cloud of the molecule may be distorted so that an instantaneous dipole or momentary dipole is produced in which one part of the molecule is slightly more negative than the other part. This momentary dipole induces dipoles in the neighbouring molecules. Thus, the force of attraction exists between them and are exactly same as between permanent dipoles. This force of attraction is known as London forces or Dispersion forces. These forces are always attractive and the interaction energy is inversely proportional to the sixth power of the distance between two interacting particles. (i.*e.* $\frac{1}{r^6}$ where r is the distance between two particles).

This can be shown by fig. given below.



Atom A
Atom B
Symmetrical distribution of electronic charge cloud

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

Atom 'B' with induced dipole

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

Atom 'B' with induced dipole

Fig. 5.2 *Dispersion forces or London forces between atoms.*

Hydrogen bonding: When hydrogen atom is attached to highly electronegative element by covalent bond, electrons are shifted towards the more electronegative atom. Thus a partial positive charge develops on the hydrogen atom. Now, the positively charged hydrogen atom of one molecule may attract the negatively charged atom of some other molecule and the two molecules can be linked together through a weak force of attraction.

$$---H^{\delta+}---F^{\delta-}---H^{\delta+}---F^{\delta-}---H^{\delta+}---F^{\delta-}$$

Thermal Energy: The energy arising due to molecular motion of the body is known as thermal energy. Since motion of the molecules is directly related to kinetic energy and kinetic energy is directly proportional to the temperature.

• The Gaseous State

Physical Properties of Gaseous State

(*i*) ases have no definite volume and they do not have specific shape.
(*ii*) Gases mix evenly and completely in all proportions without any mechanical aid.
(*iii*) Their density is much lower than solids and liquids.
(*iv*) They are highly compressible and exert pressure equally in all directions.

• Boyle's Law (Pressure-Volume Relationship)

At constant temperature, the volume of a given mass of gas is inversely proportional to its pressure.

$$V \propto \frac{1}{P} \quad \text{(At constant temp.)}$$

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

or $$PV = K \quad \text{(constant)}$$

Here, V is the volume, P is the pressure.

K = constant of proportionality.

The value of constant K, depends upon the amount of gas, the temperature of the gas and the units in which P and V are expressed.

Let V_1 be the volume of the gas at a pressure P_1. By keeping the temperature constant, if the pressure is increased to P_2 then the volume will decrease to V_2. According to the law,

$$P_1V_1 = P_2V_2 = \text{constant.}$$

or $$\frac{P_1}{P_2} = \frac{V_2}{V_1}$$

Charles' law: At constant pressure, the volume of a given mass of a gas is directly proportional to its absolute temperature.

$$V \propto T$$

or $$\frac{V}{T} = \text{constant.}$$

Let V_1 be the volume of a gas at temperature T_1. Pressure remains constant.

If the temperature of the gas increased to T_2, then the volume will also increase to V_2. Therefore, according to law,

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

• Gay Lussac's Law (Pressure-Temperature Relationship)

At constant volume, pressure of a given mass of a gas is directly proportional to the temperature.

Mathematically,

$$P \propto T$$

$$\Rightarrow \quad \frac{P}{T} = \text{constant. (At constant volume and mass)}$$

Pressure *Vs.* temperature (Kelvin) graph at constant molar volume is shown in figure.

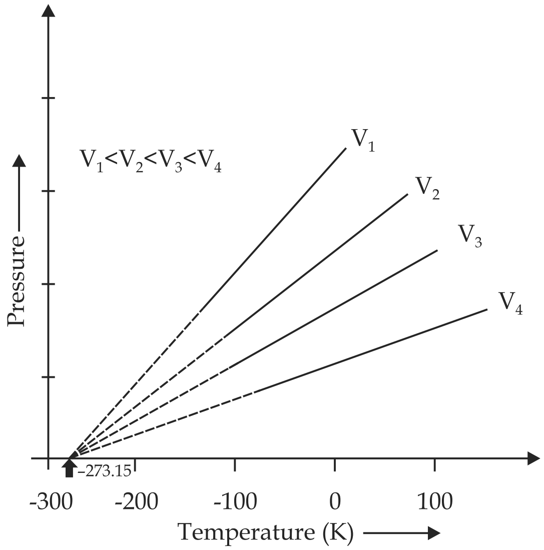


Fig. 5.3 *Pressure vs temperature (K) graph (Isochores) of a gas.*

Isochore: Each line of graph is called isochore.

• Avogadro Law (Volume-Amount Relationship)

Avogadro's law states that equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules.

$$V \propto n$$

Where n is the number of moles of the gas.

Avogadro constant: The number of molecules in one mole of a gas

$$= 6.022 \times 10^{23}$$

Ideal Gas: A gas that follows Boyle's law, Charles' law and Avogadro law strictly, is called an ideal gas.

Real gases follow these laws only under certain specific conditions. When forces of interaction are practically negligible.

• Ideal Gas Equation

This is the combined gas equation of three laws and is known as ideal gas equation.

At constant T and n : $V \propto \frac{1}{P}$ Boyle's law

At constant P and n : $V \propto T$ Charles' law

At constant P and T : $V \propto n$ Avogadro law.

$$V \propto \frac{nT}{P}$$

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

or

$$\boxed{PV = nRT}$$

R is called gas constant. It is also called Universal Gas Constant.

$$\boxed{R = \frac{PV}{nT}} \qquad R = \frac{1\,\text{atom} \times 22400\text{cm}^3}{1\,\text{mol} \times 273K}$$

$$= 82.1\ \text{cm}^3\ \text{atm K}^{-1}\ \text{mol}^{-1}$$

• Dalton's Law of Partial Pressure

When two or more non-reactive gases are enclosed in a vessel, the total pressure exerted by the gaseous mixture is equal to the sum of the partial pressure of individual gases.
Let P_1, P_2, and P_3 be the pressure of three non reactive gases *A*, *B*, and *C*. When enclosed separately in the same volume and under same condition.

$$P_{\text{total}} = P_1 + P_2 + P_3$$

Where, $P_{\text{total}} = P$ is the total pressure exerted by the mixture of gases.

• Aqueous Tension

Pressure of non reacting gases are generally collected over water and therefore are moist. Pressure of dry gas can be calculated by substracting vapour pressure of water from total pressure of moist gas.

$$P_{\text{Dry gas}} = \text{P}_{\text{Total}} - \text{Aqueous Tension}$$

• Partial Pressure in terms of Mole Fraction

Let at the temperature *T*, three gases enclosed in the volume *V*, exert partial pressure P_1, P_2 and P_3 respectively, then

$$P_1 = \frac{n_1 RT}{V}$$

$$P_2 = \frac{n_2 RT}{V}$$

$$P_3 = n_3 \frac{RT}{V}$$

$$P_{\text{total}} = P_1 + P_2 + P_3$$

$$= (n_1 + n_2 + n_3) \frac{RT}{V}$$

$$\frac{P_1}{P_{\text{total}}} = \left(\frac{n_1}{n_1 + n_2 + n_3}\right) \frac{RTV}{RTV}$$

$$= \frac{n_1}{n} = x_1$$

Where $(n = n_1 + n_2 + n_3)$

$\therefore$ $P_1 = n_1\, P_{\text{total}}.$

• Kinetic Molecular Theory of Gases

(*i*) Gases consist of large number of very small identical particles (atoms or molecules).

(*ii*) Actual volume occupied by the gas molecule is negligible in comparison to empty space between them.

(*iii*) Gases can occupy all the space available to them. This means they do not have any force of attraction between their particles.

(*iv*) Particles of a gas are always in constant random motion.

(*v*) When the particles of a gas are in random motion, pressure is exerted by the gas due to collision of the particles with the walls of the container.

(*vi*) Collision of the gas molecules are perfectly elastic. This means there is no loss of energy after collision. There may be only exchange of energy between colliding molecules.

(*vii*) At a particular temperature distribution of speed between gaseous particles remains constant.

(*viii*) Average kinetic energy of the gaseous molecule is directly proportional to the absolute temperature.

• Deviation From Ideal Gas Behaviour

Real Gas: A gas which does not follow ideal gas behaviour under all conditions of temperature and pressure, is called real gas.

Deviation with respect to pressure can be studied by plotting pressure *Vs* volume curve at a given temperature. (Boyle's law)

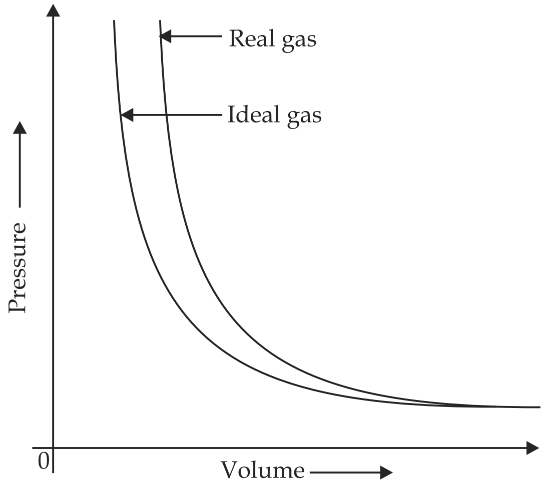


Fig. 5.4 *Plot of pressure vs volume for real gas and ideal gas.*

Compressibility factor (Z): Deviation from ideal behaviour can be measured in terms of compressibility factor, *Z*.

$$Z = \frac{PV}{nRT}$$

For ideal gas $Z = 1$

For real gas $Z \neq 1$

When $Z > 1$ Real gases show positive deviation from ideality

$Z < 1$ The gas show more compressibility and deviation is regarded as negative deviation.

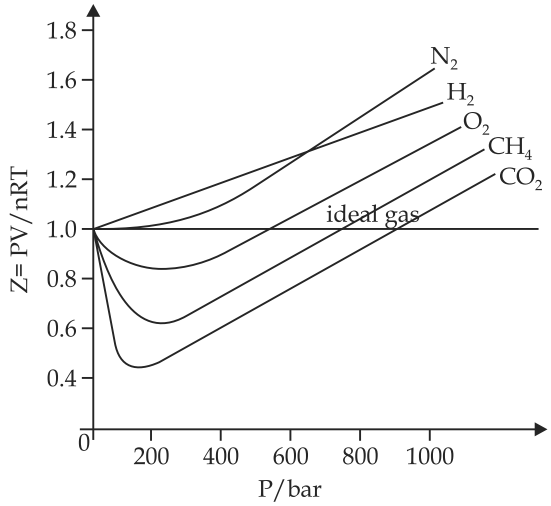


Fig. 5.5 *Variation of compressibility factor for some gases.*

• van der Waals Equation

For one mole of a gas $\left(P+\frac{a}{V^2}\right)(V-b) = RT$

Where '*a*' is a constant for molecular attraction while '*b*' is a constant for molecular volume.

(*a*) There is no force of attraction between the molecules of a gas.

(*b*) Volume occupied by the gas molecule is negligible in comparison to the total volume of the gas.

Above two assumptions of the kinetic theory of gas was found to be wrong at very high pressure and low temperature.

• Liquifaction of Gases

Liquifaction of gases can be achieved either by lowering the temperature or increasing the pressure of the gas simultaneously.

Thomas Andrews plotted isotherms of CO_2 at various temperatures shown in figure.

Critical Temperature (T_C): It is defined as that temperature above which a gas cannot be liquified however high pressure may be applied on the gas.

$$T_c = \frac{8a}{27bR}$$

(Where *a* and *b* are van der Waals constants)

Critical Pressure (P_C): It is the pressure required to Liquify the gas at the critical temperature.

$$P_c = \frac{a}{27b^2}$$

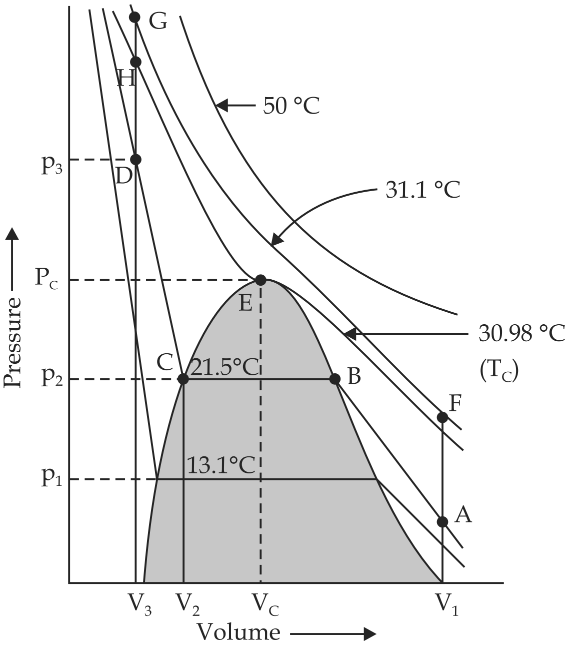


Fig. 5.6 *Isotherms of carbon dioxide at various temperatures*

The volume occupied by one mole of the gas at the critical temperature and the critical pressure is called the critical volume (V_c).

For Example. For CO_2 to Liquify.

$$T_c = 30.98°C$$
$$P_c = 73.9 \text{ atm.}$$
$$V_c = 95–6 \text{ cm}^3/\text{mole}$$

All the three are collectively called critical constants.

• Liquid State

Characteristics of Liquid State

(*i*) In liquid, intermolecular forces are strong in comparison to gas.
(*ii*) They have definite volume but irregular shapes or we can say that they can take the shape of the container.
(*iii*) Molecules of liquids are held together by attractive intermolecular forces.

Vapour Pressure: The pressure exerted by the vapour of a liquid, at a particular temperature in a state of dynamic equilibrium, is called the vapour pressure of that liquid at that temperature.

Vapour Pressure depends upon two factors:

(*i*) Nature of Liquid (*ii*) Temperature

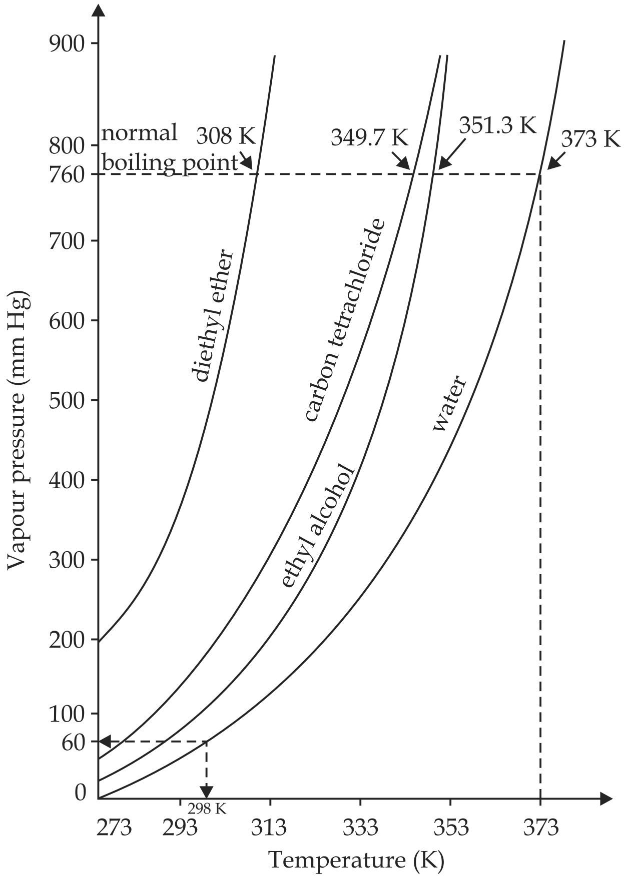


Fig. 5.7 *Vapour pressure vs temperature curve of some common liquids.*

- **Surface Tension**

It is defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid.

S.I. unit of Surface Tension = Nm^{-1}.

Surface Tension decreases with increase in temperature, because force acting per unit length decreases due to increase in kinetic energy of molecules.

- **Viscosity**

It is defined as the internal resistance to flow possessed by a liquid.

The liquids which flow slowly have very high internal resistance, which is due to strong intermolecular forces and hence are said to be more viscous.

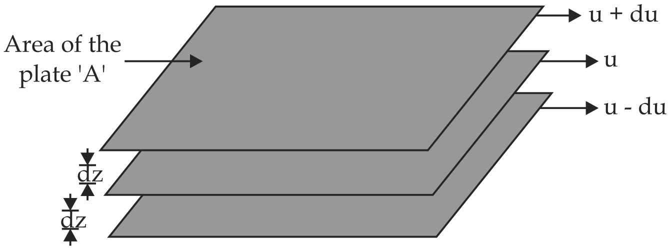


Fig. 5.8 *Gradation of velocity in the laminar flow.*

When liquid flows, the layer immediately below it tries to retard its flow while the one above tries to accelerate.

Thus, force is required to maintain the flow of layers.

This force

$$f \propto A \quad \text{(A is area of contact)}$$

$$f \propto \frac{du}{dz} \left(\frac{du}{dz} \text{ is velocity gradient } i.e., \text{ the change in velocity with distance.}\right)$$

$$f \propto A\frac{du}{dz}$$

$$f = \eta A\frac{du}{dz}$$

Where 'η' is proportionality constant and is called coefficient of viscosity.

SI unit of viscosity coefficient = 1 newton second per square metre

$= Nsm^{-2}$ = Pascal Second.

In C.G.S system, the unit of coefficient of viscosity is poise.

$$1 \text{ poise} = 1 g\ cm^{-1}\ s^{-1}.$$

Effect of Temp. on Viscosity: Viscosity of liquids decreases as the temperature rises because at high temperature, molecules have high kinetic energy and can overcome the intermolecular forces to slip past one another.

Words that Matter

- **Boyle's Law:** It states that, under isothermal conditions pressure of a given mass of a gas is inversely proportional to its volume.

$$P \propto \frac{1}{V}$$

or $PV = \text{constant}$

or $P_1V_1 = P_2V_2$

- **Charles' Law:** It states that, pressure remaining constant, volume of a fixed amount of a gas is directely proportional to its absolute temperature.

$$V \propto T$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ (At constant pressure)}$$

- **Avogadro law:** It states that, equal volumes of all gases under same conditions of temperature and pressure contain equal number of molecules.
- **Ideal gas equation:**

$$PV = nRT$$

Where R = Universal Constant.

- **Dalton's law of partial pressures:** It states that, total pressure exerted by a mixture of non-reacting gases is equal to the sum of partial pressures exerted by them.

$$P = P_1 + P_2 + P_3 \ldots\ldots$$

- **van der Waals equation**

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

- Critical Temperature: $(T_c) = \dfrac{8a}{27\,Rb}$

Critical pressure $(P_c) = \dfrac{a}{27b^2}$

Critical volume $V_c = 3b$

NCERT TEXTBOOK QUESTIONS SOLVED

Q1. *What will be the minimum pressure required to compress 500 dm³ of air at 1 bar to 200 dm³ at 30°C?*

Sol. $P_1 = 1$ bar, $V_1 = 500$ dm³

$P_2 = ?$ $V_2 = 200$ dm³

As temperature remains constant at 30°C,

$$P_1V_1 = P_2V_2$$

$$1 \text{ bar} \times 500 \text{ dm}^3 = P_2 \times 200 \text{ dm}^3 \quad \text{or} \quad P_2 = \frac{500}{200}\text{bar} = \mathbf{2.5\ bar}$$

Q2. *A vessel of 120 mL capacity contains a certain amount of gas at 35°C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 mL at 35°C. What would be its pressure?*

Sol. $V_1 = 120$ mL, $P_1 = 1.2$ bar, $T_1 = 35°C$

$V_2 = 180$ mL, $P_2 = ?$ $T_2 = 35°C$

As temperature remains constant, $P_1V_1 = P_2V_2$

$$(1.2 \text{ bar})(120 \text{ mL}) = P_2 (180\text{mL}) \quad \text{or} \quad P_2 = \mathbf{0.8\ bar}$$

Q3. *Using the equation of state PV = nRT, show that at a given temperature, density of a gas is proportional to the gas pressure P.*

Sol. According to ideal gas equation

$$PV = nRT$$

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

$$n = \frac{\text{Constant Mass of gas}}{\text{Molar mass of gas}}$$

$$P = \frac{mRT}{MV} \qquad \left[\because \rho\,(density) = \frac{m}{V}\right]$$

$$P = \frac{\rho RT}{M}$$

$P \propto \rho$ at constant temperature

Q4. *At 0°C, the density of a gaseous oxide at 2 bar is same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?*

Sol. Using the expression, $d = \frac{MP}{RT}$, at the same temperature and for same density,

$$\underset{\text{(Gaseous oxide)}}{M_1P_1} = \underset{(N_2)}{M_2P_2} \quad \text{(as } R \text{ is constant)}$$

$\therefore$ $M_1 \times 2 = 28 \times 5$ (Molecular mass of N_2 = 28 u)

or $\mathbf{M_1 = 70u}$

Q5. *Pressure of 1 g of an ideal gas A at 27°C is found to be 2 bar. When 2 g of another ideal gas B is introduced in the same flask at same temperature, the pressure becomes 3 bar. Find the relationship between their molecular masses.*

Sol. Suppose molecular masses of A and B are M_A and M_B respectively. Then their number of moles will be

$$n_A = \frac{1}{M_A}, \qquad n_B = \frac{2}{M_B}$$

P_A = 2 bar, $P_A + P_B$ = 3 bar, *i.e.*, P_B = 1 bar

Applying the relation $PV = nRT$

$$P_AV = n_ART, \qquad P_BV = n_BRT \quad \therefore \quad \frac{P_A}{P_B} = \frac{n_A}{n_B} = \frac{1/M_A}{2/M_B} = \frac{M_B}{2M_A}$$

or $$\frac{M_B}{M_A} = 2 \times \frac{P_A}{P_B} = 2 \times \frac{2}{1} = 4 \quad \text{or} \quad \mathbf{M_B = 4\,M_A}.$$

Q6. *The drain cleaner, Drainex contains small bits of aluminium which react with caustic soda to produce dihydrogen. What volume of dihydrogen at 20 °C and one bar will be released when 0.15g of aluminium reacts?*

Sol. The chemical equation for the reaction is

$2\,Al + 2\,NaOH + H_2O \rightarrow 2\,NaAlO_2 + 3H_2$ (3 × 22400 mL At N.T.P)

$2 \times 27 = 54$ g.

54 g of Al at N.T.P release

H_2 gas = 3×22400

0.15 g of Al at N.T.P release

$$H_2 \text{ gas} = \frac{3 \times 22400 \times 0.15}{54} = 186.7 \text{ mL}$$

N.T.P condition.

V_1 = 186.7 mL $\qquad V_2$ = ?

P_1 = 1.013 bar $\qquad P_2$ = 1 bar

T_1 = 0 + 273 = 273 K $\qquad T_2$ = 20 + 273 = 293 K

According to Gas equation

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

$$V_2 = \frac{1.013 \text{ bar} \times 186.7 \text{mL} \times 293\text{K}}{1 \text{ bar} \times 273\text{K}}$$

$$= \mathbf{203\ mL}$$

Q7. *What will be the pressure exerted by a mixture of 3.2g of methane and 4.4g of carbon dioxide contained in a 9 dm³ flask at 27 °C?*

Sol.

$$p = \frac{n}{V}RT = \frac{m}{M}\frac{RT}{V}$$

$$p_{CH_4} = \left(\frac{3.2}{16}\text{mol}\right)\frac{0.0821 \text{ dm}^3\text{atm K}^{-1}\text{mol}^{-1} \times 300 \text{ K}}{9 \text{ dm}^3} = 0.55 \text{ atm}$$

$$p_{CO_2} = \left(\frac{4.4}{44}\text{mol}\right)\frac{0.0821 \text{ dm}^3\text{atm K}^{-1}\text{mol}^{-1} \times 300 \text{ K}}{9 \text{ dm}^3} = 0.27 \text{ atm}$$

$$p_{\text{total}} = 0.55 + 0.27 = \mathbf{0.82\ atm}$$

In terms of SI units, $R = 8.314\ p\ m^3\ \text{K}^{-1}\ \text{mol}^{-1}$, $V = 9 \times 10^{-3}\ \text{m}^3$

$P = 5.543 \times 10^4 \text{ Pa} + 2.771 \times 10^4 \text{ Pa} = \mathbf{8.314 \times 10^4\ Pa.}$

Q8. *What will be the pressure of the gas mixture when 0.5 L of H_2 at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1 L vessel at 27 °C?*

Sol. *Calculation of partial pressure of H_2 in 1L vessel*

$P_1 = 0.8$ bar, $V_1 = 0.5$ L

$P_2 = ?$, $V_2 = 1.0$ L

As temperature remains constant, $P_1V_1 = P_2V_2$

$(0.8 \text{ bar})\ (0.5 \text{ L}) = P_2\ (1.0 \text{ L})$ or $P_2 = 0.40$ bar, *i.e.,* $p_{H_2} = 0.40$ bar

Calculation of partial pressure of O_2 in 1 L vessel

$P'_1V'_1 = P'_2V'_2$

$(0.7 \text{ bar})\ (2.0 \text{ L}) = P'_2\ (1 \text{ L})$ or $P'_2 = 1.4$ bar, *i.e.,* $p_{O_2} = 1.4$ bar

∴ Total pressure $= P_{H_2} + P_{O_2} = 0.4 \text{ bar} + 1.4 \text{ bar} = \mathbf{1.8\ bar}$

Q9. *Density of a gas is found to be 5.46 g/dm³ at 27 °C and at 2 bar pressure. What will be its density at STP?*

Sol. $d = \dfrac{MP}{RT}$. For the same gas at different temperatures and pressures, $\dfrac{d_1}{d_2} = \dfrac{P_1}{T_1} \times \dfrac{T_2}{P_2}$.

Here, $d_1 = 5.46 \text{ g dm}^{-3}$, $T_1 = 27\ °\text{C} = 300$ K, $P_1 = 2$ bar.

At STP, $d_2 = ?$, $T_2 = 0\ °\text{C} = 273$ K, $P_2 = 1$ bar

$$\therefore \quad \frac{5.46\text{g dm}^{-3}}{d_2} = \frac{2 \text{ bar}}{300 \text{ K}} \times \frac{273\ K}{1 \text{ bar}} \text{ or } d_2 = \mathbf{3\ g\ dm^{-3}}$$

Q10. *34.05 mL of phosphorus vapour weighs 0.0625g at 546°C and 1.0 bar pressure. What is the molar mass of phosphorus?*

Sol. Step I. Calculation of volume at 0°C and 1 bar pressure

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \text{ i.e., } \frac{1 \times 34.05}{546+273} = \frac{1 \times V_2}{273} \text{ or } V_2 = 11.35 \text{ mL}$$

11.35 mL of vapour at 0°C and 1 bar pressure weigh = 0.0625 g

∴ 22700 mL of vapour at 0°C and 1 bar pressure will weigh

$$= \frac{0.0625}{11.35} \times 22700 = 125 \text{ g}$$

∴ Molar mass = $\mathbf{125\ g\ mol^{-1}}$

Alternatively, using

$$R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$$

$$PV = nRT, \text{ i.e., } n = \frac{PV}{RT} = \frac{1.0 \text{ bar} \times (34.05 \times 10^{-3} \text{ dm}^3)}{0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ Mol}^{-1} \times 819 \text{ K}}$$

$$= 5 \times 10^{-4} \text{ mol}$$

∴ Mass of 1 mole $= \frac{0.0625}{5 \times 10^{-4}} g = 125$ g

∴ Molar mass = $\mathbf{125\ g\ mol^{-1}}$

Q11. *A student forgot to add the reaction mixture to the round bottomed flask at 27 °C but instead, he/she placed the flask on the flame. After a lapse of time, he realized his mistake, and using a pyrometer, he found the temperature of the flask was 477 °C. What fraction of air would have been expelled out?*

Sol. Suppose volume of vessel = V cm^3

i.e., volume of air in the flask at 27°C = V cm^3.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}, \text{ i.e., } \frac{V}{300} = \frac{V_2}{750} \text{ or } V_2 = 2.5\,V$$

∴ Volume expelled = $2.5\,V - V = 1.5\,V$

∴ Fraction of air expelled = $\frac{1.5\,V}{2.5\,V} = \mathbf{\frac{3}{5}}$

Q12. *Calculate the temperature of 4.0 moles of a gas occupying 5 dm^3 at 3.32 bar (R = 0.083 bar dm^3 K^{-1} mol^{-1})*

Sol. $PV = nRT$ or $T = \frac{PV}{nR} = \frac{3.32 \text{ bar} \times 5 \text{ dm}^3}{4.0 \text{ mol} \times 0.083 \text{ bar dm}^3 \text{ K}^{-1}\text{mol}^{-1}}$ **= 50K**

Q13. *Calculate the total number of electrons present in 1.4 g of dinitrogen gas.*

Sol. Molecular mass of N_2 = 28*g*

28 g of N_2 has No. of molecules = 6.022×10^{23}

1.4 g of N_2 has No. of molecules = $6.022 \times 10^{23} \times 1.4$ g/28 g

= 3.011×10^{22} molecules.

Atomic No. of Nitrogen (N) = 7

1 molecule of N_2 has electrons = $7 \times 2 = 14$

3.011×10^{22} molecules of N_2 have electrons

$$= 14 \times 3.011 \times 10^{22}$$

$$= \mathbf{4.215 \times 10^{23}\ electrons.}$$

Q14. *How much time would it take to distribute one Avogadro number of wheat grains if 10^{10} grains are distributed each second ?*

Sol. Time taken to distribute 10^{10}grains = 1s

Time taken to distribute = 6.022×10^{23} grains

$$= \frac{1s \times 6.022 \times 10^{23}\,\text{grains}}{10^{10}\ \text{grains}}$$

$$= \frac{6.022 \times 10^{13}}{60 \times 60 \times 24 \times 365} = \mathbf{1.9 \times 10^{6}\ yr}.$$

Q15. *Calculate the total pressure in a mixture of 8g of oxygen and 4g of hydrogen confined in a vessel of 1 dm^3 at 27°C. $R = 0.083$ bar dm^3 K^{-1} mol^{-1}.*

Sol. Molar mass of O_2 = 32 g mol^{-1} $\therefore$ 8 g $O_2 = \frac{8}{32}$ mol = 0.25 mol

Molar mass of H_2 = 2 g mol^{-1} $\therefore$ 4 g $H_2 = \frac{4}{2}$ = 2 mol

$\therefore$ Total number of moles (n) = 2 + 0.25 = 2.25

V = 1 dm^3, T = 27°C = 300 K, R = 0.083 bar dm^3 K^{-1} mol^{-1}

$$PV = nRT$$

or

$$P = \frac{nRT}{V} = \frac{(2.25\ \text{mol})\ (0.083\ \text{bar dm}^3\text{K}^{-1}\text{mol}^{-1})\ (300\ \text{K})}{1\ \text{dm}^3}$$

$$= \mathbf{56.025\ bar}$$

Q16. *Pay load is defined as the difference between the mass of the displaced air and the mass of the balloon. Calculate the pay load 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}).*

Sol. Radius of the balloon = 10 m

$\therefore$ Volume of the balloon $= \frac{4}{3}\pi r^3 = \frac{4}{3} \times \frac{22}{7} \times (10\ \text{m})^3 = 4190.5\ \text{m}^3$

Volume of He filled at 1.66 bar and 27 °C = 4190.5 m^3

Calculation of mass of He

$$PV = n\text{R}T = \frac{w}{M}\text{R}T$$

or

$$w = \frac{MPV}{RT} = \frac{(4 \times 10^{-3}\,\text{kg mol}^{-1})\ (1.66\ \text{bar})\ (4190.5 \times 10^{3}\ \text{dm}^3)}{(0.083\ \text{bar dm}^3\ \text{K}^{-1}\ \text{mol}^{-1})\ (300\ \text{K})}$$

$$= 1117.5\ \text{kg}$$

Total mass of the balloon alongwith He = 100 + 1117.5 = 1217.5 kg

Maximum mass of the air that can be displaced by balloon to go up = Volume × Density

$$= 4190.5\ \text{m}^3 \times 1.2\ \text{kg m}^{-3} = 5028.6\ \text{kg}$$

$\therefore$ Pay load = 5028.6 – 1217.5 kg = **3811.1 kg**

Q17. *Calculate the volume occupied by 8.8 g of CO_2 at 31.1 °C and 1 bar pressure. R = 0.083 bar LK^{-1} mol^{-1}.*

Sol.
$$\text{No. of moles of } CO_2\ (n) = \frac{\text{Mass of } CO_2}{\text{Molar mass}} = \frac{8.8\text{ g}}{44\text{g mol}^{-1}} = 0.2\text{ mol}$$

$$\begin{aligned}\text{Pressure of } CO_2\ (P) &= 1\text{ bar}\\ R &= 0.083\text{ bar LK}^{-1}\text{ mol}^{-1}\\ \text{Temperature } (T) &= 273 + 31.1\\ &= 304.1\text{ K}\\ \text{Since from gas eq. } PV &= nRT\\ V &= \frac{nRT}{P} = \frac{0.2\times 0.083\times 304.1}{1\text{ bar}}\\ &= \mathbf{5.048\ L}\end{aligned}$$

Q18. *2.9 g of a gas at 95°C occupied the same volume as 0.184 g of hydrogen at 17°C at the same pressure. What is the molar mass of the gas?*

Sol. As $P_1 = P_2$ and $V_1 = V_2$,

$\therefore$ $P_1V_2 = P_2V_2$, *i.e.,* $n_1RT_1 = n_2RT_2$ $\therefore$ $n_1T_1 = n_2T_2$

or $\frac{w_1}{M_1}T_1 = \frac{w_2}{M_2}T_2$

$$\frac{2.9}{M_x}\times(95+273) = \frac{0.184}{2}\times(17+273) \quad\text{or}\quad M_x = \frac{2.9\times 368\times 2}{0.184\times 290} = \mathbf{40\ g\ mol^{-1}}$$

Q19. *A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.*

Sol. As the mixture H_2 and O_2 contains 20% by weight of dihydrogen, therefore, if H_2 = 20g, then O_2 = 80g

$$n_{H_2} = \frac{20}{2} = 10\text{ moles},\quad n_{O_2} = \frac{80}{32} = 2.5\text{ moles}$$

$$p_{H_2} = \frac{n_{H_2}}{n_{H_2}+n_{O_2}}\times P_{\text{total}} = \frac{10}{10+2.5}\times 1\text{ bar} = \mathbf{0.8\ bar}$$

Q20. *What would be the SI unit for the quantity pV^2T^2/n?*

Sol.
$$\frac{(Nm^{-2})(m^3)^2(K)^2}{mol} = \mathbf{Nm^4\ K^2\ mol^{-1}}$$

Q21. *In terms of Charles' law explain why –273°C is the lowest possible temperature.*

Sol. At –273°C, volume of the gas becomes equal to zero, *i.e.,* the gas ceases to exist.

Q22. *Critical temperature for CO_2 and CH_4 are 31.1°C and –81.9°C respectively. Which of these has stronger intermolecular forces and why ?*

Sol. Higher the critical temperature, more easily the gas can be liquefied, *i.e.,* greater are the intermolecular forces of attraction. Hence, CO_2 has stronger intermolecular forces than CH_4.

Q23. *Explain the physical significance of van der Waals parameters.*

Sol. '*a*' is a measure of the magnitude of the intermolecular forces of attraction, while *b* is a measure of the effective size of the gas molecules.

MORE QUESTIONS SOLVED

I. VERY SHORT ANSWER TYPE QUESTIONS

Q1. *What is the value of the gas constant in SI units?*

Ans. 8.314 $JK^{-1}\ mol^{-1}$.

Q2. *Define boiling point of a liquid.*

Ans. The temperature at which the vapour pressure of a liquid is equal to external pressure is called boiling point of liquid.

Q3. *What is SI unit of (i) Viscosity (ii) Surface tension?*

Ans. (*i*) Unit of viscosity is Nsm^{-2}

(*ii*) Unit of surface tension is Nm^{-1}

Q4. *What is the effect of temperature on (i) surface tension and (ii) Viscosity?*

Ans. (*i*) Surface tension decreases with increase of temperature.

(*ii*) Viscosity decreases with increase of temperature.

Q5. *What is the unit of coefficient of viscosity?*

Ans. Poise.

Q6. *What do you understand by laminar flow of a liquid?*

Ans. The type of flow in which there is regular gradation of velocity in passing from one layer to the next is called laminar flow.

Q7. *What do you mean by compressibility factor?*

Ans. The deviation from ideal behaviour can be measured in terms of compressibility factor Z.

$$Z = \frac{PV}{nRT}$$

Q8. *What is Boyle Temperature?*

Ans. The temperature at which a real gas obeys ideal gas law over an appreciable range of pressure, is called Boyle temperature or Boyle point.

Q9. *What is meant by elastic collision ?*

Ans. Collision in which there is no loss of kinetic energy but there is transfer of energy, is called elastic collision.

Q10. *Define critical temperature of gas.*

Ans. The temperature above which a gas cannot be liquefied.

Q11. *What are real gases?*

Ans. A gas which can deviate from ideal gas behaviour at higher pressure and lower temperature, is called a real gas.

Q12. *Define an ideal gas.*

Ans. A gas that follows Boyle's law, Charles' law and Avogadro law strictly, is called an ideal gas.

Q13. *Name four properties of gases.*

Ans. (*i*) Gases, have no definite shape and no definite volume.

(*ii*) There is no force of attraction existing between the molecules of gases.

(*iii*) Gases are highly compressible.

(*iv*) Gases can mix evenly and can spread in whole space.

Q14. *State Dalton's law of partial pressure.*

Ans. Daltons' Law states that, total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of individual gases.

Q15. *What do you mean by aqueous tension?*

Ans. Pressure exerted by saturated water vapour is called aqueous tension.

Q16. *Give mathematical expression for ideal gas equation.*

Ans.
$$PV = nRT$$

Where R is called Gas constant.

Q17. *Write van der Waals equation for n moles of a gas.*

Ans.
$$\left[P + \frac{an^2}{V^2}\right](V - nb) = n\mathrm{R}T$$

Where '*a*' and '*b*' are van der waals constants.

Q18. *How is compressibility factor expressed in terms of molar volume of the real gas and that of the ideal gas?*

Ans.
$$\mathrm{Z} = \frac{V_{\text{real}}}{V_{\text{ideal}}}$$

Q19. *Why liquids diffuse slowly as compared to gases?*

Ans. In liquids, the molecules are more compact in comparison to gases.

Q20. *What is the effect of temperatures on the vapour pressure of a liquid?*

Ans. Vapour pressure increases with rise in temperature.

Q21. *Why falling liquid drops are spherical?*

Ans. Because of the property of surface tension, liquid tends to minimise its area.

II. SHORT ANSWER TYPE QUESTIONS

Q1. *A weather balloon has a volume of 175 dm^3 when filled with hydrogen gas at a pressure of 1.0 bar. Calculate the volume of the balloon when it rises to a height where the atmospheric pressure is 0.8 bar. Assume that temperature is constant.*

Ans. From the given data: P_1 = 1 bar, P_2 = 0.8 bar

V_1 = 175 dm^3, V_2 = ?

Since the temperature is constant, Boyle's Law can be applied

$$P_1V_1 = P_2V_2 \quad \text{or} \quad V_2 = \frac{P_1V_1}{P_2}$$

or
$$V_2 = \frac{(1\ \text{bar}) \times (175\ \text{dm}^3)}{(0.8\ \text{bar})} = \mathbf{218.75\ dm^3}$$

Q2. *A certain amount of a gas at 27°C and 1 bar pressure occupies a volume of 25 m^3. If the pressure is kept constant and the temperature is raised to 77°C, what will be the volume of the gas?*

Ans. From the available data: V_1 = 25 m^3, T_1 = 27 + 273 = 300 K

V_2 = ? T_2 = 77 + 273 = 350 K

Since the pressure of the gas is constant, Charles' law is applicable

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} \quad \text{or} \quad V_2 = \frac{V_1 \times T_2}{T_1}$$

$$V_2 = \frac{(25\ \text{m}^3) \times (350)\ \text{K}}{(300\ \text{K})} = \mathbf{29.17\ m^3.}$$

Q3. *A flask was heated from 27°C to 227°C at constant pressure. Calculate the volume of the flask if 0.1 dm^3 of air measured at 227°C was expelled from the flask.*

Ans. Let the volume of the flask = V dm^3 (after expelling the air)

$V_1 = V$ dm^3, $T_1 = 27 + 273 = 300$ K

$V_2 = (V + 0.1)$ dm^3, $T_2 = 227 + 273 = 500$ K

Since the pressure of the gas is constant, Charles' law is applicable.

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} \quad \text{or} \quad V_1 = \frac{T_1}{T_2} \times V_2$$

or $$V = \frac{(300\ \text{K}) \times (V + 0.1\ \text{dm}^3)}{(500\ \text{K})} \quad \text{or } V = \frac{3\,(V + 0.1)}{5}\ \text{dm}^3$$

or $$5V - 3V = 0.3\ \text{dm}^3 \quad \text{or} \quad 2\text{V} = 0.3\ \text{dm}^3$$

∴ $$\text{V} = \frac{0.3\ \text{dm}^3}{2} = \mathbf{0.15\ dm^3}$$

Q4. *A gas occupying a volume of 100 litres is at 20°C under a pressure of 2 bar. What temperature will it have when it is placed in an evacuated chamber of volume 175 litres? The pressure of the gas in the chamber is one-third of its initial pressure.*

Ans. From the available data:

$V_1 = 100$ L, $V_2 = 175$ L

$P_1 = 2$ bar, $P_2 = 2 \times 1/3 = 2/3$ bar

$T_1 = 20 + 273 = 293$ K, $T_2 = ?$

According to Gas equation, $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{or} \quad T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

By substituting the values, $$T_2 = \frac{(2/3\ \text{bar}) \times (175\ \text{L}) \times (293\ \text{K})}{(2\ \text{bar}) \times (100\ \text{L})}$$

$$= 170.9\ \text{K} = 170.9 - 273.0 = \mathbf{-102.1°C}$$

Q5. *At 25°C and 760 mm of Hg pressure a gas occupies 600 mL volume. What will be its pressure at a height where temperature is 10°C and volume of the gas is 640 mL?*

Ans.

$P_1 = 760$ mm Hg, $V_1 = 600$ mL

$T_1 = 25 + 273 = 298$ K

$V_2 = 640$ mL and $T_2 = 10 + 273 = 283$ K

According to combined gas law,

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

$$\Rightarrow \qquad p_2 = \frac{p_1V_1T_2}{T_1V_2}$$

$$\Rightarrow \qquad p_2 = \frac{(760 \text{ mm Hg}) \times (600 \text{ mL}) \times (283 \text{ K})}{(640 \text{ mL}) \times (298 \text{ K})}$$

$$= \mathbf{676.6 \text{ mm Hg}}$$

Q6. *A 34.0 dm*^{3} *cylinder contains 212 g of oxygen gas at 21°C. What mass of oxygen must be released to reduce the pressure in the cylinder to 1.24 bar?*

Ans. **Step I.** Calculation of no. of moles of O_2 left in cylinder.

P = 1.24 bar, $V = 34 \text{ dm}^3$

T = (21 + 273) = 294 K, $R = 0.083 \text{ dm}^3 \text{ bar K}^{-1} \text{ mol}^{-1}$

According to ideal gas equation,

$$PV = nRT$$

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

$$n = \frac{(1.24 \text{ bar}) \times (34 \text{ dm}^3)}{(0.083 \text{ dm}^3 \text{bar K}^{-1} \text{ mol}^{-1}) \times (294 \text{ K})}$$

$$= \mathbf{1.727 \text{ mol}}$$

Step II. Calculation of mass of oxygen released.

Mass of O_2 left in the cylinder = $n \times M$ = (1.727 mol) × (32g mol^{-1}) = 55.26 g

Mass of O_2 initially present = 212 g

Mass of O_2 released = (212 – 55.26) **= 156.74 g**

Q7. *The values of the van der Waal's constants for a gas are* $a = 4.10 \text{ dm}^6 \text{ bar mol}^{-2}$ *and* $b = 0.035 \text{ dm}^3 \text{ mol}^{-1}$. *Calculate the values of the critical temperature and critical pressure for the gas.*

Ans. (*i*) Calculation of critical temperature (T_c).

$a = 4.10 \text{ dm}^6 \text{ bar mol}^{-2}$

$b = 0.035 \text{ dm}^3 \text{ mol}^{-1}$

$R = 0.0821 \text{ dm}^3 \text{ bar mol}^{-1} K^{-1}$.

Now, critical temperature, $$T_c = \frac{8a}{27Rb}$$

Substituting the values, $$T_c = \frac{8 \times 4.10}{27 \times 0.0821 \times 0.035}$$

(*ii*) Calculation of critical pressure (P_c)

$$P_c = \frac{a}{27b^2} = \frac{4.10}{27 \times 0.035 \times 0.035} = \mathbf{123.96 \text{ bar}}$$

Q8. *The pressure of a mixture of* H_2 *and* N_2 *in a container is 1200 torr. The partial pressure of nitrogen in the mixture is 300 torr. What is the ratio of* H_2 *and* N_2 *molecules in the mixture?*

Ans. Total pressure of mixture = 1200 torr

Partial pressure of N_2 (P_{N_2}) = 300 torr

Partial pressure of H_2 (P_{H_2}) = 1200 – 300 = 900 torr

According to ideal gas equation,

$$PV = nRT$$

$$P_{H_2} = \frac{n_{H_2}RT}{V} = 900 \text{ torr} \qquad ...(i)$$

$$P_{N_2} = \frac{n_{N_2}RT}{V} = 300 \text{ torr} \qquad ...(ii)$$

Divide (*i*) by (*ii*), $$\frac{P_{H_2}}{P_{N_2}} = \frac{n_{H_2}}{n_{N_2}} = \frac{900}{300} = \frac{3}{1} \therefore n_{H_2} : n_{N_2} :: 3:1$$

Q9. *(a) What do you mean by 'Surface Tension' of a liquid?*

(b) Explain the factors which can affect the surface tension of a liquid.

Ans. (*a*) **Surface tension:** It is defined as the force acting per unit length perpendicular to the line drawn on the surface. It's unit is Nm^{-1}.

(*b*) Surface tension of a liquid depends upon following factors.

(*i*) **Temperature:** Surface tension decreases with rise in temperature. As the temperature of the liquid increases, the average kinetic energy of the molecules increases. Thus, there is a decrease in intermolecular force of attraction which decreases the surface tension.

(*ii*) **Nature of the liquid:** Greater the magnitude of intermolecular forces of attraction in the liquid, greater will be the value of surface tension.

Q10. *A neon-dioxygen mixture contains 70.6 g dioxygen and 167.5g neon. If pressure of the mixture of gases in the cylinder is 25 bar. What is the partial pressure of dioxygen and neon in the mixture?*

Ans. Number of moles of dioxygen $= \frac{70.6 \text{ g}}{32\text{g mol}^{-1}}$

$= 2.21$ mol

Number of moles of neon $= \frac{167.5 \text{ g}}{20 \text{ g mol}^{-1}}$

$= 8.375$ mol

Mole fraction of dioxygen $= \frac{2.21}{2.21+8.375} = \frac{2.21}{10.585}$

$= 0.21$

Mole fraction of neon $= \frac{8.375}{2.21+8.375}$

$= 0.79$

Alternatively,

mole fraction of neon = 1 – 0.21 = 0.79

Partial pressure of a gas = mole fraction × total pressure

⇒ Partial pressure of oxygen = 0.21 × (25 bar)

= **5.25 bar**

Partial Pressure of neon = 0.79 × (25 bar)

= **19.75 bar**

III. LONG ANSWER TYPE QUESTIONS

Q1. *State and explain Boyle's law. Represent the law graphically.*

Ans. It states that, the pressure of a fixed mass of a gas is inversely proportional to its volume if temperature is kept constant.

$$P \propto \frac{1}{V}$$

$$PV = \text{constant } (n \text{ and } T \text{ are constant})$$

$$P_1V_1 = P_2V_2.$$

Graphical representation:

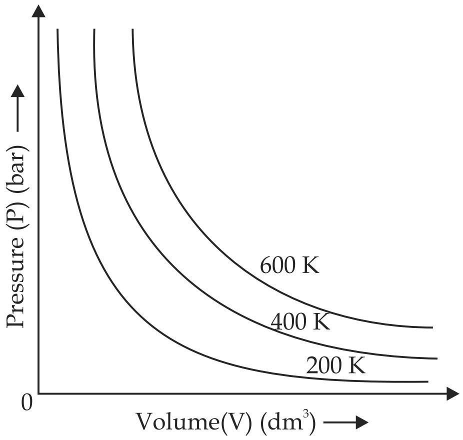


Fig. *Graph of pressure, P vs. Volume, V of a gas at different temperatures.*

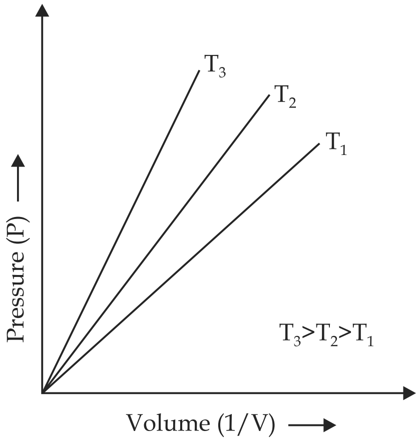


Fig. *Graph of pressure of a gas, P vs. 1/V*

Q2. *Give an expression for the van der Waals equation. Give the significance of the constants used in the equation. What are their units?*

Ans.

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

When n is the no. of moles present and 'a' and 'b' are known as van der Waals constants.

Significance of van der Waals constants

van der Waals constant '*a*' : 'a' is related to the magnitude of the attractive forces among the molecules of a particular gas. Greater the value of 'a', more will be the attractive forces.

$$\text{Unit of 'a'} = L^2\ \text{mol}^{-2}$$

van der Waals Constant '*b*': 'b' determines the volume occupied by the gas molecules which depends upon size of molecule.

$$\text{Unit of '}b\text{'} = L\ \text{mol}^{-1}.$$

Q3. *What are ideal and real gases? Out of CO_2 and NH_3 gases, which is expected to show more deviation from the ideal gas behaviour?*

Ans. **Ideal Gas:** A gas that follows Boyle's law, Charles' law and Avogadro law strictly, is called an ideal gas. It is assumed that intermolecular forces are not present between the molecules of an ideal gas.

Real Gases: Gases which deviate from ideal gas behaviour are known as real gases. NH_3 is expected to show more deviation. Since NH_3 is polar in nature and it can be liquified easily.

Q4. *State and explain Dalton's law of partial pressures. Can we apply Dalton's law of partial pressures to a mixture of carbon monoxide and oxygen?*

Ans. **Dalton's law of partial pressure:** When two or more non-reacting gases are enclosed in a vessel, the total pressure of the gaseous mixture is equal to the sum of the partial pressures that each gas will exert when enclosed separately in the same vessel at constant temperature.

$$P = P_1 + P_2 + P_3$$

Where, P is the total pressure of the three gases A, B, and C enclosed in a container. P_1, P_2 and P_3 are the partial pressures of the three gases when enclosed separately in the same vessel at a given temperature one by one.

No, the law cannot be applied. Carbon monoxide and oxygen readily combine to form carbon dioxide. The law can be applied only to the non-reacting gases.

IV. MULTIPLE CHOICE QUESTIONS

1. For one mole of a gas, the ideal gas equation is

(*a*) $PV = \frac{1}{2} RT$ (*b*) $PV = RT$ (*c*) $PV = \frac{3}{2} RT$ (*d*) $PV = \frac{5}{2} RT$

2. The average kinetic energy of the gas molecule is

(*a*) inversely proportional to its absolute temperature
(*b*) directly proportional to its absolute temperature
(*c*) equal to the square of its absolute temperature
(*d*) All of the above

3. Which of the following is the correct mathematical relation for Charles law at constant pressure?

(*a*) $V \propto T$ (*b*) $V \propto t$ (*c*) $V \propto \frac{1}{T}$ (*d*) all of above

4. At constant temperature, the pressure of the gas is reduced to one-third, the volume

(*a*) reduce to one-third (*b*) increases by three times
(*c*) remaining the same (*d*) cannot be predicted

5. With rise in temperature, the surface tension of a liquid

(*a*) decreases (*b*) increases
(*c*) remaining the same (*d*) none of the above

6. Viscosity of a liquid is a measure of

(*a*) repulsive forces between the liquid molecules
(*b*) frictional resistance
(*c*) intermolecular forces between the molecules
(*d*) none of the above

7. The cleansing action of soaps and detergents is due to

(*a*) internal friction (*b*) high hydrogen bonding
(*c*) viscosity (*d*) surface tensions

8. In van der Waals equation of state for a non-ideal gas the net force of attaction among the molecules is given by

(*a*) $\frac{an^2}{V^2}$ (*b*) $P + \frac{an^2}{V^2}$ (*c*) $P - \frac{an^2}{V^2}$ (*d*) $-\frac{an^2}{V^2}$

9. The compressibility factor, z for an ideal gas is

(*a*) zero (*b*) less than one (*c*) greater than one (*d*) equal to one

10. Which of the following gases will have the lowest rate of diffusion?

(*a*) H_2 (*b*) N_2 (*c*) F_2 (*d*) O_2

Ans. 1. (*b*) 2. (*b*) 3. (*a*) 4. (*b*) 5. (*a*)
6. (*b*) 7. (*d*) 8. (*a*) 9. (*d*) 10. (*c*)

V. HOTS QUESTIONS

Q1. *(a) Why aerated water bottles kept under water during summer?*
(b) Which property of liquid is responsible for spherical shape of drop?
(c) Why is moist air lighter than dry air?
(d) Define aqueous tension.
(e) What are units of a and b which are van der Waals constants?

Ans. (*a*) To reduce temperature, so as to reduce pressure, otherwise bottle may burst.
(*b*) Surface Tension.
(*c*) Moist air has water vapours which lowers vapour density, so it is lighter.
(*d*) It is pressure of water vapours at given temperature.
(*e*) Unit of a is $L^2\ mol^{-2}$, b is $L\ mol^{-1}$.

Q2. *Why does sharpened edge becomes smooth on heating up to melting point?*

Ans. On heating the glass, it melts and take up rounded shape at edges which has minimum surface area b/c of surface tension.

Q3. *Arrange the following in order of increasing density:*

$d = \frac{PM}{RT}$, *O_2 at 25°C, 2 atm. O_2 at 0°C, 2 atm. O_2 at 273°C, 1 atm.*

Ans. $d = \frac{PM}{RT}$, R and M are constant, so d depends upon $\frac{P}{T}$. So at 25°C, 1 atm, $\frac{P}{T} = \frac{1}{298}$. At 273° C, 1 atm, $\frac{P}{T} = \frac{1}{546}$. Hence, increasing order of density will be:

O_2 at 273° C, 1 atm < O_2 at 25° C, 2 atm.

Q4. *An O_2 cylinder has 10 L O_2 at 200 atm. If patient takes 0.50 ml of O_2 at 1 atm in one breath at 37°C, how many breaths are possible?*

Ans. $P_1 = 200$ atm, $V_1 = 10$L
$P_2 = 1$ atm, $V_2 = ?$
$P_1V_1 = P_2V_2 \Rightarrow 200 \times 10 = 1 \times V_2$ or $V_2 = 2000$ L.

$$\text{No. of breathes} = \frac{\text{Total Volume}}{\text{Volume for 1 breath}} = \frac{2000\text{L}}{0.5 \times 10^{-3}\text{L}} = 4 \times 10^{6}$$

❑❑❑