

LESSON AT A GLANCE

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

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

or $PV = \text{constant}$

or $P_1V_1 = P_2V_2$

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

$$V \propto T$$

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

- **Avogadro's 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 $R = \text{universal constant.}$

- **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.

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

- **Van der Waal's equation**

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

• **Critical Temperature:**

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

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

(Critical volume) $V_c = 3b$.

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?

Ans. Given: $p_1 = 1 \text{ bar}$ $p_2 = ?$
 $V_1 = 500 \text{ dm}^3$ $V_2 = 200 \text{ dm}^3$

At constant temperature ($t = 30 \text{ }^\circ\text{C}$)

$$p_2 V_2 = p_1 V_1$$

$$p_2 = \frac{p_1 V_1}{V_2} = \frac{(1 \text{ bar}) \times (500 \text{ dm}^3)}{(200 \text{ dm}^3)} = 2.5 \text{ bar.}$$

∴ The minimum pressure required for compression = 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?

Ans. Given: Temperature = Constant = 35 °C

$$p_1 = 1.2 \text{ bar} \quad p_2 = ?$$

$$V_1 = 120 \text{ mL} \quad V_2 = 180 \text{ mL}$$

$$p_2 V_2 = p_1 V_1$$

$$p_2 = \frac{p_1 V_1}{V_2} = \frac{(1.2 \text{ bar}) \times (120 \text{ mL})}{(180 \text{ mL})} = 0.8 \text{ bar.}$$

Q3. Using the equation of state $pV = nRT$, show that at a given temperature density of a gas is proportional to gas pressure p .

Ans. The number of moles, $n = \frac{\text{Mass}}{\text{Molar mass}} = \frac{w}{M}$

Substituting the value of n in the ideal gas equation, $pV = nRT$ gives,

$$pV = \frac{w}{M} RT$$

or
$$p = \frac{w}{V} \frac{RT}{M}$$

$$p = d \frac{RT}{M} \quad \left(\because \frac{w}{V} = d, \text{ density} \right)$$

At a given temperature (T), R and M (molar mass of the gas) are constant.

Therefore, $\frac{RT}{M}$ is constant.

So, it can be written as

$$p = \text{constant} \times d \quad \text{or} \quad p \propto d$$

i.e., at a given temperature, the pressure of a gas is directly proportional to its density.

- Q4.** 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?

Ans. Density, $d = \frac{pM}{RT}$

At constant T and for same density of two gases,

$p_1 M_1$	=	$p_2 M_2$
gaseous oxide		dinitrogen
Gaseous Oxide		Dinitrogen (N_2)
$p_1 = 2 \text{ bar}$		$p_2 = 5 \text{ bar}$
$M_1 = ?$		$M_2 = 2 \times 14 = 28 \text{ u}$
$M_1 = \frac{p_2 M_2}{p_1} = \frac{(5 \text{ bar}) \times (28 \text{ u})}{(2 \text{ bar})} = 70 \text{ u}$		

\therefore Molecular mass of gaseous oxide = 70 u.

- 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 a relationship between their molecular masses.

Ans. Given: $p_A = 2 \text{ bar}$

$p_A + p_B = 3 \text{ bar}$ (Dalton's law of partial pressures)

$$\therefore p_B = 3 - 2 = 1 \text{ bar}$$

Gas A

$$m_A = 1 \text{ g}$$

$$p_A = 2 \text{ bar}$$

Gas B

$$m_B = 2 \text{ g}$$

$$p_B = 1 \text{ bar}$$

$$V = V$$

$$V = V$$

$$T = 27 + 273 = 300 \text{ K} \quad T = 300 \text{ K}$$

From ideal gas equation,

$$pV = nRT$$

$$= \frac{m}{M} RT \quad (m = \text{mass and } M = \text{molar mass})$$

$$\text{or} \quad p = \frac{mRT}{MV}$$

$$\text{Gas A: } p_A = \frac{m_A RT}{M_A V}$$

$$\text{Gas B: } p_B = \frac{m_B RT}{M_B V}$$

Since $\frac{RT}{V} = \text{constant}$ for both the gases

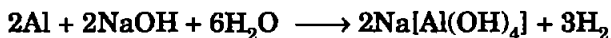
$$\frac{p_A}{p_B} = \frac{m_A}{M_A} \times \frac{M_B}{m_B} = \frac{m_A}{m_B} \cdot \frac{M_B}{M_A}$$

$$\frac{2 \text{ bar}}{1 \text{ bar}} = \frac{1 \text{ g}}{2 \text{ g}} \cdot \frac{M_B}{M_A}$$

$$\therefore \frac{M_B}{M_A} = 2 \times 2 = 4 \quad \text{or} \quad M_B = 4 M_A$$

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 0.15 g of aluminium reacts?

Ans. The reaction is



2 mol

3 mol

$$2 \times 27 = 54 \text{ g}$$

Thus, 54 g Al produces 3 mol H₂.

$$0.15 \text{ g Al would produce } \frac{3 \times 0.15}{54} = 0.00833 \text{ mol H}_2.$$

From ideal gas equation,

$$pV = nRT$$

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

Given: $T = 20 + 273 = 293 \text{ K}$

$$p = 1 \text{ bar}$$

$$R = 8.314 \times 10^{-2} \text{ L bar K}^{-1} \text{ mol}^{-1}$$

$$n = 0.00833 \text{ mol}$$

$$\begin{aligned} \therefore V &= \frac{(0.00833 \text{ mol}) \times (8.314 \times 10^{-2} \text{ L bar K}^{-1} \text{ mol}^{-1}) \times (293 \text{ K})}{(1 \text{ bar})} \\ &= 0.2029 \text{ L} = \mathbf{202.9 \text{ mL}} \end{aligned}$$

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

Ans. Given: $V = 9 \text{ dm}^3$, $T = 27 + 273 = 300 \text{ K}$

$$R = 8.314 \times 10^{-2} \text{ dm}^3 \text{ bar K}^{-1} \text{ mol}^{-1}$$

(i) **Partial pressure exerted by 3.2 g of methane**

Molar mass of methane,

$$\text{CH}_4 = 12 + 4 = 16 \text{ g mol}^{-1}$$

Number of moles of methane

$$= \frac{3.2 \text{ g}}{16 \text{ g mol}^{-1}} = 0.2 \text{ mol}$$

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

$$= \frac{(0.2 \text{ mol}) (8.314 \times 10^{-2} \text{ dm}^3 \text{ bar K}^{-1} \text{ mol}^{-1}) (300 \text{ K})}{(9 \text{ dm}^3)}$$

$$= \mathbf{0.5543 \text{ bar.}}$$

(ii) **Partial pressure exerted by 4.4 g of carbon dioxide**

Molar mass of carbon dioxide,

$$\text{CO}_2 = (12 + 2 \times 16) = 44 \text{ g mol}^{-1}$$

Number of moles of carbon dioxide

$$= \frac{4.4 \text{ g}}{44 \text{ g mol}^{-1}} = 0.1 \text{ mol}$$

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

$$= \frac{(0.1 \text{ mol}) (8.314 \times 10^{-2} \text{ dm}^3 \text{ bar K}^{-1} \text{ mol}^{-1}) (300 \text{ K})}{(9 \text{ dm}^3)}$$

$$= \mathbf{0.2771 \text{ bar.}}$$

(iii) Total pressure exerted by the mixture

$$\begin{aligned}
 p &= p_{\text{CH}_4} + p_{\text{CO}_2} \\
 &= (0.5543 + 0.2771) \text{ bar} = 0.8314 \text{ bar} \\
 &= 0.8314 \times 10^5 \text{ Pa} = \mathbf{8.314 \times 10^4 \text{ Pa}}.
 \end{aligned}$$

Q8. What will be the pressure of the gaseous 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?

Ans. (i) **Partial pressure of H_2**

$$\begin{array}{ll}
 p_1 = 0.8 \text{ bar} & p_2 = ? \\
 V_1 = 0.5 \text{ L} & V_2 = 1 \text{ L}
 \end{array}$$

$$p_2 = \frac{p_1 V_1}{V_2} = \frac{(0.8 \text{ bar})(0.5 \text{ L})}{(1 \text{ L})} = 0.4 \text{ bar} = p_{\text{H}_2}$$

(ii) **Partial pressure of O_2**

$$\begin{array}{ll}
 p_1 = 0.7 \text{ bar} & p_2 = ? \\
 V_1 = 2.0 \text{ L} & V_2 = 1 \text{ L}
 \end{array}$$

$$p_2 = \frac{p_1 V_1}{V_2} = \frac{(0.7 \text{ bar})(2.0 \text{ L})}{(1 \text{ L})} = 1.4 \text{ bar} = p_{\text{O}_2}$$

$$\begin{aligned}
 \text{Total pressure, } p &= p_{\text{H}_2} + p_{\text{O}_2} \\
 &= (0.4 + 1.4) \text{ bar} = \mathbf{1.8 \text{ bar}}.
 \end{aligned}$$

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

Ans. Given: $d_1 = 5.46 \text{ g dm}^{-3}$ $d_2 = ?$

$$\begin{aligned}
 T_1 &= 27 \text{ °C} = 27 + 273 \\
 &= 300 \text{ K}
 \end{aligned}$$

$$T_2 = 273 \text{ K}$$

$$p_1 = 2 \text{ bar}$$

$$p_2 = 1 \text{ bar}$$

$$d = \frac{pM}{RT}, d_1 = \frac{p_1 M}{RT_1}, d_2 = \frac{p_2 M}{RT_2}$$

For same gas, M and R are constant

$$\therefore \frac{d_2}{d_1} = \frac{p_2 T_1}{p_1 T_2}$$

$$d_2 = \frac{d_1 p_2 T_1}{p_1 T_2}$$

$$= \frac{(5.46 \text{ g dm}^{-3})(1 \text{ bar})(300 \text{ K})}{(2 \text{ bar})(273 \text{ K})} = \mathbf{3 \text{ g dm}^{-3}}.$$

Q10. 34.05 mL of phosphorus vapour weighs 0.0625 g at 546 °C and 0.1 bar pressure. What is the molar mass of phosphorus?

Ans. Given: $p = 0.1$ bar

$$V = 34.05 \text{ mL} = 34.05 \times 10^{-3} \text{ L}$$

$$R = 8.314 \times 10^{-2} \text{ bar L K}^{-1} \text{ mol}^{-1}$$

$$T = 546 \text{ }^\circ\text{C} = 546 + 273 = 819 \text{ K}$$

$$m = 0.0625 \text{ g}$$

Let the molar mass be M .

\therefore Number of moles of phosphorus vapours

$$n = \frac{m}{M} = \frac{0.0625}{M} \text{ mol}$$

From ideal gas equation

$$pV = nRT = \frac{mRT}{M}$$

or
$$M = \frac{mRT}{pV}$$

$$= \frac{(0.0625 \text{ g})(8.314 \times 10^{-2} \text{ bar L K}^{-1} \text{ mol}^{-1})(819 \text{ K})}{(0.1 \text{ bar})(34.05 \times 10^{-3} \text{ L})}$$

$$= 1249.8 \text{ g mol}^{-1}.$$

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?

Ans. Given: $V_1 =$ volume of air in flask = V

$$V_2 = ?$$

$$T_1 = 27 + 273 = 300 \text{ K}$$

$$T_2 = 477 + 273 = 750 \text{ K}$$

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

$$\therefore V_2 = \frac{V_1 T_2}{T_1} = \frac{V \times 750 \text{ K}}{300 \text{ K}} = 2.5 V$$

Volume left in the flask = V

Volume of air expelled = $(2.5 - 1) V = 1.5 V$

$$\text{Fraction of air expelled} = \frac{1.5V}{2.5V} = \frac{3}{5}$$

Q12. Calculate the temperature of 4.0 mol of a gas occupying 5 dm³ volume at 3.32 bar. ($R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$)

Ans. Given: $p = 3.32 \text{ bar}$ $V = 5 \text{ dm}^3$
 $T = ?$ $n = 4.0 \text{ mol}$
 $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$
 $pV = nRT$

$$T = \frac{pV}{nR}$$

$$= \frac{(3.32 \text{ bar}) \times (5 \text{ dm}^3)}{(4.0 \text{ mol}) \times (0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1})}$$

$$= 50 \text{ K.}$$

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

Ans. Atomic number of N = 7

∴ There are 7 electrons in one N atom and 14 in one N₂ molecule.

$$\text{Molar mass of N}_2 = 28 \text{ g mol}^{-1}$$

$$\text{Mass of N}_2 = 1.4 \text{ g}$$

$$\therefore \text{No. of moles of N}_2 = \frac{\text{Mass}}{\text{Molar mass}}$$

$$= \frac{1.4 \text{ g}}{28 \text{ g mol}^{-1}} = 0.05 \text{ mol}$$

Number of molecules of N₂

$$= \text{No. of moles} \times \text{Avogadro's constant}$$

$$= 0.05 \text{ mol} \times 6.022 \times 10^{23} \text{ mol}^{-1}$$

$$\text{Number of electrons} = 14 \times 0.05 \times 6.022 \times 10^{23}$$

$$= 4.2154 \times 10^{23}.$$

Q14. How much time would it take to distribute one Avogadro number of wheat grains, if 10¹⁰ grains are distributed each second?

Ans. Number of wheat grains = 6.022 × 10²³

$$\text{Time taken for distribution of } 10^{10} \text{ grains} = 1 \text{ s}$$

∴ Time taken for distribution of

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

$$= 6.022 \times 10^{13} \text{ s}$$

$$1 \text{ year} = \left(\frac{365 \text{ day}}{\text{yr}}\right) \times \left(\frac{24 \text{ hr}}{\text{day}}\right) \times \left(\frac{60 \text{ min}}{\text{hr}}\right) \times \left(\frac{60 \text{ s}}{\text{min}}\right)$$

$$= 3.1536 \times 10^7 \text{ s yr}^{-1}$$

$$\therefore \text{Time taken} = \frac{6.022 \times 10^{13} \text{ s}}{3.1536 \times 10^7 \text{ s yr}^{-1}} = 1.90956 \times 10^6 \text{ yr.}$$

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

Ans. Given: $V = 1 \text{ dm}^3$

$$T = 27 + 273 = 300 \text{ K}$$

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

(i) **Partial pressure of O_2**

$$\text{Mass of } \text{O}_2 = 8 \text{ g}$$

$$\text{Molar mass of } \text{O}_2 = 16 \times 2 = 32 \text{ g mol}^{-1}$$

$$\text{No. of moles of } \text{O}_2 = \frac{8 \text{ g}}{32 \text{ g mol}^{-1}} = 0.25 \text{ mol}$$

$$\text{Partial pressure of } \text{O}_2 = p_{\text{O}_2} = \frac{nRT}{V}$$

$$= \frac{(0.25 \text{ mol})(0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1})(300 \text{ K})}{(1 \text{ dm}^3)}$$

$$p_{\text{O}_2} = 6.225 \text{ bar.}$$

(ii) **Partial pressure of H_2**

$$\text{Mass of } \text{H}_2 = 4 \text{ g}$$

$$\text{Molar mass of } \text{H}_2 = 2 \times 1 = 2 \text{ g mol}^{-1}$$

$$\text{Number of moles of } \text{H}_2 = \frac{4 \text{ g}}{2 \text{ g mol}^{-1}} = 2 \text{ mol}$$

$$\therefore \text{Partial pressure of } \text{H}_2 = p_{\text{H}_2} = \frac{nRT}{V}$$

$$= \frac{(2 \text{ mol})(0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1})(300 \text{ K})}{(1 \text{ dm}^3)}$$

$$p_{\text{H}_2} = 49.800 \text{ bar.}$$

$$\text{Total pressure} = 6.225 + 49.800 = 56.025 \text{ bar.}$$

Alternately,

(i) Moles of $\text{O}_2 = 0.25 \text{ mol}$ (as calculated earlier)

(ii) Moles of $H_2 = 2$ mol (as calculated earlier)

(iii) Total moles = $0.25 + 2 = 2.25$ mol

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

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

$$= 56.025 \text{ bar.}$$

Q16. Pay load is defined as the difference between the mass of 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 $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$)

Ans. Radius of balloon = 10 m

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

$$= 4190.5 \text{ m}^3 = 4190.5 \times 10^3 \text{ dm}^3$$

Mass of displaced air,

$$\text{Density of air} = 1.2 \text{ kg m}^{-3}$$

$$\text{Mass} = \text{Volume} \times \text{Density}$$

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

Total mass of balloon,

$$\text{Mass of balloon} = 100 \text{ kg}$$

Mass of He in the balloon: From ideal gas equation

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

$$m = \frac{pVM}{RT}$$

Given:

$$p = 1.66 \text{ bar}$$

$$V = 4190.5 \times 10^3 \text{ dm}^3$$

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

$$T = 27 + 273 = 300 \text{ K}$$

$$M \text{ for He} = 4 \times 10^{-3} \text{ kg mol}^{-1}$$

$$\therefore m = \frac{(1.66 \text{ bar})(4190.5 \times 10^3 \text{ dm}^3)(4.0 \times 10^{-3} \text{ kg mol}^{-1})}{(0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1})(300 \text{ K})}$$

$$= 1117.5 \text{ kg}$$

$$\begin{aligned}\text{Total mass of balloon} &= \text{Mass of balloon} + \text{Mass of He filled} \\ &= 100 + 1117.5 = 1217.5 \text{ kg}\end{aligned}$$

$$\begin{aligned}\text{Payload} &= \text{Mass of displaced air} - \text{Total mass of balloon} \\ &= 5028.6 - 1217.5 = 3811.1 \text{ kg}.\end{aligned}$$

Q17. Calculate the volume occupied by 8.8 g of CO_2 at 31.1°C and 1 bar pressure. $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$.

Ans. Given: $p = 1 \text{ bar}$ $V = ?$

$$T = 31.1 + 273 = 304.1 \text{ K}$$

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

$$\text{Mass of } \text{CO}_2 = 8.8 \text{ g}$$

$$\text{Molar mass of } \text{CO}_2 = 12 + 2 \times 16 = 44 \text{ g mol}^{-1}$$

$$\begin{aligned}\text{No. of moles of } \text{CO}_2 &= \frac{\text{Mass}}{\text{Molar mass}} \\ &= \frac{8.8 \text{ g}}{44 \text{ g mol}^{-1}} = 0.2 \text{ mol}\end{aligned}$$

From ideal gas equation

$$pV = nRT$$

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

$$= \frac{(0.2 \text{ mol})(0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1})(304.1 \text{ K})}{(1 \text{ bar})}$$

$$= 5.05 \text{ dm}^3.$$

Q18. 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?

Ans. Given,

For H_2 gas,

$$p_1 = p$$

$$V_1 = V$$

$$T_1 = 17 + 273$$

$$= 290 \text{ K}$$

$$m_1 = 0.184 \text{ g}$$

$$M_1 = 2 \text{ g mol}^{-1}$$

For unknown gas,

$$p_2 = p$$

$$V_2 = V$$

$$T_2 = 95 + 273$$

$$= 368 \text{ K}$$

$$m_2 = 2.9 \text{ g}$$

$$M_2 = ?$$

From ideal gas equation,

$$pV = \frac{mRT}{M}$$

For H_2 gas,

$$pV = \frac{0.184 \times R \times 290}{2}$$

For unknown gas,

$$pV = \frac{2.9 \times R \times 368}{M_2}$$

From the above two relations (equating their right hand sides, since their left hand sides are equal),

$$\frac{0.184 \times R \times 290}{2} = \frac{2.9 \times R \times 368}{M_2}$$

$$M_2 = \frac{2.9 \times 368 \times 2}{0.184 \times 290} = 40 \text{ g mol}^{-1}.$$

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

Ans. Partial pressure of a gas = (Mole fraction of the gas) \times (Total pressure of gaseous mixture)

Mole fraction of dihydrogen, H_2

$$= \frac{\text{Number of moles of } \text{H}_2}{\text{Total no. of moles of } (\text{H}_2 + \text{O}_2)} = \frac{n_{\text{H}_2}}{n_{\text{H}_2} + n_{\text{O}_2}}$$

Let the total mass of the gaseous mixture be 100 g

Then, mass of $\text{H}_2 = \frac{20 \times 100}{100} = 20 \text{ g}$

Molar mass of $\text{H}_2 = 2 \text{ g mol}^{-1}$

Number of moles of H_2 ,

$$n_{\text{H}_2} = \frac{20 \text{ g}}{2 \text{ g mol}^{-1}} = 10 \text{ mol}$$

$$\text{Mass of } \text{O}_2 = \frac{80 \times 100}{100} = 80 \text{ g}$$

Molar mass of $\text{O}_2 = 2 \times 16 = 32 \text{ g mol}^{-1}$

Number of moles of O_2 ,

$$n_{\text{O}_2} = \frac{80 \text{ g}}{32 \text{ g mol}^{-1}} = 2.5 \text{ mol}$$

$$\text{Mole fraction of } \text{H}_2 = X_{\text{H}_2} = \frac{n_{\text{H}_2}}{n_{\text{H}_2} + n_{\text{O}_2}} = \frac{10}{10 + 2.5} = \frac{10}{12.5}$$

$$\begin{aligned}\text{Partial pressure of H}_2 &= p_{\text{H}_2} = X_{\text{H}_2} \times P_{\text{total}} \\ &= \left(\frac{10}{12.5} \right) (1 \text{ bar}) = 0.8 \text{ bar}.\end{aligned}$$

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

$$\begin{aligned}\text{Ans. SI unit of } \frac{pV^2T^2}{n} &= \frac{(\text{Nm}^{-2})(\text{m}^3)^2(\text{K})^2}{(\text{mol})} \\ &= \frac{\text{Nm}^{-2} \text{m}^6 \text{K}^2}{\text{mol}} = \text{Nm}^4 \text{K}^2 \text{mol}^{-1}.\end{aligned}$$

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

Ans. According to Charles' law,

$$V_t = V_0 \left(1 + \frac{t}{273} \right)$$

where, V_t = volume of gas at temperature $t^\circ\text{C}$

V_0 = volume of gas at 0°C

When $t = -273^\circ\text{C}$

$$V_{-273^\circ\text{C}} = V_0 \left[1 + \frac{(-273)}{273} \right] = V_0 (1 - 1) = 0.$$

Thus, according to Charles' law, the volume of any gas would become zero at -273°C . Any temperature below -273°C would result in negative value which is physically impossible. Therefore, -273°C is the lowest possible temperature.

Q22. Critical temperature for carbon dioxide and methane are 31.1°C and -81.9°C , respectively. Which of these has stronger intermolecular forces and why?

Ans. Carbon dioxide has higher critical temperature (31.1°C) as compared to that of methane (-81.9°C). It means that CO_2 can be liquefied more easily at higher temperature and it has stronger intermolecular forces than methane.

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

Ans. ' 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.

TEXTUAL PROBLEMS SOLVED

Problem 1. Calculate (i) root mean square speed (ii) average speed and (iii) most probable speed of CO_2 molecules at 700 K.

Solution. Given: $T = 700 \text{ K}$.

$$M \text{ for } \text{CO}_2 = 12 + 2 \times 16 = 44 \text{ g mol}^{-1} \\ = 0.044 \text{ kg mol}^{-1}$$

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

(i) **Root mean square speed,**

$$c = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (700 \text{ K})}{(0.044 \text{ kg mol}^{-1})}} \\ = 629.92 \text{ m s}^{-1}$$

(ii) **Average speed,**

$$\bar{c} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (700 \text{ K})}{3.14 \times (0.044 \text{ kg mol}^{-1})}} \\ = 580.50 \text{ m s}^{-1}$$

(iii) **Most probable speed,**

$$c^* = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (700 \text{ K})}{(0.044 \text{ kg mol}^{-1})}} \\ = 514.33 \text{ m s}^{-1}$$

Problem 2. A sample of a gas contains 15 molecules with a speed of 3 m s^{-1} , 25 molecules with a speed of 5 m s^{-1} and 30 molecules with a speed of 8 m s^{-1} . Calculate root mean square speed of these molecules.

Solution. **Root mean square speed**

$$c = \sqrt{\frac{N_1c_1^2 + N_2c_2^2 + N_3c_3^2}{N_1 + N_2 + N_3}} \\ c = \sqrt{\frac{(15 \times 3^2) + (25 \times 5^2) + (30 \times 8^2)}{15 + 25 + 30}} \\ c = \sqrt{\frac{(15 \times 9) + (25 \times 25) + (30 \times 64)}{15 + 25 + 30}} = 6.187 \text{ m s}^{-1}$$

Problem 3. Calculate the temperature at which the average speed of oxygen equals that of hydrogen at 20 K.

Solution. Average Speed $\bar{c} = \sqrt{\frac{8RT}{\pi M}}$

For hydrogen $T_1 = 20 \text{ K}$; $M_1 = 0.002 \text{ kg mol}^{-1}$

$$\bar{c}_{H_2} = \sqrt{\frac{8RT_1}{\pi M_1}}$$

For oxygen $T_2 = ?$, $M_2 = 0.032 \text{ kg mol}^{-1}$

$$\bar{c}_{O_2} = \sqrt{\frac{8RT_2}{M_2}}$$

Given $\bar{c}_{H_2} = \bar{c}_{O_2}$

$$\therefore \sqrt{\frac{8RT_1}{\pi M_1}} = \sqrt{\frac{8RT_2}{M_2}}$$

taking square of both the sides

$$\frac{8RT_1}{\pi M_1} = \frac{8RT_2}{M_2}$$

Or
$$\frac{T_1}{M_1} = \frac{T_2}{M_2}$$

$$T_2 = \frac{T_1 M_2}{M_1}$$

Putting for values of T_1 , M_1 and M_2 ,

$$T_2 = \frac{(20 \text{ K}) \times (0.032 \text{ kg mol}^{-1})}{(0.002 \text{ kg mol}^{-1})} = 320 \text{ K}$$

Problem 4. Calculate the temperatures at which the root mean square speed, average speed and the most probable speed of oxygen gas are all equal to 1500 ms^{-1} .

Solution. Given for oxygen $c^* = \bar{c} = c = 1500 \text{ m s}^{-1}$

$$M = 0.032 \text{ kg mol}^{-1}$$

(i) For root mean square speed,

$$c = 1500 \text{ m s}^{-1} = \sqrt{\frac{3RT}{M}}$$

Taking square of both sides

$$(1500 \text{ m s}^{-1})^2 = \frac{3RT}{M} = \frac{3 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (T)}{(0.032 \text{ kg mol}^{-1})}$$

$$T = \frac{(1500 \text{ m s}^{-1}) \times (1500 \text{ m s}^{-1}) \times (0.032 \text{ kg mol}^{-1})}{3 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1})}$$

$$= 2886.7 \text{ K}$$

(ii) For average speed,

$$\bar{c} = 1500 \text{ m s}^{-1} = \sqrt{\frac{8RT}{\pi M}}$$

Taking square of both side

$$c^2 = (1500 \text{ m s}^{-1})^2 = \frac{8RT}{\pi M} = \frac{8 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (T)}{(3.14) \times (0.032 \text{ kg mol}^{-1})}$$

$$T = \frac{(1500 \text{ m s}^{-1}) \times (1500 \text{ m s}^{-1}) \times (0.032 \text{ kg mol}^{-1})}{8 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1})}$$

$$= 3399.1 \text{ K}$$

(iii) For most probable speed,

$$\bar{c} = 1500 \text{ m s}^{-1} = \sqrt{\frac{2RT}{M}}$$

Taking square of both sides

$$c^2 = (1500 \text{ m s}^{-1})^2 = \frac{2RT}{M} = \frac{2 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (T)}{(0.032 \text{ kg mol}^{-1})}$$

$$T = \frac{(1500 \text{ m s}^{-1}) \times (1500 \text{ m s}^{-1}) \times (0.032 \text{ kg mol}^{-1})}{2 \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1})}$$

$$= 4330.0 \text{ K.}$$

□□□

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