



### TEXTBOOK QUESTIONS SOLVED

Question 1. What will be the minimum pressure required to compress 500 dm<sup>3</sup> of air at 1 bar to 200 dm<sup>3</sup> at 30°C?

Answer:

$$P_1 = 1 \text{ bar}, P_2 = ?$$

$$V_1 = 500 \text{ dm}^3, V_2 = 200 \text{ dm}^3$$

As temperature remains constant at 30°C,

$$P_1 V_1 = P_2 V_2$$

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

Question 2. 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?

Answer:

$$V_1 = 120 \text{ mL}, P_1 = 1.2 \text{ bar},$$

$$V_2 = 180 \text{ mL}, P_2 = ?$$

As temperature remains constant,  $P_1 V_1 = P_2 V_2$

$$(1.2 \text{ bar}) (120 \text{ mL}) = P_2 (180 \text{ mL})$$

Question 3. Using the equation of state  $PV = nRT$ , show that at a given temperature, density of a gas is proportional to the gas pressure  $P$ .

Answer:

According to ideal gas equation

$$PV = nRT \text{ or } PV = nRT/V$$

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

$$P = \frac{mRT}{MV}$$

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

$$P \propto \rho \text{ at constant temperature}$$

$$\left[ \because \rho (\text{density}) = \frac{m}{V} \right]$$

Question 4. 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?

Answer:

Using the expression,  $d = MP/RT$ , at the same temperature and for same density,

$$M_1 P_1 = M_2 P_2 \text{ (as } R \text{ is constant)}$$

$$(\text{Gaseous oxide}) (N_2)$$

or

$$M_1 \times 2 = 28 \times 5 \text{ (Molecular mass of } N_2 = 28 \text{ u)}$$

$$\text{or } M_1 = 70 \text{ u}$$

Question 5. 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.

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

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

$$P_A = 2 \text{ bar}, \quad P_A + P_B = 3 \text{ bar}, \quad \text{i.e., } P_B = 1 \text{ bar}$$

Applying the relation  $PV = nRT$

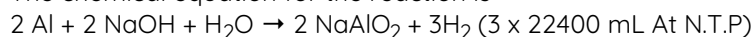
$$P_A V = n_A RT, \quad P_B V = n_B RT \quad \therefore \frac{P_A}{P_B} = \frac{n_A}{n_B} = \frac{1/M_A}{2/M_B} = \frac{M_B}{2M_A}$$

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

Question 6. 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?

Answer:

The chemical equation for the reaction is



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

54 g of Al at N.T.P release

$\text{H}_2$  gas =  $3 \times 22400$  0.15 g of Al at N.T.P release

$$\text{H}_2 \text{ gas} = \frac{3 \times 22400 \times 0.15}{54}$$

$$= 186.7 \text{ mL}$$

N.T.P condition.	$V_1 = 186.7 \text{ mL}$	$V_2 = ?$
	$P_1 = 1.013 \text{ bar}$	$P_2 = 1 \text{ bar}$
	$T_1 = 0 + 273 = 273 \text{ K}$	$T_2 = 20 + 273$
		$= 293 \text{ K}$

According to Gas equation

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \text{or} \quad V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

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

$$= 203 \text{ mL}$$

Question 7. What will be the pressure exerted by a mixture of 3.2g of methane and 4.4g of carbon dioxide contained in a 9 dm<sup>3</sup> flask at 27 °C?

Answer:

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

$$p_{\text{CH}_4} = \left( \frac{3.2}{16} \text{ mol} \right) \frac{0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{9 \text{ dm}^3} = 0.55 \text{ atm}$$

$$p_{\text{CO}_2} = \left( \frac{4.4}{44} \text{ mol} \right) \frac{0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{9 \text{ dm}^3} = 0.27 \text{ atm}$$

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

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

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

Question 8. What will be the pressure of the gas mixture when 0.5 L of  $\text{H}_2$  at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in all vessel at 27 °C?

Answer:

Calculation of partial pressure of  $\text{H}_2$  in 1L vessel  $P_1 = 0.8 \text{ bar}$ ,

$$P_2 = ?, \quad V_1 = 0.5 \text{ L}, \quad V_2 = 1.0 \text{ L}$$

As temperature remains constant,  $P_1 V_1 = P_2 V_2$

$$(0.8 \text{ bar})(0.5 \text{ L}) = P_2 (1.0 \text{ L}) \quad \text{or} \quad P_2 = 0.40 \text{ bar, i.e., } P_{\text{H}_2} = 0.40 \text{ bar}$$

Calculation of partial pressure of  $\text{O}_2$  in 1 L vessel

$$P_1' V_1 = P_2' V_2'$$

(0.7 bar) (2.0 L) =  $P_2$  (1L) or  $P_2' = 1.4$  bar, i.e.,  $P_{O_2} = 1.4$  bar  
 Total pressure =  $P_{H_2} + P_{O_2} = 0.4 \text{ bar} + 1.4 \text{ bar} = 1.8 \text{ bar}$

Question 9. Density of a gas is found to be  $5.46 \text{ g/dm}^3$  at  $27^\circ\text{C}$  and at 2 bar pressure. What will be its density at STP?

Answer:

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

Here,  $d_1 = 5.46 \text{ g dm}^{-3}$ ,  $T_1 = 27^\circ\text{C} = 300 \text{ K}$ ,  $P_1 = 2 \text{ bar}$ .  
 At STP,  $d_2 = ?$ ,  $T_2 = 0^\circ\text{C} = 273 \text{ K}$ ,  $P_2 = 1 \text{ bar}$

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

Question 10. 34.05 mL of phosphorus vapour weighs 0.0625 g at  $546^\circ\text{C}$  and 1.0 bar pressure. What is the molar mass of phosphorus?

Answer:

Step I. Calculation of volume at  $0^\circ\text{C}$  and 1 bar pressure

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \text{ i.e., } \frac{1 \times 34.05}{546 + 273} = \frac{1 \times V_2}{273} \text{ or } V_2 = 11.35 \text{ mL}$$

11.35 mL of vapour at  $0^\circ\text{C}$  and 1 bar pressure weigh = 0.0625 g

$\therefore$  22700 mL of vapour at  $0^\circ\text{C}$  and 1 bar pressure will weigh

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

$\therefore$  Molar mass =  **$125 \text{ g mol}^{-1}$**

Alternatively, using

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

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

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

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

$\therefore$  Molar mass =  **$125 \text{ g mol}^{-1}$**

\*\*\*\*\* END \*\*\*\*\*