

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 dm³ of air at 1 bar to 200 dm³ at 30°C?

Answer :

Given,

Initial pressure, $p_1 = 1$ bar

Initial volume, $V_1 = 500$ dm³

Final volume, $V_2 = 200$ dm³

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,

$$pV = nRT \dots\dots\dots (i)$$

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

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

Where,

m ' Mass of gas

M ' Molar mass of gas

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

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

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

Given,

$$\begin{aligned}
 m_A &= 1 \text{ g} \\
 p_A &= 2 \text{ bar} \\
 m_B &= 2 \text{ g} \\
 p_B &= (3 - 2) = 1 \text{ bar}
 \end{aligned}$$

(Since total pressure is 3 bar)

Substituting these values in equation (v), we have

$$\begin{aligned}
 \frac{2 \times M_A}{1} &= \frac{1 \times M_B}{2} \\
 \Rightarrow 4M_A &= M_B
 \end{aligned}$$

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

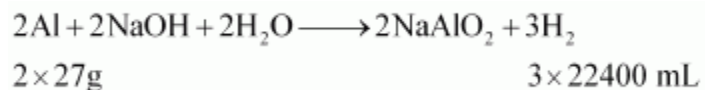
$$4M_A = M_B$$

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 × 27 g) of Al gives 3 × 22400 mL of H₂.

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

At STP,

$$\begin{aligned}
 p_1 &= 1 \text{ atm} \\
 V_1 &= 186.67 \text{ mL} \\
 T_1 &= 273.15 \text{ K}
 \end{aligned}$$

Let the volume of dihydrogen be V_2 at $p_2 = 0.987 \text{ atm}$ (since 1 bar = 0.987 atm) and $T_2 = 20^\circ\text{C} = (273.15 + 20) \text{ K} = 293.15 \text{ K}$.

Now,

$$\begin{aligned}\frac{p_1 V_1}{T_1} &= \frac{p_2 V_2}{T_2} \\ \Rightarrow V_2 &= \frac{p_1 V_1 T_2}{p_2 T_1} \\ &= \frac{1 \times 186.67 \times 293.15}{0.987 \times 273.15} \\ &= 202.98 \text{ mL} \\ &= 203 \text{ mL}\end{aligned}$$

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 dm³ flask at 27 °C ?

Answer :

It is known that,

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

For methane (CH₄),

$$\begin{aligned}p_{\text{CH}_4} &= \frac{3.2}{16} \times \frac{8.314 \times 300}{9 \times 10^{-3}} \left[\begin{array}{l} \text{Since } 9 \text{ dm}^3 = 9 \times 10^{-3} \text{ m}^3 \\ 27^\circ\text{C} = 300\text{K} \end{array} \right] \\ &= 5.543 \times 10^4 \text{ Pa}\end{aligned}$$

For carbon dioxide (CO₂),

$$\begin{aligned}p_{\text{CO}_2} &= \frac{4.4}{44} \times \frac{8.314 \times 300}{9 \times 10^{-3}} \\ &= 2.771 \times 10^4 \text{ Pa}\end{aligned}$$

Total pressure exerted by the mixture can be obtained as:

$$\begin{aligned}p &= p_{\text{CH}_4} + p_{\text{CO}_2} \\ &= (5.543 \times 10^4 + 2.771 \times 10^4) \text{ Pa} \\ &= 8.314 \times 10^4 \text{ Pa}\end{aligned}$$

Hence, the total pressure exerted by the mixture is 8.314 × 10⁴ Pa.

Q8 :

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 $27^\circ C$?

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,

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

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 g/dm^3 at $27^\circ C$ at 2 bar pressure. What will be its density at STP?

Answer :

Given,

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

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

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

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

$$T_2 = 273 K$$

$$d_2 = ?$$

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} .

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,

$$p = 0.1 \text{ bar}$$

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

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

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

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

$$pV = nRT$$

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

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

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

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

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°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,

$$V_1 = V$$

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

$$V_2 = ?$$

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

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 = $2.5 V - V = 1.5 V$

$$= \frac{1.5V}{2.5V} = \frac{3}{5}$$

Hence, fraction of air expelled out

Q12 :

Calculate the temperature of 4.0 mol of a gas occupying 5 dm^3 at 3.32 bar .

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

Answer :

Given,

$$n = 4.0 \text{ mol}$$

$$V = 5 \text{ dm}^3$$

$$p = 3.32 \text{ bar}$$

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

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 (N_2) = 28 g mol^{-1}

$$\begin{aligned} \text{Thus, } 1.4 \text{ g of } \text{N}_2 &= \frac{1.4}{28} = 0.05 \text{ mol} \\ &= 0.05 \times 6.02 \times 10^{23} \text{ number of molecules} \\ &= 3.01 \times 10^{23} \text{ number of molecules} \end{aligned}$$

Now, 1 molecule of N_2 contains 14 electrons.

$$\begin{aligned} \text{Therefore, } 3.01 \times 10^{23} \text{ molecules of } \text{N}_2 &\text{ contains } = 1.4 \times 3.01 \times 10^{23} \\ &= 4.214 \times 10^{23} \text{ electrons} \end{aligned}$$

Q14 :

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

Answer :

$$\text{Avogadro number} = 6.02 \times 10^{23}$$

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 1.909×10^6 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 dm³ at 27°C. R = 0.083 bar dm³ K⁻¹ mol⁻¹.

Answer :

Given,

Mass of dioxygen (O₂) = 8 g

$$\text{O}_2 = \frac{8}{32} = 0.25 \text{ mole}$$

Thus, number of moles of

Mass of dihydrogen (H₂) = 4 g

$$\text{H}_2 = \frac{4}{2} = 2 \text{ mole}$$

Thus, number of moles of

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

Given,

$$V = 1 \text{ dm}^3$$

$$n = 2.25 \text{ mol}$$

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

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

Total pressure (p) can be calculated as:

$$pV = nRT$$

$$\begin{aligned}
 \Rightarrow p &= \frac{nRT}{V} \\
 &= \frac{2.25 \times 0.083 \times 300}{1} \\
 &= 56.025 \text{ bar}
 \end{aligned}$$

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°C.

(Density of air = 1.2 kg m⁻³ and R = 0.083 bar dm³ K⁻¹ mol⁻¹).

Answer :

Given,

Radius of the balloon, $r = 10$ m

$$\begin{aligned}\therefore \text{Volume of the balloon} &= \frac{4}{3}\pi r^3 \\ &= \frac{4}{3} \times \frac{22}{7} \times 10^3 \\ &= 4190.5 \text{ m}^3 \text{ (approx)}\end{aligned}$$

Thus, the volume of the displaced air is 4190.5 m³.

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,

$$\begin{aligned}m &= \frac{MpV}{RT} \\ \text{Here,} \\ M &= 4 \times 10^{-3} \text{ kg mol}^{-1} \\ p &= 1.66 \text{ bar} \\ V &= \text{Volume of the balloon} \\ &= 4190.5 \text{ m}^3 \\ R &= 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1} \\ T &= 27^\circ\text{C} = 300\text{K} \\ \text{Then, } m &= \frac{4 \times 10^{-3} \times 1.66 \times 4190.5 \times 10^3}{0.083 \times 300} \\ &= 1117.5 \text{ kg (approx)}\end{aligned}$$

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.

Q17 :

Calculate the volume occupied by 8.8 g of CO₂ 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 L K⁻¹ mol⁻¹

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

$M = 44$ g

$p = 1$ 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,

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

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

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} .

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, $n_{\text{H}_2} = \frac{20}{2} = 10 \text{ moles}$ and the number of moles of dioxygen, $n_{\text{O}_2} = \frac{80}{32} = 2.5 \text{ moles}$.

Given,

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

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 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,

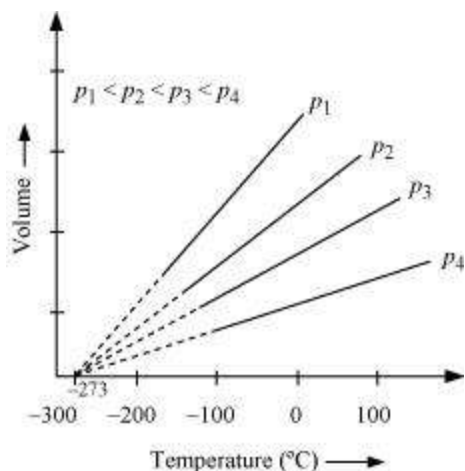
$$\begin{aligned}
 &= \frac{(\text{Nm}^{-2})(\text{m}^3)^2(\text{K})^2}{\text{mol}} \\
 &= \text{Nm}^4\text{K}^2\text{mol}^{-1}
 \end{aligned}$$

Q21 :

In terms of Charles' law explain why -273°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.

Q22 :

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 CO₂.

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.