Chapter 5. States of Matter

Question-1

What will be the minimum pressure required to compress 500 dm^3 of air at 1 bar to 200 dm^3 at 30^0 c?

Solution:

P_1 = 1 bar

P_2 = ?

V_1 = 500 dm^3

V_2 = 200dm^3

$P_1V_1 = P_2v_2$

$1 \times 500 = P_2 \times 200$ or

$P_2 = \frac{1\times 500}{200} = 2.5$ 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?

Solution:

Let V_1 = 120 ml

V_2 = 180 ml

P_1 = 1.2 bar

P_2 = ?

$P_1V_1 = P_2V_2$

$1.2 \times 120 = P_2 \times 180$

$P_2 = \frac{120}{180} \times 1.2$

P_2 = 0.8 bar.

Question-3

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

Solution:

PV = nRT

(or) P (Pressure of gas) = $\frac{n.RT}{V}$

We know $\frac{1}{V} = \rho$ (density)

$\therefore$ P = n. ρ RT

$\therefore$ Pressure of gas (P) α density of the gas (d).

Question-4

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

Solution:

Density of nitrogen, $\rho = \frac{Pm}{RT}$

$= P = 5 \times 0.987$ atm, $m = 28$

$= \rho = \frac{5 \times 0.987 \times 28}{0.0821 \times 273}$

Density of gaseous oxide

$\frac{5 \times 0.987 \times 28}{0.0821 \times 273} = \frac{2 \times 0.987 \times x}{0.0821 \times 273}$

$x = \frac{0.0821 \times 273 \times 5 \times 0.987 \times 28}{0.0821 \times 273 \times 2 \times 0.987}$

or $x = \frac{5 \times 28}{2} = 70$ g/mol.

Question-5

Pressure of 1g of an ideal gas A at 27^0 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

Solution:

Amount of ideal gas, A = W_A = 1 g

Its pressure, P_A = 2bar

Amount of another gas, B = W_B = 2 g

Total pressure of the flask = 3 bar

$\therefore$ Partial Pressure of gas B in the gas mixture = 3-2 = 1 bar

PV = nRT

$= \frac{W.RT}{M}$

For gas A, $P_A = \frac{W_A\ .R.T}{M_A V}$

$2 = \frac{1g.RT}{M_A.V}$ ....(i)

For gas B,

$P_B = \frac{W_B \times RT}{M_B.V}$

$1 = \frac{2gR.T}{M_A.V}$ --------(ii)

Divide equation (i) by (ii)

$\frac{2g}{1g} = \frac{1g.RT}{M_A \times V} \times \frac{M_B.V}{2g.RT}$

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

$4M_A = M_B$

The molecular mass of M_b is four times of M_a.

Question-6

The drain cleaner, Drainex contains small bits of aluminium which react with caustic soda to produce hydrogen. What volume of hydrogen at 20^0c and one bar will be released when 0.15 g of aluminium reacts?

Solution:

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

2×27 g of Al $\quad$ 3×22.4 litre of H_2

= 54 g

54g Al gives = 3×22.4 litre H_2 at N.T.P.

0.15 g Al gives = $\frac{0.15 \times 3 \times 22.4}{54}$ = 0.186 litre at N.T.P.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Normal Pressure P_1 = 1 atm

P_2 = 1 bar = 0.987 atm

V_1 = 0.186L

V_2 =?

T_1 = 273K

T_2 = 273 + 20 = 293K

$$\frac{1 \times 0.186}{273} = \frac{0.987 \times V_2}{293}$$

V_2 = $\frac{1 \times 0.186 \times 293}{0.987 \times 273}$ = 0.202 litre = 202 ml.

Question-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 27^0C?

Solution:

w_{CH_4} = 3.2 g

Mol. Weight of CH_4 = 16

Moles of CH_4 = $\frac{3.2}{16}$=0.2

w_{CO_2} = 4.4 g

Mol. Weight of CO_2 = 44

Moles of CO_2 = $\frac{4.4}{44}$ = 0.1 mol

Total moles present in mixture = 0.2 + 0.1 = 0.3 mol

Volume of flask = 9 dm^3 = 9 L

Temperature = 273 + 27 = 300 K

Now for pressure of mixture = $P \times V = nRT$

$P \times 9 = 0.3 \times 0.0821 \times 300$

$P = \frac{0.3 \times 0.0821 \times 300}{9}$=0.821 atom

$0.821 \times 1.01 \times 10^5\ P_a = 8.29 \times 10^4\ P_a$.

Question-8

What will be the pressure of the gas mixture when 0.5 L of H_2 at 0.8 bars and 2.0 L of oxygen at 0.7 bar are introduced in a 1 L vessel are 27^0C?

Solution:

Volume of H_2, (V_1) 0.5 l

Pressure of H_2, (P_1) = 0.8 bar

Volume of oxygen, (V_2) = 2.0 l

Pressure of O_2, (P_2) = 0.7 bar

Total Volume of the flask ($V_1 + V_2$) = 1l

Total pressure of the gaseous mixture is P

$P (V_1 + V_2) = P_1V_1 + P_2V_2$

$P \times 1 = (0.8 \times 0.5) + (0.7 \times 2.0)$

$P \times 1 = (0.4 + 1.4)$

Total Pressure of the mixture of gases , P = 1.8/1 = 1.8 bar.

Question-9

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

Solution:

$$\frac{d_1 T_1}{P_1} = \frac{d_2 T_2}{P_2}$$

d_1 = 5.46 g/dm^3 d_2 = ?

T_1 = 273 + 27 = 300 K T_2 = 273 K

P_1 = 2 bar

$$\frac{5.46 \times 300}{2} = \frac{d_2 \times 273}{1}$$

$d_2 = \frac{5.46 \times 300}{2 \times 273}$ = 3g/dm^3.

Question-10

34.05 mL of phosphorous vapour weighs 0.00625 g at 546° C and 0.1 bar pressure. What is the molar mass of phosphorous?

Solution:

P = 0.1 bar , V = 34.05 mL = 0.03405

R = 0.0821 litre-atm K^{-1} mol^{-1}

T = 273 + 546 = 819 K

PV = nRT

$n = \frac{PV}{RT} = \frac{0.1 \times 0.03405}{.0821 \times 819}$

$= 0.00005\ mol^{-1}$

Weight of phosphorous vapour = 0.00625g

Molar mass of phosphorus = $\frac{0.00625g}{0.00005 mol^{-1}}$ = 125 g.

Question-11

A student forgot to add the reaction mixture to the round bottomed flask at 27°C but put it on the flame. After a lapse of time, he realized his mistake, using a pyrometer he found the temperature of the flask was 477°C. What fraction of air would have been expelled out?

Solution:

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

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

$T_1 = 27+273=300$ K

$T_2 = 477 + 273 = 750$ K

$\frac{V_1}{300} = \frac{V_2}{750}$

$\frac{V_1}{V_2} = \frac{300}{750} = \frac{2}{5}$

If the total volume of flask is taken as 1

Fraction of air expelled = $1 - \frac{2}{5} = \frac{3}{5}$.

Question-12

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

Solution:

PV = nRT

P = 3.32 bar, V = 5dm^3

n = 4 moles

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

$\therefore 3.32 \times 5 = 4 \times 0.083 \times T$

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

Question-13

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

Solution:

Atomic weight of Nitrogen = 14 g

weight of Nitrogen gas = 1.4 g

Number of gram Moles of Nitrogen atoms = $\frac{W}{M} = \frac{1.4}{14}$ = 0.1 mol.

No. of atoms of Nitrogen, present in 1 mol of Nitrogen gas = 6.022×10^{23} atoms

1 atom of Nitrogen contains = 7 electrons

No. of electrons, present in 1 mol of Nitrogen atoms = $7 \times 6.022 \times 10^{23}$

No. of electrons in 0.1 mol = $0.1 \times 7 \times 6.02 \times 10^{23} = 4.22 \times 10^{23}$.

Question-14

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

Solution:

10^{10} grains are distributed I s Time in seconds, taken to distribute 6.023×10^{23} grains = $\frac{6.023 \times 10^{23}}{10^{10}}$ s = $\frac{6.023 \times 10^{23}}{10^{10} \times 60 \times 60 \times 24 \times 365}$ years = 1,90,800 years.

Question-15

Calculate the total pressure in a mixture of 8 g of oxygen and 4 g of hydrogen confined in a vessel of 1 dm^3 at 27°C, R = 0.083 bar dm^3 K^{-1} mol^{-1}.

Solution:

Weight of oxygen W_{O_2} = 8 g

Moles of O_2 = $\frac{8}{32} = \frac{1}{4}$

W_{H_3} = 4 g

Moles of H_2 = $\frac{4}{2} = 2$

Total moles present in mixture = $\frac{1}{4} + 2 = \frac{9}{4}$

Volume of vessel = 1 dm^3 = 1 L

Temperature = 27°C = 27 + 273 = 300 K

Now for pressure of mixture = P × V = n × RT

P × 1 = $\frac{9}{4} \times 0.083 \times 300$

P = 56.025 bar.

Question-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 10m, mass 100kg 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}).

Solution:

Volume of balloon, V = $\frac{4}{3}\pi r^3$

r = 10 m

V = $\frac{4}{3} \times 3.14 \times 1(10)^3 = 4187 m^3$

Volume of air displaced = Volume of balloon = 4187 m^3

M = Vol × Density

Mass of displaced air = 4187 m^3 × 1.2 kg m^{-3} = 5024. 4 kg

Temp, T = 27+273=300 K

Moles of gas present, n = $\frac{PV}{RT}$

= $\frac{1.66 \times 4187}{0.083 \times 300} = 279.1 \times 10^3$ mol.

Mass = No. of moles × Molecular weight

Mass of He present = 279 × 10^3 × 4

= 1116.4 × 10^3 g

= 1116.4 kg

Mass of filled balloon = 100 + 1116.4 = 1216.4 Kg

Pay load = Mass of displaced air - Mass of balloon

= 5024.4 - 1216.4

= 3808 kg.

Question-17

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

Solution:

P = 1 bar , T = 31.1 + 273 = 304.1K

Mol. wt of CO_2 = 44

Weight of CO_2 = 8.8 g

PV = (W/m)RT

$l \times V = \frac{8.8}{44} \times 0.083 \times 304.1$

V = 5.05 L.

Question-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?

Solution:

Ideal gas equation, PV = nRT

1^{st} Case

Let the molar mass of gas (M) = M(g mol^{-1})

Mass of gas (m) = 2.9 g

No. of moles (m) = $\frac{\text{Mass}}{\text{Molarmass}} = \frac{2.9}{M}$ mol

T = 273 + 95 = 368

Pressure = P, Volume = V

PV = nRT

PV = $\frac{2.9}{M} \times R \times 368$ ....(i)

2^{nd} Case

Mass of hydrogen = 0.184 g

Number of moles of H_2 = $\frac{0.184}{2}$ mol

T = 273 + 17 = 290

PV = $\frac{0.184}{2} \times R \times 290$ ....(ii)

In both gases PV is same

From (i) and (ii)

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

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

Question-19

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

Solution:

Percentage of H_2 = 20%, say 20g

No. of moles of H_2 = $\frac{20}{2}$ = 10 moles

Mass of oxygen = (100-20) = 80 g)

No. of moles of O_2 = $\frac{80}{32}$ = 2.5 moles

Mole fraction of H_2 = $\frac{10}{(10+2.5)}$ = $\frac{10}{(12.5)}$

Partial pressure of H_2 =Total Pressure × Mole fraction of H_2

Total Pressure = 1 bar

$$= 1 \times \frac{10}{12.5} = 0.8 \text{ bar.}$$

Question-20

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

Solution:

S.I. unit = $\frac{Pa \times (m^3)^2 K^2}{mol}$ = Pa $m^6K^2mol^{-1}$.

Question-21

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

Solution:

Charles law states that Pressure, remaining constant the volume of a given mass of gas increases or decreases by 1/273 of its volume at 0^0c for every 1^0c rise or fall in temperature .

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

If $t = -273^0$c

$V_t = V_0 \left(1 - \frac{273}{273}\right)$
$= 0$

Thus at -273^0c any volume of a given mass of gas becomes zero. Thus -273^0c is said to be limiting temperature and it is equal to OK (absolute scale).

Question-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?

Solution:

CO_2 can more easily be liquified than Methane which shows that intermolecular forces in CO_2 are more stronger than in Methane, because in CO_2 polarity exists.

Question-23

Explain the physical significance of van der Waals parameters.

Solution:

Vander Waals equation of state for mol of gas = $\left(P + \frac{a}{V^2}\right)$(v-b) = RT

For n moles of gas = $\left(P + \frac{n^2a}{V^2}\right)$(v-nb) = RT

The constant 'a' tells the magnitude of attractive forces between the molecules of the gas. In other words larger the value of 'a', greater the intermolecular attraction.

The constant 'b' relates to non compressible volume of gas, as gas molecules occupy some volume, however small may be, at high pressures and low temperatures.∴Actual volume of gas = Available volume - Non compressible volume = (V - b).