NCERT TEXTUAL EXERCISE (SOLVED)

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

Ans. $P_1 = 1$ bar, $V_1 = 500\ dm^3$

$P_2 = ?$, $V_2 = 200\ dm^3$

As temperature remains constant at 30°C,

$$P_1V_1 = P_2V_2$$

1 bar × 500 dm³ = P_2 × 200 dm³ or $P_2 = \frac{500}{200}$ bar = 2.5 bar

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?

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

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

As température remains constant, $P_1V_1 = P_2V_2$

(1.2 bar) (120 mL) = P_2 (180 mL) or $P_2 = 0.8$ bar

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.

Ans. Using $PV = nRT$

$$PV = \frac{wRT}{M}, \text{ where } n = \text{number of moles} = \frac{\text{weight } (w)}{\text{Mol. Mass } (M)}$$

$$\therefore P = \frac{wRT}{VM}$$

$$P = d\frac{RT}{M}, \text{ where } d = \text{density} = \text{weight/volume} \quad or\ d = \frac{MP}{RT}$$

Thus, for a given temperature (M & R are constant), the density of gas is proportional to gas pressure i.e.

$$d \propto P.$$

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

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

M_1P_1 (Gaseous Oxide) = M_2P_2 (N_2) (as R is constant)

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

or $M_1 = 70$ μm.

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.

Ans. 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 $PV = nRT$

$P_AV = n_ART$, $P_BV = n_BRT$

$$\therefore \frac{P_A}{P_B} = \frac{n_A}{n_B} = \frac{1/M_A}{2/M_B} = \frac{M_B}{2M_A} \text{ or } \frac{M_B}{M_A} = 2\times\frac{P_A}{P_B} = 2\times\frac{2}{1} = 4 \text{ or } M_B = 4\,M_A$$

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 1 bar pressure will be released when 0.15 g of aluminium reacts?

Ans. $$\underset{2\times 27}{2Al} + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + \underset{3\times 22{,}400\text{ mL}}{3H_2}$$

H_2 produced at STP from 0.15 g Al $= \dfrac{3\times 22400\times 0.15}{54}$ mL $= 186.7$ mL

$$\underset{\text{(STP)}}{\frac{P_1V_1}{T_1}} = \underset{\text{(Required conditions)}}{\frac{P_2V_2}{T_2}}$$

We know 1 bar = 0.987 atm

$$\frac{1\ \text{atm}\times 186.7\ \text{mL}}{273\ \text{K}} = \frac{0.987\ \text{atm}\times V_2}{293\ \text{K}} \text{ or } V_2 = 203\ \text{mL}$$

7. 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. $$p = \frac{n}{V}RT = \frac{wRT}{MV}$$

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

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

$P_{total} = 0.55 + 0.27 = 0.82$ atm

8. 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 oxygen at 0.7 bar are introduced in a l L vessel at 27°C?

Ans. Partial pressure of H_2 in l L vessel

$P_1 = 0.8$ bar, $V_1 = 0.5$ L, $P_2 = ?$ $V_2 = 1.0$ L

Temperature remaining constant $P_1P_1 = P_2 P_2$

or, $0.8 \text{ bar} \times 0.5 \text{ L} = P_2 \times 1.0 \text{ L}$ or, $P_{H_2} = 0.40$ bar

Partial pressure of O_2 in 1 L vessel

$P'_1 = 0.7$ bar, $V'_1 = 2.0$ L, $P'_2 = ?$ $V'_2 = 1.0$ L

Temperature remaining constant

$P'_1V'_1 = P'_2 V'_2$

or, $0.7 \text{ bar} \times 2.0 \text{ L} = P'_2 \times 1.0 \text{ L}$ or, $P'_2 = 1.4$ bar or, $P_{O_2} = 1.4$ bar

Total Pressure = $P_{H_2} + P_{O_2} = 0.4 \text{ bar} + 1.4 \text{ bar} = 1.8$ bar

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

Ans. $d = \dfrac{MP}{RT}$

For the same gas at different temperature and pressure $\dfrac{d_1}{d_2} = \dfrac{P_1}{T_1} \times \dfrac{T_2}{P_2}$

Given that

$d_1 = 5.46$ g/dm³, $T_1 = 27°C = 300$ K, $P_1 = 2$ bar

At S.T.P.

$d_2 = ?$, $T_2 = 0°C = 273$ K, $P_2 = 1$ bar

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

10. 340.5 mL of phosphorus vapour weighs 0.0625 g at 546°C temperature and 0.1 bar pressure. What is the molar mass of phosphorus?

Ans. Volume at 0°C and 1 bar pressure $= \dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}$ or, $\dfrac{0.1 \times 340.5}{546 + 273} = \dfrac{1 \times V_2}{273}$

$$V_2 = \frac{0.1 \times 340.5 \times 273}{(546 - 273) \times 1} \text{ or, } \quad V_2 = 11.35 \text{ mL.}$$

$\therefore$ Weight of 11.35 mL of vapour at 0°C and 1 bar pressure = 0.0625 g

$\therefore$ Weight of 22,700 mL of vapour at 0°C and 1 bar pressure

$$= \frac{0.0625}{11.35} \times 22,700 = 125 \text{ g}$$

$\therefore$ Molar mass = 125 g/mol

11. 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 realised his mistake, and using a pyrometre, he found the temperature of the flask was 477°C. What fraction of air would have been expelled out?

Ans. Let us assume that volume of vessel = V cm^3

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

Since $\frac{V_1}{T_1} = \frac{V_2}{T_2} or, \frac{V}{300} = \frac{V_2}{750}$ or, V_2 = 2.5 V

∴ Volume expelled = 2.5 V – V = 1.5 V

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

12. Calculate the temperature of 4.0 mol of a gas occupying 5 dm^3 at 3.32 bar (*R* = 0.083 bar dm^3/K/mol)

Ans. $PV = nRT$ or, $T = \frac{PV}{nR} = \frac{3.32\text{ bar} \times 5\text{ dm}^3}{4.0\text{mol} \times 0.083\text{ bar dm}^3\text{ / K / mol}}$ = 50 K

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

Ans. Gram molecular mass of nitrogen = 28 g

28 g of nitrogen contains 6.023 × 10^{23} molecules.

∴ 1.4 g of nitrogen would contain $= \frac{6.023 \times 10^{23} \times 1.4}{28}$ = 3.011 × 10^{22} molecules

Since 1 atom of nitrogen consists of 7 electrons then one nitrogen molecule would contain 14 electrons

3.011 × 10^{22} molecules of nitrogen would contain = 14 × 3.011 × 10^{22} = 4.215 × 10^{23} electrons.

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

Ans. One Avogadro number of wheat grains= 6.022 ×10^{23} grains

Time required to distribute 10^{10} grains = 1 s

Time required to distribute 6.022 × 10^{23} grains $= \frac{6.023 \times 10^{23}}{10^{10}}$s = 6.022× 10^{13} s

$= \frac{6.022 \times 10^{13}}{60 \times 60 \times 24 \times 365}$ years = 19. 096 × 10^5 years.

15. Calculate the total pressure in a mixture of 8 g of oxygen and 4 g of hydrogen confined in a vessel of 1 dm³ at 27°C. *R* = 0.083 bar dm³/K/mol.

Ans. Molar mass of O_2 = 32 g/mol

∴ 8 g of O_2 = 8/32 g/mol = 0.25 mol

Molar mass of H_2 = 2 g/mol

∴ 4 g of H_2 = 4/2 g/mol = 2 mol

∴ Total number of moles (n) = 2 mol + 0.25 mol = 2.25 mol

V = 1 dm³, T = 27° C = 300 K, R = 0.083 bar dm³/K/mol

$$PV = n\,RT \text{ or, } P = \frac{nRT}{V} = \frac{2.25 \text{ mol} \times 0.083 \text{ bar dm}^3/\text{K}/\text{mol} \times 300 \text{ K}}{1 \text{ dm}^3}$$

= 56.025 bar

16. 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/m³ and *R* = 0.083 bar dm³/K/ mol)

Ans. Radius of balloon = 10 m

Volume of the balloon = $4/3\pi r^3$ = 4/3 × 22/7 × (10 m)³ = 4190.5 m³

Volume of He filled at 1.66 bar and 27°C = 4190.5 m³

Calculation of mass of He

$PV = nRT = (w/M)\,RT$

$$\text{or, } P = \frac{nRT}{V} = \frac{2.25 \text{ mol} \times 0.083 \text{ bar dm}^3/\text{K}/\text{mol} \times 300 \text{ K}}{1 \text{ dm}^3} = 1117.5 \text{ kg}$$

∴ Total mass of (balloon + He) = 100 + 1117.5 = 1217.5 kg

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

or, 4190.5 m³×1.2 kg/m³ = 5028.6 kg

payload = 5028.6 – 1217.5 kg = 3811.1 kg

17. Calculate the volume occupied by 8.8 g of CO_2 at 31.1°C and 1 bar pressure (*R* = 0.083 bar L/K/mol)

Ans. $$PV = nRT \text{ or, } V = \frac{nRT}{P} = \frac{wRT}{MP} \left(\text{because } n = \frac{w}{M}\right)$$

$$V = \frac{8.8\text{g}}{44 \text{ g/mol}} \times \frac{0.083 \text{ bar L/K/mol} \times (273 + 31.1) \text{ K}}{1 \text{ bar}} = 5.05 \text{ L}$$

18. 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?

Ans. Let the molar mass of the gas be M_x

We know $P_1 = P_2$ and $V_1 = V_2$ or, $P_1V_1 = P_2V_2$

or, $n_1RT_1 = n_2RT_2$ or, $n_1T_1 = n_2T_2$ or, $\frac{W_1}{M_1}T_1 = \frac{W_2}{M_2}T_2$

$$\therefore \frac{2.9}{M_x}\times(95+273) = \frac{0.184}{2}\times(17+273)$$

$$\text{or, } M_x = \frac{2.9\times368\times2}{0.184\times290} = 40\text{ g/mol}$$

19. A mixture of hydrogen and oxygen at one bar pressure contains 20% by weight of hydrogen. Calculate the partial pressure of hydrogen.

Ans. $nH_2 = 20/2 = 10$ mol, $nO_2 = 80/32 = 2.5$ mol

$$P_{H_2} = \frac{^{n}H_2}{^{n}H_2 + ^{n}O_2}\times P_{Total} = \frac{10}{10+2.5}\times1\text{ bar} = 0.8\text{ bar.}$$

20. What would he the SI unit for the quantity PV^2T^2/n?

Ans. $$\frac{N/m^2\times(m^3)^2\times K^2}{mol} = N\,m^4\,K/mol$$

21. In terms of Charles's law explain why –273°C is the lowest temperature?

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

22. 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?

Ans. 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.

23. Explain the significance of the van der Waals' parametres?

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.