DCAM classes
Dynamic Classes for Academic Mastery

Chemistry

(Chapter 5)(States of Matter)

XI

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* = *n*R*T*; 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:

$$\underset{2 \times 27\text{g}}{2Al} + 2NaOH + 2H_2O \longrightarrow 2NaAlO_2 + \underset{3 \times 22400 \text{ mL}}{3H_2}$$

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),

$$p_{CH_4} = \frac{3.2}{16} \times \frac{8.314 \times 300}{9 \times 10^{-3}} \left[\begin{array}{l} \text{Since } 9 \text{ dm}^3 = 9 \times 10^{-3} \text{ m}^3 \\ 27°\text{C} = 300\text{K} \end{array}\right]$$
$$= 5.543 \times 10^4 \text{ Pa}$$

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,

$p_1 = 0.8$ bar $\qquad p_2 = p_{H_2}$

$V_1 = 0.5$ L $\qquad V_2 = 1$ L $\quad = ?$

It is known that,

$$p_1 V_1 = p_2 V_2$$
$$\Rightarrow p_2 = \frac{p_1 V_1}{V_2}$$
$$\Rightarrow p_{H_2} = \frac{0.8 \times 0.5}{1}$$
$$= 0.4 \text{ bar}$$

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

Now,

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

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

$$p_1 V_1 = p_2 V_2$$

$$\Rightarrow p_2 = \frac{p_1 V_1}{V_2}$$

$$\Rightarrow p_{O_2} = \frac{0.7 \times 20}{1}$$

$$= 0.4 \text{ bar}$$

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

$$p_{total} = p_{H_2} + p_{O_2}$$

$$= 0.4 + 1.4$$

$$= 1.8 \text{ bar}$$

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,

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

$p_1 = 2 \text{ bar}$

$T_1 = 27°\text{C} = (27 + 273) \text{ K} = 300 \text{ K}$

$p_2 = 1 \text{ bar}$

$T_2 = 273 \text{ K}$

$d_2 = ?$

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

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

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

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

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

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

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

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 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:

$$pV = nRT$$

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

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

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

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}.

Question 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?

Answer

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,

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

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

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

$$= 2.5 \text{ V}$$

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

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

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 $O_2 = \frac{8}{32} = 0.25$ mole

Mass of dihydrogen (H_2) = 4 g

Thus, number of moles of $H_2 = \frac{4}{2} = 2$ 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*R*T*

$$\Rightarrow p = \frac{nRT}{V}$$
$$= \frac{225 \times 0.083 \times 300}{1}$$
$$= 56.025 \text{ bar}$$

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

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,

$$m = \frac{MpV}{\text{R}T}$$

Here,

$$M = 4 \times 10^{-3}\,\text{kg mol}^{-1}$$
$$p = 1.66\,\text{bar}$$
$$V = \text{Volume of the balloon}$$
$$= 4190.5\ \text{m}^3$$
$$\text{R} = 0.083\,\text{bar dm}^3\,K^{-1}\text{mol}^{-1}$$
$$T = 27°\text{C} = 300\text{K}$$

$$\text{Then, } m = \frac{4 \times 10^{-3} \times 1.66 \times 4190.5 \times 10^3}{0.083 \times 300}$$
$$= 1117.5\,\text{kg}\,(\text{approx})$$

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,

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

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

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,

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

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

Question 5.19:

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

Answer

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

Then, the number of moles of dihydrogen, $n_{H_2} = \frac{20}{2} = 10 \text{ moles}$ and the number of moles of dioxygen, $n_{O_2} = \frac{80}{32} = 2.5 \text{ moles}$.

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.

Question 5.20:

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

Answer

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,

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

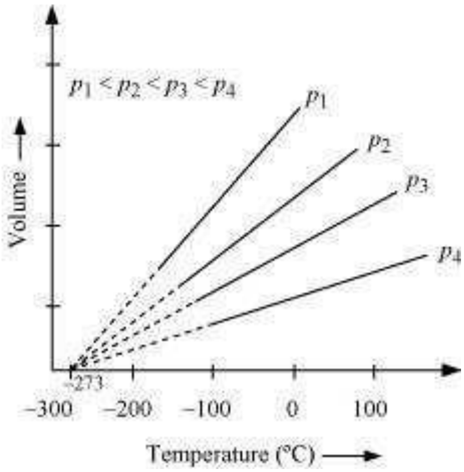
$$= Nm^4K^2\,mol^{-1}$$

Question 5.21:

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

Answer

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.

Question 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?

Answer

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.

Question 5.23:

Explain the physical significance of Van der Waals parameters.

Answer

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.