SpeedLabs

Science

CBSE 11th

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Q.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 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.

Q.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 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,

$$p_1V_1 = p_2V_2$$

$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$

$$= \frac{1.2 \times 120}{180}\text{bar}$$

$$= 0.8 \text{ bar}$$

Q.3 Using the equation of state pV = nRT; show that at a given temperature density of a gas is proportional to gas pressurep.

Ans 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

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

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

$\Rightarrow d \propto p$

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

Q.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 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

$p_1 = 1.2$ bar

$p_2 = 120$ mL

Molecular mass of nitrogen, $M_2 = 28$ g/mol

$$M_1 = \frac{M_2 p_2}{p_1}$$

$$M_1 = \frac{28 \times 5}{2}$$

$$= 70 \text{ g/mol.}$$

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

Q.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 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

$m_A = 1g$

$p_A = 2$ bar

$m_B = 2g$

$p_B = (3 - 2) = 1$ 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

$$4M_A = M_B$$

Q.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 The reaction of aluminium with caustic soda can be represented as:

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

2×27 g $\qquad\qquad\qquad\qquad$ 3×22400 mL

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 $\frac{3 \times 22400 \times 0.15}{54}$ mL of H_2 i.e., 186.67 mL of H2.

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,

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

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

$$V_2 = \frac{1 \times 186.67 \times 293.15}{0.987 \times 273.15}$$

$$V_2 = 202.98mL$$

$$V_2 = 203mL$$

Therefore, 203 mL of dihydrogen will be released.

Q.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 It is known that,

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

For methane (CH_4),

$$p_{CH_4}=\frac{3.2}{16}\times\frac{8.314\times300}{9\times10^{-3}}\begin{bmatrix}\text{Since } 9\,dm^3=9\times10^{-3}m^3\\ 27°C=300K\end{bmatrix}$$

$= 5.543 \times 10^4$ Pa

For carbon dioxide (CO_2),

$$p_{CO_2}=\frac{4.4}{44}\times\frac{8.314\times300}{9\times10^{-3}}$$

$= 2.771 \times 10^4$ Pa

Total pressure exerted by the mixture can be obtained as:

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

$$p=\left(5.543\times10^4+2.771\times10^4\right)Pa$$

$$p=8.314\times10^4\,Pa$$

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

Q.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?

Ans 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

It is known that,

$$p_1V_1=p_2V_2$$

$$\Rightarrow p_2=\frac{p_1V_1}{V_2}$$

$$\Rightarrow p_{H_2}=\frac{0.8\times0.5}{1}$$

$= 0.4$ bar

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

Now,

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

$V_1 = 2.0$ L $\quad\quad$ $V_2 = 1$ L

$$p_1V_1 = p_2V_2$$

$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$

$$p_{O_2} = \frac{0.7 \times 2.0}{1}$$

$$= 1.4 \text{ bar}$$

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

Q.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 Given,

$$d_1 = 5.46 g/dm^3$$

$$p_1 = 2 bar$$

$$T_1 = 27°C = (27 + 273)K = 300K$$

$$p_2 = 1 bar$$

$$T_2 = 273K$$

$$d_2 = ?$$

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

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

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

$$\frac{d_1}{d_2} = \frac{p_1T_1}{p_2T_2}$$

$$d_2 = \frac{p_2T_1d_1}{p_1T_2}$$

$$d_2 = \frac{1\times300\times5.46}{2\times273}$$

$$d_2 = 3\text{gdm}^{-3}$$

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

Q.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 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:

$$pV = nRT$$

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

$$n = \frac{0.1\times34.05\times10^{-3}}{0.083\times819}$$

$$= 5.01\times10^{-5}\text{mol}$$

$$\text{Therefore, molar mass of phosphorus } = \frac{0.06250}{5.01\times10^{-5}} = 1247.5\text{gmol}^{-1}$$

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

Q.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 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,

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

$$V_2=\frac{V_1T_2}{T_1}$$

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

$$V_2=2.5V$$

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

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

Q.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 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:

$$pV=nRT$$

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

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

$$T=50K$$

Hence, the required temperature is 50 K.

Q.13 Calculate the total number of electrons present in 1.4 g of dinitrogen gas.

Ans 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

= 3.01×10^{23} number of molecules

Now, 1 molecule of N_2 contains 14 electrons.

Therefore, 3.01×10^{23} molecules of N_2 contains = $1.4 \times 3.01 \times 10^{23}$

= 4.214×10^{23} electrons

Q.14 How much time would it take to distribute one Avogadro number of wheat grains, if 10^{10} grains are distributed each second?

Ans Avogadro number = 6.02×10^{23}

Thus, time required

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

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

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

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

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

Q.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 Given,

Mass of dioxygen (O_2) = 8 g

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

Mass of dihydrogen (H_2) = 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,

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

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

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

$$p = \frac{225 \times 0.083 \times 300}{1} = 56.025$$

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

Q.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 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 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 = 4×10^{-3} kg mol^{-1}

p = 1.66 bar

V = Volume of the bottom

= 4190.5 m^3

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

T = 27°C = 300K

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

$= 1117.5$ 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.

Q.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 It is known that,

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

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

$$\text{Thus, volume (V) } = \frac{8.8 \times 0.083 \times 304.1}{44 \times 1}$$

$= 5.04806L$

$= 5.05$ L

Hence, the volume occupied is 5.05 L.

Q.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 Volume (V) occupied by dihydrogen is given by,

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

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

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

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

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

According to the question,

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

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

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

$$\Rightarrow M = 40 \text{gmol}^{-1}$$

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

Q.19 A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

Ans 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 bar

Hence, the partial pressure of dihydrogen is 0.8 bar.

Q.20 What would be the SI unit for the quantity pV^2T^2/n?

Ans 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.

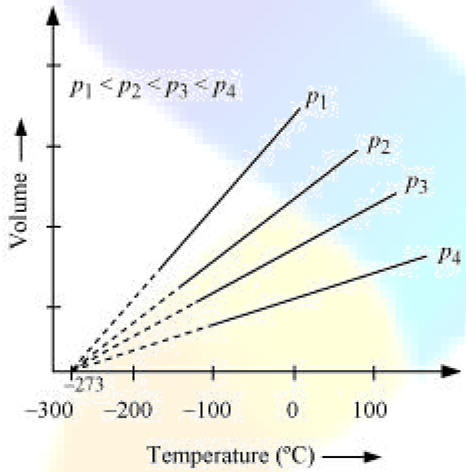
The SI unit for the number of moles, n is mol.

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

$$= \frac{\left(Nm^{-2}\right)\left(m^3\right)\left(K\right)^2}{mol} = Nm^4K^2mol^{-1}$$

Q.21 In terms of Charles' law explain why –273°C is the lowest possible temperature.

Ans 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.

Q.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 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.

Q.23 Explain the physical significance of Van der Waals parameters.

Ans 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.