

## NCERT Solutions for Class 11 Chemistry Chapter 5

### States of Matter Class 11

#### Chapter 5 States of Matter Exercise Solutions

Exercise : Solutions of Questions on Page Number : 152

**Q1 :**

**What will be the minimum pressure required to compress 500 dm<sup>3</sup> of air at 1 bar to 200 dm<sup>3</sup> at 30°C?**

**Answer :**

Given,

Initial pressure,  $p_1 = 1$  bar

Initial volume,  $V_1 = 500$  dm<sup>3</sup>

Final volume,  $V_2 = 200$  dm<sup>3</sup>

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

$$\begin{aligned} p_1 V_1 &= p_2 V_2 \\ \Rightarrow p_2 &= \frac{p_1 V_1}{V_2} \\ &= \frac{1 \times 500}{200} \text{ bar} \\ &= 2.5 \text{ bar} \end{aligned}$$

Therefore, the minimum pressure required is 2.5 bar.

**Q2 :**

**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,

$$\begin{aligned}
 p_1 V_1 &= p_2 V_2 \\
 p_2 &= \frac{p_1 V_1}{V_2} \\
 &= \frac{1.2 \times 120}{180} \text{ bar} \\
 &= 0.8 \text{ bar}
 \end{aligned}$$

Therefore, the pressure would be 0.8 bar.

**Q3 :**

**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,

$$pV = nRT \dots\dots\dots (i) \text{ Where, } p$$

' Pressure of gas

' 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

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

Where,

$m$  ' Mass of gas

$M$  ' Molar mass of gas

$$\frac{m}{V} = d$$

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  $(T)$ ,  $\frac{M}{RT} = \text{constant}$ .

$$d = (\text{constant}) p$$
$$\Rightarrow d \propto p$$

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

**Q4 :**

**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,

$$p_1 = 2 \text{ bar}$$

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

Molecular mass of nitrogen,  $M_2 = 28 \text{ g/mol}$

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

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

**Q5 :**

**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,

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

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,

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

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

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

From equation (ii), we have

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

Where,  $M_A$  and  $M_B$  are the molecular masses of gases A and B respectively. Now, from equations (iii) and (iv), we have

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

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

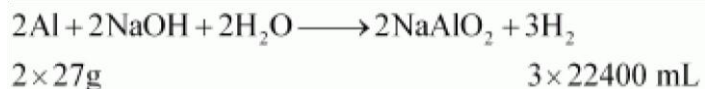
$$4M_A = M_B$$

**Q6 :**

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:



At STP (273.15 K and 1 atm), 54 g ( $2 \times 27$  g) of Al gives  $3 \times 22400$  mL of  $\text{H}_2$ .

$$\therefore 0.15 \text{ g Al gives } \frac{3 \times 22400 \times 0.15}{54} \text{ mL of H}_2 \text{ i.e., } 186.67 \text{ mL of H}_2$$

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  $T_2 = 20^\circ\text{C} = (273.15 + 20) \text{ K} = 293.15 \text{ K}$ .

Now,

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

$$\Rightarrow V_2 = \frac{p_1 V_1 T_2}{p_2 T_1}$$

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

$$= 202.98 \text{ mL}$$

$$= 203 \text{ mL}$$

Therefore, 203 mL of dihydrogen will be released.

**Q7 :**

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  $9 \text{ dm}^3$  flask at  $27^\circ\text{C}$  ?

**Answer :**

It is known that,

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

For methane ( $\text{CH}_4$ ),

$$p_{\text{CH}_4} = \frac{3.2}{16} \times \frac{8.314 \times 300}{9 \times 10^{-3}} \left[ \begin{array}{l} \text{Since } 9 \text{ dm}^3 = 9 \times 10^{-3} \text{ m}^3 \\ 27^\circ\text{C} = 300\text{K} \end{array} \right]$$

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

For carbon dioxide ( $\text{CO}_2$ ),

$$p_{\text{CO}_2} = \frac{4.4}{44} \times \frac{8.314 \times 300}{9 \times 10^{-3}}$$

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

Total pressure exerted by the mixture can be obtained as:

$$p = p_{\text{CH}_4} + p_{\text{CO}_2}$$

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

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

Hence, the total pressure exerted by the mixture is  $8.314 \times 10^4 \text{ Pa}$ .

**Q8 :**

What will be the pressure of the gaseous mixture when 0.5 L of  $\text{H}_2$  at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1L vessel at  $27^\circ\text{C}$ ?

**Answer :**

Let the partial pressure of  $\text{H}_2$  in the vessel be  $p_{\text{H}_2}$ .

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,

$$p_1 V_1 = p_2 V_2$$

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

$$\begin{aligned} \Rightarrow p_{H_2} &= \frac{0.8 \times 0.5}{1} \\ &= 0.4 \text{ bar} \end{aligned}$$

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

Now,

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

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

$$p_1 V_1 = p_2 V_2$$

**Q9 :**

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

**Answer :**

Given,

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

$$p_1 = 2 \text{ bar}$$

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

$$p_2 = 1 \text{ bar}$$

$$T_2 = 273 \text{ K}$$

$$d_2 = ?$$

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



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

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

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

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

Hence, the density of the gas at STP will be 3 g dm<sup>3</sup>.

**Q10 :**

**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

$V = 34.05 \text{ mL} = 34.05 \times 10^{-3} \text{ L} = 34.05 \times 10^{-3} \text{ dm}^3$

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

$T = 546^\circ\text{C} = (546 + 273) \text{ K} = 819 \text{ K}$

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

$$pV = nRT$$

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

$$\begin{aligned} &= \frac{0.1 \times 34.05 \times 10^{-3}}{0.083 \times 819} \\ &= 5.01 \times 10^{-5} \text{ mol} \end{aligned}$$

$$= \frac{0.0625}{5.01 \times 10^{-5}}$$

Therefore, molar mass of phosphorus =  $1247.5 \text{ g mol}^{-1}$

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

**Q11 :**

**A student forgot to add the reaction mixture to the round bottomed flask at  $27^\circ \text{C}$  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  $477^\circ \text{C}$ . 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  $27^\circ \text{C}$  is  $V$ .

Now,

$$V_1 = V$$

$$T_1 = 27^\circ \text{C} = 300 \text{ K}$$

$$V_2 = ?$$

$$T_2 = 477^\circ \text{C} = 750 \text{ K}$$

According to Charles's law,

$$\begin{aligned} \frac{V_1}{T_1} &= \frac{V_2}{T_2} \\ \Rightarrow V_2 &= \frac{V_1 T_2}{T_1} \\ &= \frac{750V}{300} \\ &= 2.5 V \end{aligned}$$

Therefore, volume of air expelled out =  $2.5 V - V = 1.5 V$

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

**Q12 :**

**Calculate the temperature of 4.0 mol of a gas occupying  $5 \text{ dm}^3$  at 3.32 bar.**

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

**Answer :**

Given,  $n =$

4.0 mol  $V =$

5 dm<sup>3</sup>  $p =$

3.32 bar

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

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

$$\begin{aligned} pV &= nRT \\ \Rightarrow T &= \frac{pV}{nR} \\ &= \frac{3.32 \times 5}{4 \times 0.083} \\ &= 50 \text{ K} \end{aligned}$$

Hence, the required temperature is 50 K.

**Q13 :**

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

**Answer :**

Molar mass of dinitrogen ( $\text{N}_2$ ) = 28 g mol<sup>-1</sup>

$$\text{Thus, } 1.4 \text{ g of } \text{N}_2 = \frac{1.4}{28} = 0.05 \text{ mol}$$

$$\begin{aligned} &= 0.05 \times 6.02 \times 10^{23} \text{ number of molecules} \\ &= 3.01 \times 10^{23} \text{ number of molecules} \end{aligned}$$

Now, 1 molecule of  $\text{N}_2$  contains 14 electrons.

Therefore,  $3.01 \times 10^{23}$  molecules of  $\text{N}_2$  contains =  $1.4 \times 3.01 \times 10^{23}$

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

**Q14 :**

**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 =  $6.02 \times 10^{23}$

Thus, time required

$$\begin{aligned}
 &= \frac{6.02 \times 10^{23}}{10^{10}} \text{ s} \\
 &= 6.02 \times 10^{23} \text{ s} \\
 &= \frac{6.02 \times 10^{23}}{60 \times 60 \times 24 \times 365} \text{ years} \\
 &= 1.909 \times 10^6 \text{ years}
 \end{aligned}$$

Hence, the time taken would be  $1.909 \times 10^6$  years .

**Q15 :**

**Calculate the total pressure in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of 1 dm<sup>3</sup> at 27°C. R = 0.083 bar dm<sup>3</sup> K<sup>-1</sup> mol<sup>-1</sup>.**

**Answer :**

Given,

Mass of dioxygen (O<sub>2</sub>) = 8 g

Thus, number of moles of

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

Mass of dihydrogen (H<sub>2</sub>) = 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<sup>3</sup> n = 2.25

mol

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

$$T = 27^\circ\text{C} = 300 \text{ K}$$

Total pressure ( $p$ ) can be calculated as:  $pV$

$$= nRT$$

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

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

**Q16 :**

**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  $27^\circ\text{C}$ . (Density of air =  $1.2 \text{ kg m}^{-3}$  and  $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$ ).**

**Answer :**

Given,

Radius of the balloon,  $r = 10 \text{ m}$

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

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

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

Thus, the volume of the displaced air is  $4190.5 \text{ m}^3$ .

Given,

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

Then, mass of displaced air =  $4190.5 \times 1.2 \text{ kg}$

$$= 5028.6 \text{ kg}$$

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

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

Here,

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

$$p = 1.66 \text{ bar}$$

$V$  = Volume of the balloon

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

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

$$T = 27^\circ\text{C} = 300\text{K}$$

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

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

$$= 1217.5 \text{ kg}$$

Hence, pay load = (5028.6 – 1217.5) kg

$$= 3811.1 \text{ kg}$$

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

**Q17 :**

**Calculate the volume occupied by 8.8 g of CO<sub>2</sub> at 31.1°C and 1 bar pressure.**

**R = 0.083 bar L K<sup>-1</sup> mol<sup>-1</sup>.**

**Answer :**

It is known that,

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

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

Here,  $m$

$$= 8.8 \text{ g}$$

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

$$T = 31.1^\circ\text{C} = 304.1 \text{ K}$$

$$M = 44 \text{ g } p = 1 \text{ bar}$$

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

Hence, the volume occupied is 5.05 L.

**Q18 :**

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,

$$\begin{aligned}V &= \frac{m RT}{M p} \\ &= \frac{0.184}{2} \times \frac{R \times 290}{p}\end{aligned}$$

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

$$\begin{aligned}V &= \frac{m RT}{M p} \\ &= \frac{2.9}{M} \times \frac{R \times 368}{p}\end{aligned}$$

According to the question,

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

Hence, the molar mass of the gas is 40 g mol<sup>-1</sup>.

**Q19 :**

**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,  $n_{\text{H}_2} = \frac{20}{2} = 10$  moles and the number of moles of

dioxygen,  $n_{\text{O}_2} = \frac{80}{32} = 2.5$  moles.

Given,

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

Then, partial pressure of dihydrogen,

$$\begin{aligned} p_{\text{H}_2} &= \frac{n_{\text{H}_2}}{n_{\text{H}_2} + n_{\text{O}_2}} \times P_{\text{total}} \\ &= \frac{10}{10 + 2.5} \times 1 \\ &= 0.8 \text{ bar} \end{aligned}$$

Hence, the partial pressure of dihydrogen is **0.8 bar**.

**Q20 :**

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

**Answer :**

The SI unit for pressure,  $p$  is  $\text{Nm}^{-2}$ .

The SI unit for volume,  $V$  is  $\text{m}^3$ .

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



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

$$\frac{pV^2T^2}{n}$$

Therefore, the SI unit for quantity  $n$  is given by,

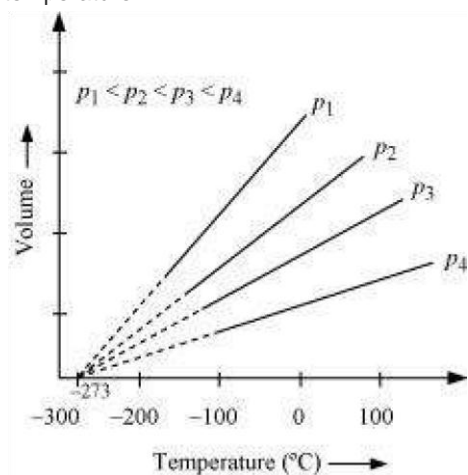
$$\begin{aligned} &= \frac{(\text{Nm}^{-2})(\text{m}^3)^2(\text{K})^2}{\text{mol}} \\ &= \text{Nm}^4\text{K}^2\text{mol}^{-1} \end{aligned}$$

**Q21 :**

**In terms of Charles' law explain why  $-273^\circ\text{C}$  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  $^\circ\text{C}$ ) is a straight line. If this line is extended to zero volume, then it intersects the temperature-axis at  $-273^\circ\text{C}$ . In other words, the volume of any gas at  $-273^\circ\text{C}$  is zero. This is because all gases get liquefied before reaching a temperature of  $-273^\circ\text{C}$ . Hence, it can be concluded that  $-273^\circ\text{C}$  is the lowest possible temperature.

**Q22 :**

**Critical temperature for carbon dioxide and methane are  $31.1^\circ\text{C}$  and  $-81.9^\circ\text{C}$  respectively. Which of these has stronger intermolecular forces and why?**



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**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  $\text{CO}_2$ .

**Q23 :**

**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.