NCERT Solutions for Class 11 Chemistry : Chapter 5 (States of Matter)

Q1. 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:

Initial pressure, P_1 = 1 bar

Initial volume, V_1 = 500 dm^3

Final volume, V_2 = 200 dm^3

As the temperature remains same,the final pressure (P_2) can be calculated with the help of Boyle's law.

Acc. Boyle's law,

$P_1V_1 = P_2V_2$

$P_2 = \frac{P_1V_1}{V_2}$

$= \frac{1 \times 500}{200}$

= 2.5 bar

∴ the *minimum pressure required to compress 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:

Initial pressure, P_1 = 1.2 bar

Initial volume, V_1 = 120 mL

Final volume, V_2 = 180 mL

As the temperature remains same, final pressure (P_2) can be calculated with the help of Boyle's law.

According to the Boyle's law,

$P_1V_1 = P_2V_2$

$P_2 = \frac{P_1V_1}{V_2}$

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

= 0.8 bar

Therefore, the min pressure required is 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 ....(1)

Where, p = pressure

V = volume

N = number of moles

R = Gas constant

T = temp

$\frac{n}{V} = \frac{p}{RT}$

Replace n with $\frac{m}{M}$, therefore,

$\frac{m}{MV} = \frac{p}{RT}$ ....(2)

Where, m = mass

M = molar mass

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

Where, d = density

Therefore, from equation (2), we get

$\frac{d}{M} = \frac{p}{RT}$

d = ($\frac{M}{RT}$) p

d $\propto$ p

Therefore, at a given temp, the density of gas (d) 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 temp (T) can be given by,

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

Now, density of oxide (d_1) is as given,

$d_1 = \frac{M_1p_1}{RT}$

Where, M_1 = mass of the oxide

p_1 = pressure of the oxide

Density of dinitrogen gas (d_2) is as given,

$d_2 = \frac{M_1p_2}{RT}$

Where, M_2 = mass of the oxide

p_2 = pressure of the oxide

Acc to the question,

$d_1 = d_2$

Therefore, $M_1p_1 = M_2p_2$

Given:

p_1 = 2 bar

p_2 = 5 bar

Molecular mass of nitrogen, M_2 = 28 g/mol

Now, M_1

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

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

= 70 g/mol

Therefore, 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,

$p_X V = n_X RT$ ......(1)

Where p_X and n_X represent the pressure and number of moles of gas X.

For ideal gas Y, the ideal gas equation is given by,

$p_Y V = n_Y RT$ ......(2)

Where, p_Y and n_Y represent the pressure and number of moles of gas Y.

[V and T are constants for gases X and Y]

From equation (1),

$p_X V = \frac{m_X}{M_X}$ RT

$\frac{p_X M_X}{m_X}$ = $\frac{RT}{V}$ ......(3)

From equation (2),

$p_Y V = \frac{m_Y}{M_Y}$ RT

$\frac{p_Y M_Y}{m_Y}$ = $\frac{RT}{V}$ ...... (4)

Where, M_X and M_Y are the molecular masses of gases X and Y respectively.

Now, from equation (3) and (4),

$\frac{p_X M_X}{m_X} = \frac{p_Y M_Y}{m_Y}$ ..... (5)

Given,

m_X = 1 g

p_X = 2 bar

m_Y = 2 g

p_Y = (3 − 2) = 1 bar (Since total pressure is 3 bar)

Substituting these values in equation (5),

$\frac{2 \times M_X}{1} = \frac{1 \times M_Y}{2}$

4 $M_X = M_Y$

Therefore, the relationship between the molecular masses of X and Y is,

4 $M_X = M_Y$

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 aluminum with caustic soda is as given below:

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

At Standard Temperature Pressure (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

= 186.67 mL of H_2

At Standard Temperature Pressure,

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

Now,

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

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

= 202.98 mL

= 203 mL

Hence, 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 (CH_4),

p_{CH_4}

= $\frac{3.2}{16}$ × $\frac{8.314 \times 300}{9 \times 10^{-3}}$ [Since 9 dm^3 = 9×10^{-3} m^3 $27°$ C = 300 K]

= 5.543 × 10^4 Pa

For carbon dioxide (CO_2),

p_{CO_2}

= $\frac{4.4}{44}$ × $\frac{8.314 \times 300}{9 \times 10^{-3}}$

= 2.771 × 10^4 Pa

Total pressure exerted by the mixture can be calculated as:

p = p_{CH_4} + p_{CO_2}

= (5.543 × 10^4 + 2.771 × 10^4) Pa

= 8.314 × 10^4 Pa

Q8. What will be the pressure of the gaseous mixture when 0.5 L of H2 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 container be p_{H_2}.

Now,

p_1 = 0.8 bar

p_2 = p_{H_2} V_1 = 0.5 L

V_2 = 1 L

It is known that,

$p_1 \; V_1 = p_2 \; V_2 \; p_2 = \frac{p_1 \times V_1}{V_2} \; p_{H_2} = \frac{0.8 \times 0.5}{1}$

= 0.4 bar

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

Now,

p_1 = 0.7 bar

$p_2 = p_{O_2} \; V_1$ = 2.0 L

V_2 = 1 L

$p_1 \; V_1 = p_2 \; V_2 \; p_2 = \frac{p_1 \times V_1}{V_2} \; p_{O_2} = \frac{0.7 \times 20}{1}$

= 1.4 bar

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

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

= 0.4 + 1.4

= 1.8 bar

Q9. Density of a gas is found to be 5.46 g/dm3 at 27 °C at 2 bar pressure. What will be its density at STP?

Answer:

Given,

d_1 = 5.46 g/dm^3

p_1 = 2 bar

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

p_2 = 1 bar

T_2 = 273 K

d_2 = ?

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

d = $\frac{Mp}{RT} \; \frac{d_1}{d_2} = \frac{\frac{M \, p_1}{R \, T_1}}{\frac{M \, p_2}{R \, T_2}} \; \frac{d_1}{d_2} = \frac{p_1 \; T_2}{p_2 \; T_1}$

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

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

= 3 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 bar

V = 34.05 mL = 34.05 × 10^{-3} dm^3

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

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

The no of moles (n) can be calculated using the ideal gas equation as:

pV = nRT

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

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

= 5.01 × 10^{-5} mol

Therefore, molar mass of phosphorus = $\frac{0.0625}{5.01 \times 10^{-5}}$

= 1247.5 g mol^{-1}

Q11. A student forgot to add the reaction mixture to the container at 27° C but instead, he placed the container on the flame. After a lapse of time, he came to know about his mistake, and using a pyrometer he found the temp of the container 477° C. What fraction of air would have been expelled out?

Answer:

Let the volume of the container be V.

The volume of the air inside the container at 27° C is V.

Now,

V_1 = V

T_1 = 27° C = 300 K V_2 = ?

T_2 = 477° C = 750 K

Acc to Charles's law,

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

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

= 2.5 V

Therefore, volume of air expelled out

= 2.5 V − V = 1.5 V

Hence, fraction of air expelled out

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

= $\frac{3}{5}$

Q12. 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}).

Given,

N= 4.0 mol

V = 5 dm^3

p = 3.32 bar

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

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

pV = nRT

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

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

= 50 K

Therefore, the required temp 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}

Thus, 1.4 g of N_2

= $\frac{1.4}{28}$

= 0.05 mol

= 0.05 × 6.02 × 10^{23} no of molecules

= 3.01 × 10^{23} no. of molecules

Now, 1 molecule of N_2 has 14 electrons.

Therefore, 3.01 × 10^{23} molecules of N_2 contains,

= 14 × 3.01 × 1023

= 4.214 × 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 no. = 6.02 × 10^{23}

Therefore, time taken

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

= 6.02 × 10^{13} s

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

= 1.909×10^6 years

Therefore, 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 dm^3 at 27°C. R = 0.083 bar dm^3 K^{-1} mol^{-1}

Answer:

Given:

Mass of O_2 = 8 g

No. of moles

= $\frac{8}{32}$

= 0.25 mole

Mass of H_2 = 4 g

No. of moles

= $\frac{4}{2}$

= 2 mole

Hence, total no 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 at K^{-1} mol^{-1}

T = 27° C = 300 K

Total pressure :

pV = nRT

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

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

= 56.025 bar

Therefore, 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 dm^3 K^{-1} mol^{-1})

Answer:

Given:

r = 10 m

Therefore, volume of the balloon

= $\frac{4}{3}$ πr^3

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

= 4190.5 m^3 (approx.)

Therefore, the volume of the displaced air

= 4190.5 × 1.2 kg

= 5028.6 kg

Mass of helium,

= $\frac{MpV}{RT}$

Where, M = 4 × 10^{-3} kg mol^{-1}

p = 1.66 bar

V = volume of the balloon

= 4190.5 m^3

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

T = 27 °C = 300 K

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 with helium,

= (100 + 1117.5) kg

= 1217.5 kg

Therefore, pay load,

= (5028.6 − 1217.5)

= 3811.1 kg

Therefore, 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 dm^3 K^{-1} mol^{-1}.

Answer:

pVM = mRT

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

Given:

m = 8.8 g

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

T = 31.1 °C = 304.1 K

M = 44 g

p = 1 bar

Thus, Volume (V),

= $\frac{8.8 \times 0.083 \times 304.1}{44 \times 1}$

= 5.04806 L

= 5.05 L

Therefore, 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 = $\frac{mRT}{Mp}$

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

Let M be the molar mass of the unknown gas.

Volume occupied by the unknown gas is,

= $\frac{mRT}{Mp}$

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

According to the ques,

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

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

= 40 g mol^{-1}

Therefore, 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.

Let the weight of dioxygen be 80 g.

No. of moles of dihydrogen (n_{H2}),

= $\frac{20}{2}$

= 10 moles

No. of moles of dioxygen (n_{O2}),

= $\frac{80}{32}$

= 2.5 moles

Given:

p_{total} = 1 bar

Therefore, partial pressure of dihydrogen (p_{H2}),

= $\frac{n_{H_2}}{n_{H_2} + n_{O_2}}$ × p_{total}

= $\frac{10}{10 + 2.5} \times 1$

= 0.8 bar

Therefore, the partial pressure of dihydrogen is 0.8 bar.

Q20. What will be the SI unit for the quantity $\frac{pV^2T^2}{n}$?

Answer:

SI unit of pressure, p = Nm^{-2}

SI unit of volume, V = m^3

SI unit of temp, T = K

SI unit of number of moles, n = mol

Hence, SI unit of $\frac{pV^2T^2}{n}$ is,

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

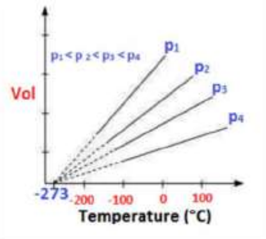
= $Nm^4K^2mol^{-1}$

Q21. In terms of Charles' law explain why -273° C is the lowest possible temp.

Answer:

According to Charles' law

At constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temp.



It was found that for all gasses (at any given pressure), the plot of volume vs. temp. (in $^\circ$ C) is a straight line.

If we extend the line to zero volume, then it intersects the temp-axis at -273° C. That is the volume of any gas at -273° C is 0. This happens because all gasses get transferred into liquid form before reaching -273° C.

Therefore, it can be said that -273° C is the lowest possible temp.

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:

If the critical temp of a gas is higher then it is easier to liquefy. That is the intermolecular forces of attraction among the molecules of gas are directly proportional to its critical temp.

Therefore, in CO_2 intermolecular forces of attraction are stronger.

Q23. Explain the physical significance of Van der Waals parameters?

Answer:

The physical significance of 'a':

The magnitude of intermolecular attractive forces within gas is represented by 'a'.

The physical significance of 'b':

The volume of a gas molecule is represented by 'b'.