Chapter – 5

States of Matter

NCERT Back Exercises :

Ques 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?

Ans 5.1: 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,

$$P_1V_1 = P_2V_2$$

$$P_2 = \frac{P_1V_1}{V_2}$$

$$= \frac{1\times500}{200} \text{ bar}$$

= 2.5 bar

Therefore, the minimum pressure required is 2.5 bar.

Ques 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?

Ans 5.2: 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.

Ques 5.3: Using the equation of state *pV* = *n*R*T*; show that at a given temperature density of a gas is proportional to gas pressure *p*.

Ans 5.3:

The equation of state is given by,

$pV = n\mathrm{R}T$…………………………………………………………………………………(i)

Where,

$p \rightarrow$ Pressure of gas

$V \rightarrow$ Volume of gas

$n \rightarrow$ Number of moles of gas

R $\rightarrow$ Gas constant

$T \rightarrow$ Temperature of gas

From equation (i) we have,

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

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

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

Where,

$m \rightarrow$ Mass of gas

$M \rightarrow$ Molar mass of gas

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

Thus, from equation (ii), we have

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

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

$d \propto p$

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

Ques 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?

Ans 5.4: 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 di nitrogen 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,

$d_1 = d_2$

$M_1 p_1 = M_2 p_2$

Given,

$p_1 = 2\ bar$

$p_2 = 5\ bar$

Molecular mass of nitrogen, M_2 = 28 g/mol

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.

Ques 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.

Ans 5.5: For ideal gas A, the ideal gas equation is given by,

$p_A V = n_A RT \ldots\ldots\ldots\ldots\ldots\ldots (i)$

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,

$p_B V = n_B RT \ldots\ldots\ldots\ldots\ldots\ldots (ii)$

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

$p_A V = \frac{m_A}{M_A} RT \implies \frac{p_A M_A}{m_A} = \frac{RT}{V} \ldots\ldots\ldots\ldots\ldots\ldots (iii)$

From equation (ii), we have

$p_B V = \frac{m_B}{M_B} RT \implies \frac{p_B M_B}{m_B} = \frac{RT}{V} \ldots\ldots\ldots\ldots\ldots\ldots (iv)$ Where, M_A and M_B are the molecular masses of gases A and B respectively.

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

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

Given,

$m_A = 1g$, $m_B = 2\ g$

$p_A = 1\ bar$, $p_B = (3 - 2) = 1\ bar$

(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 $4M_A = M_B$.

Ques 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?

Ans 5.6: The reaction of aluminium with caustic soda can be represented as:

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

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

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

At STP,

$p_1 = 1\ atm$

$V_1 = 186.67\ mL$

$T_1 = 273.15\ K$

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,

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

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

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

= 202.98 mL

= 203 mL

Therefore, 203 mL of dihydrogen will be released.

Ques 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 ?

Ans 5.7 : It is known that,

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

For methane (CH_4),

$$p_{CH_4} = \frac{3.2}{16} \times \frac{8.314 \times 300}{9 \times 10^{-3}} \qquad [Since\ 9\ dm^3 = 9 \times 10^{-3} m^3]$$

$$= 5.543 \times 10^4 Pa$$

For carbon dioxide (CO_2),

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

$$= 2.771 \times 10^4 Pa$$

Total pressure exerted by the mixture can be obtained as:

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

$$= (5.543 \times 10^4 Pa + 2.771 \times 10^4 Pa)$$

$$= 8.314 \times 10^4 Pa$$

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

Ques 5.8: What will be the pressure of the gaseous mixture when 0.5 L of H2 at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1L vessel at 27°C?

Ans 5.8: Let the partial pressure of H_2 in the vessel be p_{H_2} .

Now,

$$p_1 = 0.8\ bar$$

$$p_2 = p_{H_2} = ?$$

$$V_1 = 0.5\ L$$

$$V_2 = 1\ L$$

It is known that,

$$p_1V_1 = p_2V_2$$

$$p_2 = \frac{P_1V_1}{V_2}$$

$p_{H_2} = \frac{0.8 \times 0.5}{1} = 0.4$ bar

Now, let the partial pressure of O2 in the vessel be p_{CO_2} .

Now,

$p_1 = 0.7\ bar$

$p_2 = p_{CO_2} = ?$

$V_1 = 2.0\ L$

$V_2 = 1\ L$

$p_1V_1 = p_2V_2$

$p_2 = \frac{P_1V_1}{V_2}$

$p_{CO_2} = \frac{0.7 \times 20}{1} = 1.4$ bar

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

$p_{total} = p_{H_2} + p_{CO_2}$

$= 0.4 + 1.4$

$= 1.8\ bar$

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

Ques 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?

Ans 5.9: Given,

$d_1 = 5.46\ \text{g/dm}^3$

$p_1 = 2\ bar$

$T_1 = 27^{\circ}C = (27 + 273)K = 300K$

$p_2 = 1\ bar$

$T_2 = 273K$

$d_2 = ?$

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

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

$\frac{d_1}{d_2} = \frac{\frac{Mp_1}{RT_1}}{\frac{Mp_2}{RT_2}}$

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

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

$= \frac{1 \times 300 \times 5.46}{2 \times 273}$

$= 3 \text{ g } dm^{-3}$

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

Ques 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?

Ans 5.10: Given,

p = 0.1 bar

V = 34.05 mL = 34.05×10^{-3} L = 34.05×10^{-3} dm^3

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

T = 546°C = (546 + 273) K = 819 K

The number of moles (n) can be calculated using the ideal gas equation as:

$pV = nRT$

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

$= \frac{0.1 \times 34.05 \times 10^{-3}}{0.083 \times 819}$

$= 5.01 \times 10^{-5}$ mol

Therefore, molar mass of phosphorus $\frac{0.0625}{5.01 \times 10^{-5}}$ = 1247.5 g mol^{-1}

Hence, the molar mass of phosphorus is 1247.5 g mol^{-1} .

Ques 5.11: 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 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 5.11: Let the volume of the round bottomed flask be V.

Then, the volume of air inside the flask at 27° C is V.

Now,

$V_1 = V$

$T_1 = 27°C = 300$ K

$V_2 = ?$

$T_2 = 477°$ C $= 750$ K

According to Charles's law,

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

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

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

$$= 2.5\ V$$

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

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

Ques 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}).

Ans 5.12: 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:

$pV = nRT$

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

$$= \frac{3.32 \times 5}{4 \times 0.083} = 50\ K$$

Hence, the required temperature is 50 K.

Ques 5.13: Calculate the total number of electrons present in 1.4 g of dinitrogen gas.

Ans 5.13:

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

Thus , 1.4 g of $N_2 = \frac{1.4}{28}$ = 0.05 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 = $14 \times 3.01 \times 10^{23}$

= 4.214×10^{23} electrons

Ques 5.14: How much time would it take to distribute one Avogadro number of wheat grains, if 10^{10} grains are distributed each second?

Ans 5.14: Avogadro number = 6.02×10^{23}

Thus, time required

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

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

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

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

Ques 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}.

Ans 5.15: Given,

Mass of dioxygen (O_2) = 8 g

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

Mass of dihydrogen (H_2) = 4 g

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

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:

$pV = nRT$

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

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

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

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

Ques 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}).***

Ans 5.16: Given,

Radius of the balloon, r = 10 m

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

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

$= 4190.5\ m^3$(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,

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

p = 1.66 bar

V = Volume of the ballon = 4190.5 m^3

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

T = 27°C = 300 K

Then,

$$m = \frac{4\times10^{-3}\times1.66\times4190.5\times10^3}{0.083\times300}$$

= 1117.5 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.

Ques 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}.

Ans 5.17:

It is known that,

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

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

Here,

m = 8.8 g

R = 0.083 bar L K^{-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.

Ques 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?

Ans 5.18:

Volume (*V*) occupied by dihydrogen is given by,

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

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

Let M be the molar mass of the unknown gas.

Volume (V) occupied by the unknown gas can be calculated as:

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

$$= \frac{2.9}{M} \times \frac{R \times 368}{p}$$

According to the Question ,

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

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

$$M = \frac{2.9{\times}368{\times}2}{0.148{\times}\ 290}$$

$$= 40\ g\ mol^{-1}$$

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

Ques 5.19: A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

Ans 5.19:

Let the weight of dihydrogen be 20 g and the weight of dioxygen be 80 g.

Then, the number of moles of dihydrogen,

$n_{H_2} = \frac{20}{2} = 10$ moles and

the number of moles of dioxygen,

$n_{CO_2} = \frac{80}{32} = 2.5$ moles

Given,

Total pressure of the mixture, $p_{total} = 1$ bar

Then, partial pressure of dihydrogen,

$$p_{H_2} = \frac{n_{H_2}}{n_{H_2}+n_{O_2}} \times p_{total}$$

$$= \frac{10}{10+2.5} \times 1 = 0.8 \text{ bar}$$

Hence, the partial pressure of dihydrogen is 0.8 bar .

Ques 5.20: What would be the SI unit for the quantity $\frac{pV^2T^2}{n}$?

Ans 5.20: The SI unit for pressure, *p* is Nm^{-2}.

The SI unit for volume, *V* is m^3.

The SI unit for temperature, *T* is K.

The SI unit for the number of moles, *n* is mol.

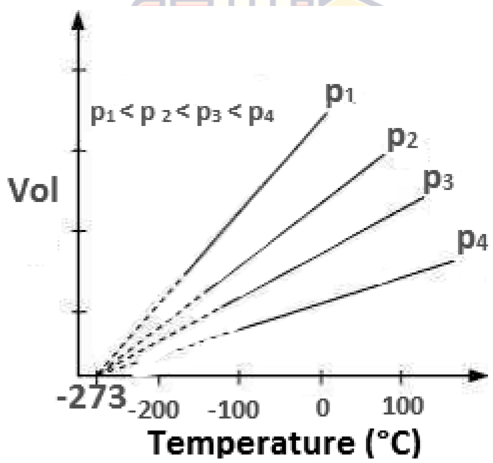
Therefore, the SI unit for quantity $\frac{pV^2T^2}{n}$ is given by,

$$= \frac{(Nm^{-2})(m^3)^2(K)^2}{mol}$$

$$= Nm^4K^2\ mol^{-1}$$

Ques 5.21: In terms of Charles' law explain why –273°C is the lowest possible temperature.

Ans 5.21: Charles' law states that at constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature.



It was found that for all gases (at any given pressure), the plots of volume vs. temperature (in °C) is a straight line. If this line is extended to zero volume, then it intersects the temperature-axis at – 273°C. In other words, the volume of any gas at – 273°C is zero. This is because all gases get liquefied before reaching a temperature of – 273°C.

Hence, it can be concluded that – 273°C is the lowest possible temperature.

Ques 5.22: 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 5.22: Higher is the critical temperature of a gas, easier is its liquefaction. This means that the intermolecular forces of attraction between the molecules of a gas are directly proportional to its critical temperature. Hence, intermolecular forces of attraction are stronger in the case of CO_2.

Ques 5.23: Explain the physical significance of Van der Waals parameters.

Ans 5.23: Physical significance of 'a':

'a' is a measure of the magnitude of intermolecular attractive forces within a gas.

Physical significance of 'b':

'b' is a measure of the volume of a gas molecule.