Question 5.1:

What will be the minimum pressure required to compress 500 dm^3 of air at 1 bar to 200 dm^3 at 30°C?

Answer

Given,

Initial pressure, p_1 = 1 bar

Initial volume, V_1 = 500 dm^3

Final volume, V_2 = 200 dm^3

Since the temperature remains constant, the final pressure (p_2) can be calculated using Boyle's law.

According to Boyle's law,

$$p_1V_1 = p_2V_2$$

$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$

$$= \frac{1 \times 500}{200} \text{ bar}$$

$$= 2.5 \text{ bar}$$

Therefore, the minimum pressure required is 2.5 bar.

Question 5.2:

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

Given,

Initial pressure, p_1 = 1.2 bar

Initial volume, V_1 = 120 mL

Final volume, V_2 = 180 mL

Since the temperature remains constant, the final pressure (p_2) can be calculated using Boyle's law.

According to Boyle's law,

$$p_1V_1 = p_2V_2$$

$$p_2 = \frac{p_1V_1}{V_2}$$

$$= \frac{1.2 \times 120}{180} \text{ bar}$$

$$= 0.8 \text{ bar}$$

Therefore, the pressure would be 0.8 bar.

Question 5.3:

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

Answer

The equation of state is given by,

$pV = nRT$ ……….. (i)

Where,

p → Pressure of gas

V → Volume of gas

n → Number of moles of gas

R → Gas constant

T → Temperature of gas

From equation (i) we have,

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

Replacing n with $\frac{m}{M}$, we have

$$\frac{m}{MV} = \frac{p}{RT} \text{ ....(ii)}$$

Where,

m → Mass of gas

M → Molar mass of gas

But, $\frac{m}{V} = d$ (d = density of gas)

Thus, from equation (ii), we have

$$\frac{d}{M} = \frac{p}{RT}$$

$$\Rightarrow d = \left(\frac{M}{RT}\right)p$$

Molar mass (M) of a gas is always constant and therefore, at constant temperature $(T), \frac{M}{RT}$ = constant.

$$d = (\text{constant})\,p$$

$$\Rightarrow d \propto p$$

Hence, at a given temperature, the density (d) of gas is proportional to its pressure (p)

Question 5.4:

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?

Answer

Density (d) of the substance at temperature (T) can be given by the expression,

$d = \frac{Mp}{RT}$

Now, density of oxide (d_1) is given by,

$$d_1 = \frac{M_1 p_1}{RT}$$

Where, M_1 and p_1 are the mass and pressure of the oxide respectively.

Density of dinitrogen gas (d_2) is given by,

$$d_2 = \frac{M_2 p_2}{RT}$$

Where, M_2 and p_2 are the mass and pressure of the oxide respectively.

According to the given question,

$$d_1 = d_2$$

$$\therefore M_1 p_1 = M_2 p_2$$

Given,

$$p_1 = 2\,bar$$

$$p_2 = 5\,\text{bar}$$

Molecular mass of nitrogen, M_2 = 28 g/mol

$$\text{Now, } M_1 = \frac{M_2 p_2}{p_1} = \frac{28 \times 5}{2} = 70 \text{ g/mol}$$

Hence, the molecular mass of the oxide is 70 g/mol.

Question 5.5:

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 a relationship between their molecular masses.

Answer

For ideal gas A, the ideal gas equation is given by,

$$p_A V = n_A RT \text{ ....(i)}$$

Where, p_A and n_A represent the pressure and number of moles of gas A.

For ideal gas B, the ideal gas equation is given by,

$$p_B V = n_B RT \text{ ....(ii)}$$

Where, p_B and n_B represent the pressure and number of moles of gas B.

[V and T are constants for gases A and B]

From equation (i), we have

$$p_A V = \frac{m_A}{M_A} RT \Rightarrow \frac{p_A M_A}{m_A} = \frac{RT}{V} \text{ ....(iii)}$$

From equation (ii), we have

$$p_B V = \frac{m_B}{M_B} RT \Rightarrow \frac{p_B M_B}{m_B} = \frac{RT}{V} \text{ ....(iv)}$$

Where, M_A and M_B are the molecular masses of gases A and B respectively.

Now, from equations (iii) and (iv), we have

$$\frac{p_A M_A}{m_A} = \frac{p_B M_B}{m_B} \text{ ....}(v)$$

Given,

$m_A = 1\text{ g}$

$p_A = 2\text{ bar}$

$m_B = 2\text{g}$

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

(Since total pressure is 3 bar)

Substituting these values in equation (v), we have

$$\frac{2 \times M_A}{1} = \frac{1 \times M_B}{2}$$

$$\Rightarrow 4M_A = M_B$$

Thus, a relationship between the molecular masses of A and B is given by

$4M_A = M_B$.

Question 5.6:

The drain cleaner, Drainex contains small bits of aluminum 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 aluminum reacts?

Answer

The reaction of aluminium with caustic soda can be represented as:

$$2Al + 2NaOH + 2H_2O \longrightarrow 2NaAlO_2 + 3H_2$$

$$2 \times 27\text{g} \qquad\qquad\qquad\qquad 3 \times 22400\text{ mL}$$

At STP (273.15 K and 1 atm), 54 g (2 × 27 g) of Al gives 3 × 22400 mL of H_2..

$\therefore$ 0.15 g Al gives $\frac{3 \times 22400 \times 0.15}{54}\text{ mL of } H_2$ i.e., 186.67 mL of H_2.

At STP,

$p_1 = 1\text{ atm}$

$V_1 = 186.67\text{ mL}$

$T_1 = 273.15\text{ K}$

Let the volume of dihydrogen be V_2 at p_2 = 0.987 atm (since 1 bar = 0.987 atm) and T_2 = 20°C = (273.15 + 20) K = 293.15 K..

Now,

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

$$\Rightarrow V_2 = \frac{p_1V_1T_2}{p_2T_1}$$

$$= \frac{1\times 186.67\times 293.15}{0.987\times 273.15}$$

$$= 202.98\,\text{mL}$$

$$= 203\ \text{mL}$$

Therefore, 203 mL of dihydrogen will be released.

Question 5.7:

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 dm^3 flask at 27 °C ?

Answer

It is known that,

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

For methane (CH_4),

$$p_{CH_4} = \frac{3.2}{16}\times\frac{8.314\times 300}{9\times 10^{-3}}\left[\begin{array}{l}\text{Since } 9\ \text{dm}^3 = 9\times 10^{-3}\,\text{m}^3\\ 27^\circ\text{C} = 300\text{K}\end{array}\right]$$

$$= 5.543\times 10^4\,\text{Pa}$$

For carbon dioxide (CO_2),

$$p_{CO_2} = \frac{4.4}{44}\times\frac{8.314\times 300}{9\times 10^{-3}}$$

$$= 2.771\times 10^4\ \text{Pa}$$

Total pressure exerted by the mixture can be obtained as:

$$p = p_{CH_4} + p_{CO_2}$$

$$= \left(5.543\times 10^4 + 2.771\times 10^4\right)\ \text{Pa}$$

$$= 8.314\times 10^4\ \text{Pa}$$

Hence, the total pressure exerted by the mixture is 8.314×10^4 Pa.

Question 5.8:

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 1L vessel at 27°C?

Answer

Let the partial pressure of H_2 in the vessel be p_{H_2}.

Now,

$p_1 = 0.8 \text{ bar}$ $\quad p_2 = p_{H_2} = ?$

$V_1 = 0.5 \text{ L}$ $\quad V_2 = 1 \text{ L}$

It is known that,

$$p_1V_1 = p_2V_2$$
$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$
$$\Rightarrow p_{H_2} = \frac{0.8 \times 0.5}{1}$$
$$= 0.4 \text{ bar}$$

Now, let the partial pressure of O_2 in the vessel be p_{O_2}.

Now,

$p_1 = 0.7 \text{ bar}$ $\quad p_2 = p_{O_2} = ?$

$V_1 = 2.0 \text{ L}$ $\quad V_2 = 1 \text{ L}$

$$p_1V_1 = p_2V_2$$
$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$
$$\Rightarrow p_{O_2} = \frac{0.7 \times 20}{1}$$
$$= 0.4 \text{ bar}$$

Total pressure of the gas mixture in the vessel can be obtained as:

$$p_{\text{total}} = p_{H_2} + p_{O_2}$$
$$= 0.4 + 1.4$$
$$= 1.8 \text{ bar}$$

Hence, the total pressure of the gaseous mixture in the vessel is 1.8 bar.

Question 5.9:

Density of a gas is found to be 5.46 g/dm^3 at 27 °C at 2 bar pressure. What will be its density at STP?

Answer

Given,

$$d_1 = 5.46 \text{ g/dm}^3$$
$$p_1 = 2 \text{ bar}$$
$$T_1 = 27°\text{C} = (27 + 273)\text{K} = 300 \text{ K}$$
$$p_2 = 1 \text{ bar}$$
$$\text{T}_2 = 273 \text{ K}$$
$$d_2 = ?$$

The density (d_2) of the gas at STP can be calculated using the equation,

$$d = \frac{Mp}{\text{R}T}$$

$$\therefore \frac{d_1}{d_2} = \frac{\frac{Mp_1}{\text{R}T_1}}{\frac{Mp_2}{\text{R}T_2}}$$

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

$$\Rightarrow d_2 = \frac{p_2 T_1 d_1}{p_1 T_2}$$
$$= \frac{1 \times 300 \times 5.46}{2 \times 273}$$
$$= 3 \text{ g dm}^{-3}$$

Hence, the density of the gas at STP will be 3 g dm^{-3}.

Question 5.10:

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?

Answer

Given,

p = 0.1 bar

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

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

$$= \frac{750V}{300}$$

$$= 2.5\ \text{V}$$

Therefore, volume of air expelled out = 2.5 *V* – *V* = 1.5 *V*

Hence, fraction of air expelled out $= \frac{1.5V}{2.5V} = \frac{3}{5}$

Question 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^{-1} mol^{-1}).

Answer

Given,

n = 4.0 mol

V = 5 dm^3

p = 3.32 bar

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

The temperature (T) can be calculated using the ideal gas equation as:

$$pV = n\text{R}T$$

$$\Rightarrow T = \frac{pV}{n\text{R}}$$

$$= \frac{3.32 \times 5}{4 \times 0.083}$$

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

Hence, the required temperature is 50 K.

Question 5.13:

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

Answer

Molar mass of dinitrogen (N_2) = 28 g mol^{-1}

Thus, 1.4 g of $N_2 = \frac{1.4}{28} = 0.05 \text{ mol}$

$= 0.05 \times 6.02 \times 10^{23}$ number of molecules

$= 3.01 \times 10^{23}$ number of molecules

Now, 1 molecule of N_2 contains 14 electrons.

Therefore, 3.01×10^{23} molecules of N_2 contains = 14 × 3.01 × 1023

$= 4.214 \times 10^{23}$ electrons

Question 5.14:

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

Answer

Avogadro number = 6.02×10^{23}

Thus, time required

$$= \frac{6.02 \times 10^{23}}{10^{10}} \text{ s}$$

$$= 6.02 \times 10^{23} \text{ s}$$

$$= \frac{6.02 \times 10^{23}}{60 \times 60 \times 24 \times 365} \text{ years}$$

$$= 1.909 \times 10^{6} \text{ years}$$

Hence, the time taken would be 1.909×10^{6} years.

Question 5.15:

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 bar dm^3 K^{-1} mol^{-1}.

Answer

Given,

Mass of dioxygen (O_2) = 8 g

Thus, number of moles of $O_2 = \frac{8}{32} = 0.25 \text{ mole}$

Mass of dihydrogen (H_2) = 4 g

Thus, number of moles of $H_2 = \frac{4}{2} = 2\,\text{mole}$

Therefore, total number of moles in the mixture = 0.25 + 2 = 2.25 mole

Given,

V = 1 dm^3

n = 2.25 mol

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

T = 27°C = 300 K

Total pressure (p) can be calculated as:

$pV = n\text{R}T$

$$\Rightarrow p = \frac{n\text{R}T}{V}$$

$$= \frac{225 \times 0.083 \times 300}{1}$$

$$= 56.025 \text{ bar}$$

Hence, the total pressure of the mixture is 56.025 bar.

Question 5.16:

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 bar dm^3 K^{-1} mol^{-1}).

Answer

Given,

Radius of the balloon, r = 10 m

$\therefore$ Volume of the balloon $= \frac{4}{3}\pi r^3$

$$= \frac{4}{3} \times \frac{22}{7} \times 10^3$$

$$= 4190.5\,\text{m}^3\,(\text{approx})$$

Thus, the volume of the displaced air is 4190.5 m^3.

Given,

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

Then, mass of displaced air = 4190.5 × 1.2 kg

= 5028.6 kg

Now, mass of helium (*m*) inside the balloon is given by,

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

Here,

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

$p = 1.66 \text{ bar}$

$V = \text{Volume of the balloon}$

$\quad = 4190.5 \text{ m}^3$

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

$T = 27°\text{C} = 300\text{K}$

$$\text{Then, } m = \frac{4 \times 10^{-3} \times 1.66 \times 4190.5 \times 10^3}{0.083 \times 300}$$

$$= 1117.5 \text{ kg (approx)}$$

Now, total mass of the balloon filled with helium = (100 + 1117.5) kg

= 1217.5 kg

Hence, pay load = (5028.6 – 1217.5) kg

= 3811.1 kg

Hence, the pay load of the balloon is 3811.1 kg.

Question 5.17:

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^{-1} mol^{-1}.

Answer

It is known that,

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

$$\Rightarrow V = \frac{mRT}{Mp}$$

Here,

m = 8.8 g

R = 0.083 bar LK^{-1} mol^{-1}

T = 31.1°C = 304.1 K

M = 44 g

p = 1 bar

$$\text{Thus, volume}(V) = \frac{8.8 \times 0.083 \times 304.1}{44 \times 1}$$
$$= 5.04806 \text{ L}$$
$$= 5.05 \text{ L}$$

Hence, the volume occupied is 5.05 L.

Question 5.18:

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?

Answer

Volume (V) occupied by dihydrogen is given by,

$$V = \frac{m}{M}\frac{\text{R}T}{p}$$
$$= \frac{0.184}{2} \times \frac{\text{R} \times 290}{p}$$

Let M be the molar mass of the unknown gas. Volume (V) occupied by the unknown gas can be calculated as:

$$V = \frac{m}{M}\frac{\text{R}T}{p}$$
$$= \frac{2.9}{M} \times \frac{\text{R} \times 368}{p}$$

According to the question,

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

Hence, the molar mass of the gas is 40 g mol^{-1}.