States of Matter

Question1

Which of the following forces is involved in dinitrogen?

MHT CET 2025 5th May Evening Shift

Options:

A.

Dipole - dipole interaction

B.

Dipole - induced dipole interaction

C.

London dispersion force

D.

Hydrogen bonding

Answer: C

Solution:

Dispersion forces are also called as London forces or van der Waals forces. It is the weakest intermolecular force that develops due to interaction between two nonpolar molecules.

Question2

A container consists mixture of four gases as $5\ \mathrm{gH_2}$, $8\ \mathrm{gHe}$, $50\ \mathrm{gCO_2}$ and 20 g Ne at a certain temperature. Which of the following gases

exerts minimum partial pressure?

MHT CET 2025 26th April Evening Shift

Options:

A.

H_2

B.

He

C.

CO_2

D.

Ne

Answer: D

Solution:

$$n_{H_2} = \frac{5\ g}{2\ g\ mol^{-1}} = 2.5\ mol$$
$$n_{He} = \frac{8\ g}{4 gmol^{-1}} = 2\ mol$$
$$n_{CO_2} = \frac{50\ g}{44\ g\ mol^{-1}} = 1.143\ mol$$
$$n_{Ne} = \frac{20\ g}{20\ g\ mol^{-1}} = 1\ mol$$

Partial pressure of any gas $(A) = P_A$

$= x_A \times P_{Total} \quad = \frac{n_A}{n_{Total}} \times P_{Total}$

P_{Total} and n_{Total} is same, for all gases.

$\therefore \quad P_A \propto n_A.$

As Ne has the least number of moles, it exerts minimum partial pressure.

Question3

Equal masses of helium and oxygen are mixed in an empty container at $25^{\circ}\mathrm{C}$. What is the fraction of the total pressure exerted by helium?

MHT CET 2025 26th April Morning Shift

Options:

A. $\frac{1}{2}$

B. $\frac{1}{4}$

C. $\frac{8}{9}$

D. $\frac{7}{9}$

Answer: C

Solution:

Step 1: Recall Dalton's law of partial pressures

The partial pressure of a gas in a mixture is proportional to its mole fraction in the mixture.

$$\frac{P_{\mathrm{He}}}{P_{\mathrm{total}}} = \frac{n_{\mathrm{He}}}{n_{\mathrm{He}} + n_{\mathrm{O}_2}}$$

Step 2: Compute number of moles of each gas

Suppose each gas has **mass** m.

- For helium ($M_{\mathrm{He}} = 4\,\mathrm{g/mol}$):

$$n_{\mathrm{He}} = \frac{m}{4}$$

- For oxygen (O_2, $M_{\mathrm{O}_2} = 32\,\mathrm{g/mol}$):

$$n_{\mathrm{O}_2} = \frac{m}{32}$$

Step 3: Mole fraction of He

$$x_{\mathrm{He}} = \frac{n_{\mathrm{He}}}{n_{\mathrm{He}} + n_{\mathrm{O}_2}} = \frac{\frac{m}{4}}{\frac{m}{4} + \frac{m}{32}}$$

Simplify:

$$x_{\mathrm{He}} = \frac{1/4}{1/4+1/32} = \frac{1/4}{8/32+1/32} = \frac{1/4}{9/32} = \frac{32}{36} = \frac{8}{9}$$

Step 4: Final Answer

The fraction of the total pressure exerted by helium is:

$$\boxed{\frac{8}{9}}$$

Correct Option: C ($\frac{8}{9}$) ✅

Question4

A container contains $4\ \mathrm{g} \mathrm{H}_2$, $4\ \mathrm{g} \mathrm{He}$ and certain amount of 'Ne' at a certain temperature. What is the mass of 'Ne' required so that the partial pressure exerted by 'Ne' is equal to the partial pressure of He ?

MHT CET 2025 25th April Evening Shift

Options:

A. 4 g

B. 8 g

C. 10 g

D. 20 g

Answer: D

Solution:

Step 1: Relation between partial pressure and moles

Partial pressure of a gas:

$$P_i \propto n_i \quad \text{at fixed T and V}$$

Thus, having equal partial pressure of Ne and He means:

$$n_{\mathrm{Ne}} = n_{\mathrm{He}}$$

Step 2: Moles of He

Molar mass of He = 4 g/mol

Mass of He = 4 g

$n_{He} = \frac{4}{4} = 1$ mol

Step 3: Required moles of Ne

So we require:

$n_{Ne} = 1$ mol

Molar mass of Ne = 20 g/mol

$m_{Ne} = n_{Ne} \times M_{Ne} = 1 \times 20 = 20$ g

Step 4: Final Answer

$\boxed{20 \text{ g}}$

Correct option = **D (20 g)**

Question5

A container contains equal masses of H_2, He, CO_2 and Ne at a certain temperature. Which of the following gases exerts minimum partial pressure?

MHT CET 2025 23rd April Evening Shift

Options:

A. H_2

B. He

C. CO_2

D. Ne

Answer: C

Solution:

Step 1: Recall the relations

For an ideal gas:

$PV = nRT,$

so the **partial pressure of each gas** under identical conditions depends on the number of moles present:

$P_i \propto n_i = \frac{\text{given mass}}{M_i},$

where M_i = molar mass.

Since all gases are given *equal masses*, the number of moles is **inversely proportional** to their molar mass:

$n \propto \frac{1}{M}.$

So gases with smaller molar mass will have more moles and therefore higher partial pressure; conversely, larger molar mass → fewer moles → lower partial pressure.

Step 2: Identify molar masses

- H_2: $M = 2\ \text{g/mol}$
- He: $M = 4\ \text{g/mol}$
- CO_2: $M = 44\ \text{g/mol}$
- Ne: $M = 20\ \text{g/mol}$

Step 3: Compare

For equal masses:

$n_{\text{H2}} \propto \frac{1}{2}, \quad n_{\text{He}} \propto \frac{1}{4}, \quad n_{\text{Ne}} \propto \frac{1}{20}, \quad n_{\text{CO2}} \propto \frac{1}{44}.$

Clearly,

$n_{\text{CO2}} < n_{\text{Ne}} < n_{\text{He}} < n_{\text{H2}} .$

So CO_2 (highest molar mass) will exert the **minimum partial pressure**.

✅ Final Answer:

Option C: CO_2

Question6

A container contains equal masses of H_2, He, CO_2 and Ne at a certain temperature. Which of the following gas exerts maximum partial pressure?

MHT CET 2025 23rd April Morning Shift

Options:

A. H_2

B. He

C. CO_2

D. Ne

Answer: A

Solution:

according to Dalton's law of partial pressure,

$P_i = x_i P_{total}$

Where P_i = partial pressure of gas

x_i = mole fraction of gas

P_{total} = total pressure of the gas mixture.

Since the masses of H_2, He, CO_2 and Ne are equal,

Number of moles $\propto \frac{1}{\text{molar mass}} \propto$ Mole fraction.

Since, H_2 has the highest mole fraction because of lowest molar mass, it will exert the maximum partial pressure.

Question7

What is the number of moles present in $1 \mathrm{~m}^3$ of any gas at STP?

MHT CET 2025 23rd April Morning Shift

Options:

A. 32.4

B. 54.6

C. 44.6

D. 28.2

Answer: C

Solution:

We are asked: **Number of moles in $1\,\mathrm{m}^3$ of any gas at STP**.

Step 1: Recall conditions at STP

- At STP (Standard Temperature and Pressure):
- Temperature = $273\,K$ (0 °C)
- Pressure = $1\,atm = 1.013 \times 10^5\,Pa$
- Molar volume of an ideal gas = $22.4\,L = 22.4 \times 10^{-3}\,m^3$

This means **1 mole** of gas occupies $0.0224\,m^3$.

Step 2: Find moles in $1\,m^3$

$n = \frac{1\,m^3}{0.0224\,m^3/mol} = \frac{1}{0.0224}$

$n \approx 44.64\,mol$

Step 3: Match with options

Closest value: **44.6**

✅ **Answer: Option C (44.6)**

Question8

The mass of $4.48\mathrm{dm}^3$ of certain gas is 5.6 g at STP. Assuming ideal behaviour, identify the probable gas from following.

MHT CET 2025 21st April Evening Shift

Options:

A. Cl_2

B. O_2

C. N_2

D. CH_4

Answer: C

Solution:

Step 1: Recall molar volume at STP.

At STP (Standard Temperature and Pressure: 0 °C, 1 atm), 1 mole of any ideal gas occupies 22.4 dm^3.

Step 2: Find number of moles in given volume.

Given volume = 4.48 dm^3.

$n = \frac{4.48}{22.4} = 0.2 \text{ mol}$

Step 3: Find molar mass using given mass.

Given mass = 5.6 g.

Molar mass:

$M = \frac{\text{mass}}{\text{moles}} = \frac{5.6}{0.2} = 28 \text{ g/mol}$

Step 4: Identify the gas.

Compare with molar masses of options:

- Cl_2: 71 g/mol
- O_2: 32 g/mol
- N_2: 28 g/mol ✅
- CH_4: 16 g/mol

Thus, the gas is **nitrogen** (N_2).

Final Answer:

$\boxed{N_2 \text{ (Option C)}}$

Question9

What is the SI unit of rate of diffusion for a gas?

MHT CET 2025 21st April Morning Shift

Options:

A. dm^{-3} s

B. $dm^3 s^{-1}$

C. $mL^{-1} s^{-1}$

D. mL s

Answer: B

Solution:

The rate of diffusion is defined as the volume of a gas diffusing per unit time.

So, its SI unit will be:

- Volume: dm^3 (cubic decimetre)
- Time: s^{-1} (per second)

Therefore, the correct SI unit is $dm^3 s^{-1}$.

Correct answer:

Option B $dm^3 s^{-1}$

Question10

A container consisting mixture of 28 g N_2, 8 gHe and 40 g Ne at 25°C . If the total pressure exerted by the gaseous mixture is 20 bar , What is the partial pressure exerted by N_2 ?

MHT CET 2025 20th April Evening Shift

Options:

A. 2 bar

B. 5 bar

C. 4 bar

D. 8 bar

Answer: C

Solution:

Step 1: **Calculate number of moles of each gas**

For N_2:

Molar mass of $N_2 = 28\,\text{g/mol}$

$n_{N_2} = \frac{28\text{ g}}{28\text{ g/mol}} = 1\,\text{mol}$

For He:

Molar mass of He $= 4\,\text{g/mol}$

$n_{He} = \frac{8\text{ g}}{4\text{ g/mol}} = 2\,\text{mol}$

For Ne:

Molar mass of Ne $= 20\,\text{g/mol}$

$n_{Ne} = \frac{40\text{ g}}{20\text{ g/mol}} = 2\,\text{mol}$

Step 2: **Find the total number of moles**

$n_{\text{total}} = n_{N_2} + n_{He} + n_{Ne} = 1 + 2 + 2 = 5\,\text{mol}$

Step 3: **Calculate mole fraction of** N_2

$x_{N_2} = \frac{n_{N_2}}{n_{\text{total}}} = \frac{1}{5} = 0.2$

Step 4: **Apply Dalton's law to find partial pressure**

According to Dalton's law:

$p_{N_2} = x_{N_2} \times p_{\text{total}}$

Given $p_{\text{total}} = 20$ bar,

$p_{N_2} = 0.2 \times 20\text{ bar} = 4\text{ bar}$

Final Answer:

Option C

$\boxed{4\text{ bar}}$

Question11

If compressibility factor of real gas is 1.05 at STP. What is molar volume of real gas?

MHT CET 2025 20th April Morning Shift

Options:

A. $22.40 \mathrm{dm}^3$

B. $21.33 \mathrm{dm}^3$

C. $23.52 \mathrm{dm}^3$

D. $23.05 \mathrm{dm}^3$

Answer: C

Solution:

Compressibility factor $(\mathrm{Z})=1.05$

Under ideal condition, 1 mole of gas accupies 22.4 L volume.

$$Z=\frac{V_{\text{real}}}{V_{\text{ideal}}}$$

$$1.05=\frac{V_{\text{real}}}{V_{\text{ideal}}}=\frac{V_{\text{real}}}{22.4 \mathrm{dm}^3}$$

$$V_{\text{real}} \quad=1.05 \times 22.4 \mathrm{dm}^3=23.52 \mathrm{dm}^3$$

Question12

Which from following is true according to Gay-Lussac's law?

MHT CET 2025 19th April Evening Shift

Options:

A. $\frac{\mathrm{V}}{\mathrm{T}}=$ constant at constant pressure and for fixed mass of gas.

B. $\frac{\mathrm{P}}{\mathrm{d}}=$ constant at constant temperature and fixed mass of gas.

C. $\frac{\mathrm{P}}{\mathrm{T}}=$ constant at constant volume and fixed mass of gas.

D. $\mathrm{P} \times \mathrm{V}=$ constant at constant temperature and for fixed mass of gas.

Answer: C

Solution:

Gay-Lussac's law states that for a fixed mass of gas at constant volume, the pressure of the gas is directly proportional to its absolute temperature.

Mathematically,

$\frac{P}{T}=\text{constant} \quad \text{(at constant volume and fixed mass of gas)}$

Now, let's check the options:

- **Option A:** $\frac{V}{T}=$ constant at constant pressure and for fixed mass of gas.

  (This is Charles's law.)

- **Option B:** $\frac{P}{d}=$ constant at constant temperature and fixed mass of gas.

  (This is incorrect.)

- **Option C:** $\frac{P}{T}=$ constant at constant volume and fixed mass of gas.

  (This is Gay-Lussac's law.)

- **Option D:** $P \times V=$ constant at constant temperature and for fixed mass of gas.

  (This is Boyle's law.)

Correct answer:

Option C $\frac{P}{T}=$ constant at constant volume and fixed mass of gas.

Question13

Four vessels of same volume consist equal masses of four gases $H_2, Cl_2,\ N_2$, and O_2 separately at same temperature. The pressure exerted by the gas is maximum for

MHT CET 2025 19th April Morning Shift

Options:

A. H_2

B. Cl_2

C. N_2

D. O_2

Answer: A

Solution:

Let us use the ideal gas equation:

$PV = nRT$

Where:

- P = pressure,
- V = volume (same for all),
- n = number of moles,
- R = universal gas constant,
- T = temperature (same for all).

Given:

- Volume (V) is same for all.
- Temperature (T) is same for all.
- Equal masses of all four gases.

The pressure will depend on the number of moles n.

Number of moles is:

$n = \frac{\text{Given mass}}{\text{Molar mass}}$

Since the given mass is same for all gases, n will be **maximum** for the gas with **minimum molar mass**.

Now, check molar mass:

- H_2: $2\,\mathrm{g/mol}$
- Cl_2: $71\,\mathrm{g/mol}$
- N_2: $28\,\mathrm{g/mol}$
- O_2: $32\,\mathrm{g/mol}$

H_2 has the **lowest molar mass** $\rightarrow$ it will have the **maximum number of moles**.

Thus, by $PV = nRT$, H_2 will exert the **maximum pressure**.

Final Answer:

Option A: H_2

Question14

Which from following compounds is NOT in gaseous phase at $25^{\circ}C$?

MHT CET 2024 16th May Morning Shift

Options:

A. ClF

B. BrF

C. IF_3

D. ClF_3

Answer: C

Solution:

ClF	Colourless gas
BrF	Pale brown gas
ClF_3	Colourleas gas
IF_3	Yellow powder

Question15

What is the numerical value of gas constant R in terms of $\mathrm{LatmK}^{-1}\ \mathrm{mol}^{-1}$?

MHT CET 2024 16th May Morning Shift

Options:

A. 0.085

B. 0.082

C. 8.314

D. 1.987

Answer: B

Solution:

The numerical value of the gas constant R in terms of $\mathrm{L\,atm\,K^{-1}\,mol^{-1}}$ is

$R = 0.0821\,\mathrm{L\,atm\,K^{-1}\,mol^{-1}}$.

This value is used in the ideal gas law equation:

$PV = nRT$

where:

P is the pressure in atmospheres,

V is the volume in liters,

n is the number of moles,

T is the temperature in Kelvin.

Note: The mentioned options do not include the exact value 0.0821, indicating a slight oversight in the options provided. Ensure to use $R = 0.0821$ for calculations involving the ideal gas constant in this unit system.

Question16

Which of the following equations gives combined relationship of Boyle's law and Charle's law?

MHT CET 2024 16th May Morning Shift

Options:

A. $\frac{P_1 V_2}{T_1}=\frac{P_2 V_1}{T_2}$

B. $\mathrm{n}=\frac{\mathrm{RT}}{\mathrm{PV}}$

C. $\frac{\mathrm{P}_1 \mathrm{~V}_1}{\mathrm{~T}_1}=\frac{\mathrm{P}_2 \mathrm{~V}_2}{\mathrm{~T}_2}$

D. $\mathrm{p}=\frac{\mathrm{RT}}{\mathrm{nV}}$

Answer: C

Solution:

The equation that represents the combined relationship of Boyle's Law and Charles's Law is given by:

Option C:

$\frac{P_1 V_1}{T_1}=\frac{P_2 V_2}{T_2}$

This equation is derived from the ideal gas law by combining the two individual gas laws:

Boyle's Law: For a given mass of gas at constant temperature, the product of pressure (P) and volume (V) is constant, i.e.,

$P_1 V_1=P_2 V_2$.

Charles's Law: At constant pressure, the volume of a gas is directly proportional to its temperature (T), i.e.,

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

By combining these two laws and assuming a constant amount of gas, the relation involving changes in pressure, volume, and temperature together is encapsulated by this combined gas law equation.

Question17

What is the volume occupied by 2.5 mol of ammonia gas at STP?

MHT CET 2024 15th May Evening Shift

Options:

A. $22.4 \mathrm{~dm}^3$

B. $25.0 \mathrm{~dm}^3$

C. $33.6 \mathrm{~dm}^3$

D. $56.0 \mathrm{~dm}^3$

Answer: D

Solution:

At standard temperature and pressure (STP), which is defined as a temperature of $273.15 \mathrm{~K}$ (0°C) and a pressure of 1 atm, one mole of an ideal gas occupies a volume of approximately $22.4 \mathrm{dm}^3$.

Using this information, the volume occupied by a given amount of gas at STP can be calculated using the formula:

$V=n \times V_m$

where:

V is the volume of the gas,

n is the number of moles of the gas,

V_m is the molar volume at STP, which is $22.4 \mathrm{dm}^3/\mathrm{mol}$.

For 2.5 moles of ammonia gas, the volume V can be calculated as follows:

$V=2.5 \mathrm{~mol} \times 22.4 \mathrm{dm}^3/\mathrm{mol}=56.0 \mathrm{dm}^3$

Thus, the volume occupied by 2.5 moles of ammonia gas at STP is:

Option D: $56.0 \mathrm{dm}^3$

Question18

3.4 moles of an ideal gas occupies volume of 68 mL at 300 K . What would be the pressure of gas? $\left(\mathrm{R}=8.314 \mathrm{~J} \mathrm{~K}^{-1} \mathrm{~mol}^{-1}\right)$

MHT CET 2024 15th May Evening Shift

Options:

A. 1.247×10^2 kPa

B. 2.431×10^3 kPa

C. 1.031×10^5 kPa

D. 3.247×10^5 kPa

Answer: A

Solution:

We use the **ideal gas equation:**

$$PV = nRT$$

Given:

- $n = 3.4\,\text{mol}$
- $V = 68\,\text{mL} = 68 \times 10^{-6}\,\text{m}^3$
- $T = 300\,\text{K}$
- $R = 8.314\,\text{J K}^{-1}\,\text{mol}^{-1}$

Calculate pressure:

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

$$P = \frac{3.4 \times 8.314 \times 300}{68 \times 10^{-6}}$$

$$P \approx 1.247 \times 10^5\,\text{Pa}$$

Convert to kPa:

$$P = 1.247 \times 10^2\,\text{kPa}$$

✅ **Correct answer:**

A — $1.247 \times 10^2\,\text{kPa}$

Question19

Which among the following gases is difficult to liquify?

MHT CET 2024 15th May Evening Shift

Options:

A. SO_2

B. Cl_2

C. NH_3

D. O_2

Answer: D

Solution:

Critical temperature (T_c) of O_2 is much below $0^{\circ}C$ and its critical pressure (P_c) value is also high. Consequently, it is difficult to liquify O_2.

Question20

Calculate the partial pressure exerted by dioxygen from a mixture of $32\ gO_2, 80\ gAr$ (mol.mass 40) and 4 g dihydrogen $\left(P_{\text{total}} \quad = 10\text{bar}\right)$.

MHT CET 2024 15th May Morning Shift

Options:

A. 1 bar

B. 2 bar

C. 4 bar

D. 5 bar

Answer: B

Solution:

To calculate the partial pressure of dioxygen (O_2) in a gas mixture, we can use Dalton's Law of Partial Pressures, which states that the partial pressure of a gas in a mixture is proportional to its mole fraction in the mixture:

$$P_i = X_i \cdot P_{\text{total}}$$

where P_i is the partial pressure of the gas, X_i is the mole fraction of the gas, and P_{total} is the total pressure of the mixture.

Step 1: Calculate moles of each gas

Moles of O_2:

Molar mass of $O_2 = 32\ \text{g/mol}$

$$n_{O_2} = \frac{32\ \text{g}}{32\ \text{g/mol}} = 1\ \text{mol}$$

Moles of Ar (Argon):

Molar mass of $\text{Ar} = 40\ \text{g/mol}$

$$n_{\text{Ar}} = \frac{80\ \text{g}}{40\ \text{g/mol}} = 2\ \text{mol}$$

Moles of H_2 (Dihydrogen):

Molar mass of $H_2 = 2\ \text{g/mol}$

$$n_{H_2} = \frac{4\ \text{g}}{2\ \text{g/mol}} = 2\ \text{mol}$$

Step 2: Calculate total moles of gas mixture

$$n_{\text{total}} = n_{O_2} + n_{\text{Ar}} + n_{H_2} = 1\ \text{mol} + 2\ \text{mol} + 2\ \text{mol} = 5\ \text{mol}$$

Step 3: Calculate mole fraction of O_2

$$X_{O_2} = \frac{n_{O_2}}{n_{\text{total}}} = \frac{1\ \text{mol}}{5\ \text{mol}} = 0.2$$

Step 4: Calculate partial pressure of O_2

Given $P_{\text{total}} = 10\ \text{bar}$,

$$P_{O_2} = X_{O_2} \times P_{\text{total}} = 0.2 \times 10\ \text{bar} = 2\ \text{bar}$$

Therefore, the partial pressure exerted by dioxygen is **2 bar**.

Option B: 2 bar

Question21

What is the volume of a gas at $1.032 \times 10^5\ \text{Nm}^{-2}$ if it occupies $1\ \text{dm}^3$ of volume at normal temperature and pressure?

MHT CET 2024 10th May Evening Shift

Options:

A. 0.982 dm^3

B. 1.3 dm^3

C. 1.5 dm^3

D. 1.7 dm^3

Answer: A

Solution:

The relationship between pressure and volume of a gas at constant temperature is described by Boyle's Law, which states that the product of the initial pressure and volume is equal to the product of the final pressure and volume. Mathematically, this can be expressed as:

$P_1 V_1 = P_2 V_2$

where:

P_1 is the initial pressure (normal atmospheric pressure, approximately $1.013 \times 10^5 \, \mathrm{Nm}^{-2}$),

V_1 is the initial volume ($1 \, \mathrm{dm}^3$),

P_2 is the final pressure ($1.032 \times 10^5 \, \mathrm{Nm}^{-2}$),

V_2 is the final volume.

Rearranging the formula to solve for V_2:

$V_2 = \frac{P_1 V_1}{P_2}$

Substituting the known values:

$V_2 = \frac{(1.013 \times 10^5 \, \mathrm{Nm}^{-2}) \times (1 \, \mathrm{dm}^3)}{1.032 \times 10^5 \, \mathrm{Nm}^{-2}}$

$V_2 = \frac{1.013 \times 10^5}{1.032 \times 10^5} \, \mathrm{dm}^3$

$V_2 = \frac{1.013}{1.032} \, \mathrm{dm}^3$

$V_2 \approx 0.982 \, \mathrm{dm}^3$

Thus, the volume of the gas at $1.032 \times 10^5 \, \mathrm{Nm}^{-2}$ is approximately $0.982 \, \mathrm{dm}^3$. Hence, the correct option is:

Option A: 0.982 dm^3

Question22

A gas occupies $11.2 \mathrm{~dm}^3$ at 105 kPa What is the volume if pressure is increased to 210 kPa ?

MHT CET 2024 10th May Morning Shift

Options:

A. $5.6 \mathrm{~dm}^3$

B. $16.8 \mathrm{~dm}^3$

C. $22.4 \mathrm{~dm}^3$

D. $33.6 \mathrm{~dm}^3$

Answer: A

Solution:

To solve this problem, we will use Boyle's Law, which states that for a given mass of gas at constant temperature, the product of pressure and volume is constant. Mathematically, Boyle's Law is expressed as:

$P_1 V_1=P_2 V_2$

Where:

P_1 is the initial pressure,

V_1 is the initial volume,

P_2 is the final pressure,

V_2 is the final volume.

Given that:

$P_1=105 \mathrm{kPa}$,

$V_1=11.2 \mathrm{dm}^3$,

$P_2=210 \mathrm{kPa}$.

We need to find V_2.

Rearranging the formula to solve for V_2, we have:

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

Substituting the known values:

$V_2 = \frac{105 \text{ kPa } \times 11.2 \text{ dm }^3}{210 \text{ kPa}}$

Performing the calculation:

$V_2 = \frac{1176 \text{ kPa } \cdot \text{dm }^3}{210 \text{ kPa}}$

$V_2 = 5.6 \text{ dm}^3$

Thus, the volume of the gas when the pressure is increased to 210 kPa is 5.6 dm^3, which corresponds to Option A.

Question23

Which of the following formula is used to calculate compressibility factor?

MHT CET 2024 4th May Evening Shift

Options:

A. $\mathrm{Z} = \frac{\mathrm{PVT}}{\mathrm{R}}$

B. $\mathrm{Z} = \frac{\mathrm{nRT}}{\mathrm{PV}}$

C. $\mathrm{Z} = \frac{\mathrm{PV}}{\mathrm{nRT}}$

D. $\mathrm{Z} = \left(\frac{\mathrm{PV}}{\mathrm{RT}}\right) \times \mathrm{n}$

Answer: C

Solution:

The compressibility factor, denoted as Z, is used to describe how much a real gas deviates from ideal gas behavior. The formula to calculate the compressibility factor is:

$Z = \frac{PV}{nRT}$

Where:

P is the pressure of the gas,

V is the volume of the gas,

n is the number of moles,

R is the ideal gas constant,

T is the temperature in Kelvin.

Thus, the correct option is **Option C**.

This equation is derived from the ideal gas law, which is $PV = nRT$. The compressibility factor approaches 1 for an ideal gas. For real gases, deviations from 1 reflect the influence of intermolecular forces and the size of gas particles.

Question24

Calculate the volume of gas at 1.25 atmosphere, if volume occupied by gas at 1 atmosphere and at same temperature is 25 mL .

MHT CET 2024 4th May Morning Shift

Options:

A. 15 mL

B. 20 mL

C. 25 mL

D. 35 mL

Answer: B

Solution:

The volume of a gas under different pressures, while keeping the temperature constant, can be determined using Boyle's Law. Boyle's Law states that the pressure and volume of a gas are inversely proportional when temperature is held constant. Mathematically, this is represented as:

$P_1V_1 = P_2V_2$

Where:

P_1 and V_1 are the initial pressure and volume.

P_2 and V_2 are the final pressure and volume.

Given:

$P_1 = 1\,\mathrm{atm}$

$V_1 = 25\,\mathrm{mL}$

$P_2 = 1.25\,\mathrm{atm}$

We need to find V_2.

Substitute the given values into Boyle's Law:

$1 \times 25 = 1.25 \times V_2$

Solve for V_2:

$V_2 = \frac{1\times25}{1.25}$

$V_2 = \frac{25}{1.25}$

$V_2 = 20\,\mathrm{mL}$

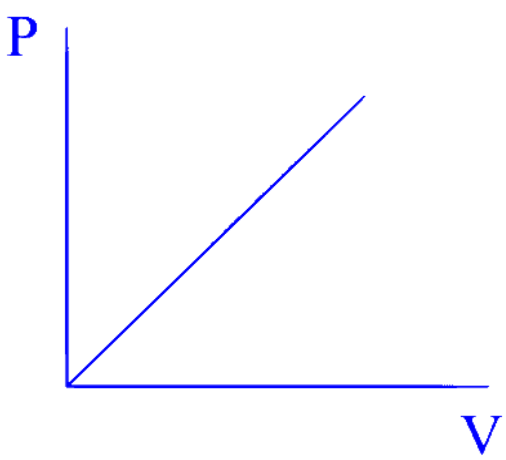
Therefore, the volume of the gas at 1.25 atmospheres is 20 mL, which corresponds to Option B.

Question25

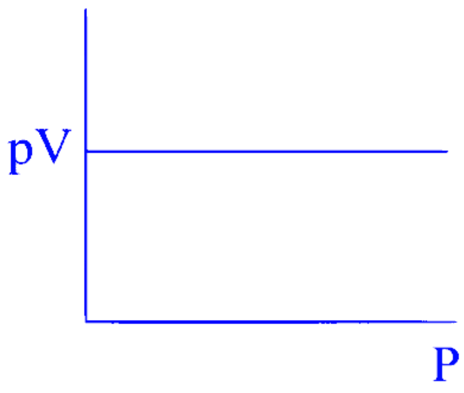
Which of the following graphs explains Boyle's law?

MHT CET 2024 3rd May Evening Shift

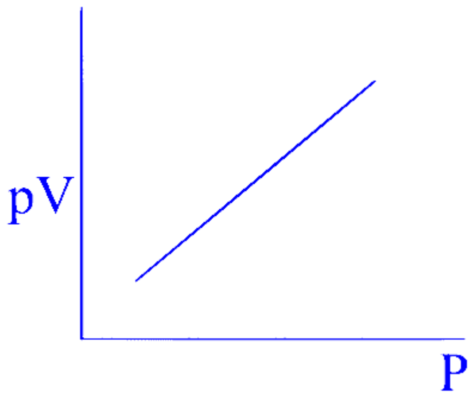
Options:

A. 

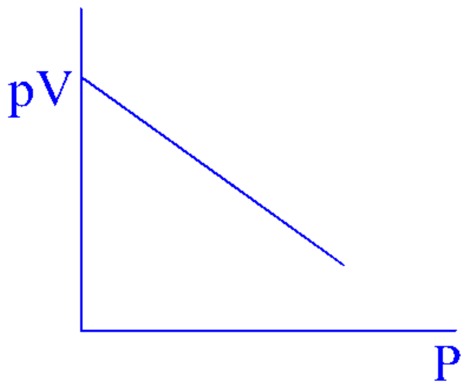
B.



C.



D.



Answer: B

Solution:

According to Boyle's law,

$P \propto \frac{1}{V}$

(at constant T and n)

$PV = \text{constant}$

Hence, the graph of PV vs P will be constant.

Question26

How many moles of dioxygen are present in $8.314 \times 10^{-3}\ \mathrm{m}^3$ **of it at 318 K having pressure** $3.18 \times 10^5 \mathrm{Nm}^{-2}$? $\left(\mathrm{R} = 8.314\ \mathrm{J\ K}^{-1}\ \mathrm{mol}^{-1}\right)$

MHT CET 2024 3rd May Morning Shift

Options:

A. 0.1 mole

B. 1.0 mole

C. 1.5 mole

D. 2.0 mole

Answer: B

Solution:

To find the number of moles of dioxygen, use the ideal gas equation:

$PV = nRT$

where:

$P = 3.18 \times 10^5\ \mathrm{N/m}^2$ (Pressure)

$V = 8.314 \times 10^{-3}\ \mathrm{m}^3$ (Volume)

n is the number of moles (what we are solving for)

$R = 8.314\ \mathrm{J\ K}^{-1}\mathrm{mol}^{-1}$ (Ideal gas constant)

$T = 318\ \mathrm{K}$ (Temperature)

Rearrange the ideal gas equation to solve for n:

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

Substitute the known values into the equation:

$n = \frac{(3.18 \times 10^5) \times (8.314 \times 10^{-3})}{8.314 \times 318}$

Calculate the numerator:

$3.18 \times 10^5 \times 8.314 \times 10^{-3} = 2644.602$

Calculate the denominator:

$8.314 \times 318 = 2645.652$

Now, calculate n:

$n = \frac{2644.602}{2645.652} \approx 0.9996$

Thus, the number of moles of dioxygen is approximately 1 mole. Therefore, the correct answer is:

Option B: 1.0 mole

Question27

Which of the following equations is true for 8.8×10^{-2} kg of carbon dioxide gas?

MHT CET 2024 2nd May Evening Shift

Options:

A. $\mathrm{PV} = 1.5\mathrm{RT}$

B. $\mathrm{PV} = \mathrm{RT}$

C. $\mathrm{PV} = 2\mathrm{RT}$

D. $\mathrm{PV} = 3\mathrm{RT}$

Answer: C

Solution:

To determine which equation is true for 8.8×10^{-2} kg of carbon dioxide gas, we need to use the ideal gas law:

$\mathrm{PV} = n\mathrm{RT}$

We need to calculate the number of moles n of the carbon dioxide (CO_2) gas. The molar mass of CO_2 is approximately 44 g/mol (since carbon has an atomic mass of about 12 and oxygen 16, and there are two oxygen atoms in carbon dioxide: $12 + 16 \times 2 = 44$ g/mol).

Given the mass of carbon dioxide is 0.088 kg (or 88 g), we calculate the number of moles as follows:

$n = \frac{\text{mass in grams}}{\text{molar mass}} = \frac{88 \text{ g}}{44 \text{ g/mol}} = 2 \text{ mol}$

Substitute $n = 2$ into the ideal gas law:

$\mathrm{PV} = n\mathrm{RT} = 2\mathrm{RT}$

Therefore, the equation that is true for 8.8×10^{-2} kg of carbon dioxide gas is:

Option C

$\mathrm{PV} = 2\mathrm{RT}$

Question28

Which from following gases of same mass exerts highest pressure at constant temperature?

MHT CET 2024 2nd May Morning Shift

Options:

A. H_2

B. N_2

C. O_2

D. Cl_2

Answer: D

Solution:

Let us consider the mass of each gas = 1 g

Order of molecular weight of the given gases:

$H_2 < N_2 < O_2 < Cl_2$

$\therefore$ The order of no. of moles: $H_2 > N_2 > O_2 > Cl_2$

From ideal gas equation, $n \propto V$

$\therefore$ Volume occupied by the gases is as follows: $H_2 > N_2 > O_2 > Cl_2$

At constant temperature, (Boyle's law)

$P \propto \frac{1}{V}$

Since volume occupied by Cl_2 gas is the least,

$\therefore$ Cl_2 gas exerts the highest pressure.

Question29

A neon-dioxygen mixture contains 64 g O_2 and 160 g Ne. If the total pressure is 25 bar, calculate the partial pressure of dioxygen.

MHT CET 2023 14th May Evening Shift

Options:

A. 5 bar

B. 7.5 bar

C. 10 bar

D. 20 bar

Answer: A

Solution:

$$
\begin{aligned}
& n_{O_2}=\frac{64}{32}=2 \text{ mol}; \quad n_{N_c}=\frac{160}{20}=8 \text{ mol} \\
& x_{O_2}=\frac{n_{O_2}}{n_{\text{Total}}}=\frac{2}{2+8}=\frac{2}{10}=0.2 \\
& P_{O_2}=x_{O_2} \times P_{\text{Total}} \quad=0.2 \times 25=5 \text{ bar}
\end{aligned}
$$

Question30

Find the temperature in degree Celsius if volume and pressure of 2 mole ideal gas is $20 \mathrm{~dm}^3$ and 4.926 atmospheres respectively. ($\mathrm{R}=0.0821 \mathrm{~dm}^3 \mathrm{~atm} \mathrm{~K}^{-1} \mathrm{~mol}^{-1}$)

MHT CET 2023 14th May Morning Shift

Options:

A. 273

B. 327

C. 600

D. 453

Answer: B

Solution:

To find the temperature of an ideal gas, we can use the ideal gas equation:

$PV = nRT$

where:

- P is the pressure in atmospheres,
- V is the volume in liters (dm^3),
- n is the number of moles,
- R is the gas constant ($0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$), and
- T is the temperature in Kelvin.

Given:

- $P = 4.926$ atmospheres,
- $V = 20 \text{ dm}^3 = 20 \text{ L}$,
- $n = 2$ moles, and
- $R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$.

We can rearrange the ideal gas equation to solve for T:

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

Plugging in the given values:

$T = \frac{(4.926 \text{ atm})(20 \text{ dm}^3)}{(2 \text{ mol})(0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1})}$

$T = \frac{98.52 \text{ atm}\cdot\text{dm}^3}{0.1642 \text{ dm}^3 \text{ atm mol}^{-1}}$

$T = 600 \text{ K}$

To convert the temperature from Kelvin to Celsius, use the conversion formula:

$\text{Celsius} = \text{Kelvin} - 273.15$

For our calculation:

$\text{Celsius} = 600 \text{ K} - 273.15$

$\text{Celsius} = 326.85 \approx 327\ ^\circ\text{C}$

Therefore, the temperature of the gas in degree Celsius is closest to Option B: 327.

Question31

A closed container contains mixture of non-reacting gases A and B. Partial pressure of A and B are 4.5 bar and 5.5 bar respectively. Find mole fractions of A and B respectively?

MHT CET 2023 13th May Evening Shift

Options:

A. 0.035 and 0.065

B. 0.055 and 0.045

C. 0.45 and 0.55

D. 0.55 and 0.45

Answer: C

Solution:

From Dalton's law of particle pressure

$$\begin{aligned} p_A &= \chi_A p_{\text{total}} \quad \text{ and } p_B = \chi_B p_{\text{total}} \\ p_{\text{total}} &= p_A + p_B \\ &= 4.5 + 5.5 \text{ bar} = 10 \text{ bar} \end{aligned}$$

Substituting the values in above formula,

$$\begin{aligned} &4.5 = \chi_A \times 10 \\ &\chi_A = 0.45 \\ &5.5 = \chi_B \times 10 \\ &\chi_B = 0.55 \end{aligned}$$

So, the mole fraction of A and B are 0.45 and 0.55 respectively.

Question32

The volume of a gas is $4 \mathrm{~dm}^3$ at $0^{\circ} \mathrm{C}$. Calculate new volume at constant pressure when the temperature is increased by $10^{\circ} \mathrm{C}$.

MHT CET 2023 13th May Morning Shift

Options:

A. $2.07 \mathrm{~dm}^3$

B. $3.21 \mathrm{~dm}^3$

C. $4.14 \mathrm{~dm}^3$

D. $6.54 \mathrm{~dm}^3$

Answer: C

Solution:

According to Charles' law,

$\frac{V_1}{T_1}=\frac{V_2}{T_2}$ (at constant P and n)

$T_1=0^{\circ} \mathrm{C}=273 \mathrm{~K}$

$T_2=10^{\circ} \mathrm{C}=283 \mathrm{~K}$

$$\therefore \quad \frac{4 \mathrm{dm}^3}{273 \mathrm{~K}}=\frac{V_2}{283 \mathrm{~K}}$$

$$\therefore \quad V_2=\frac{4 \times 283}{273}=4.1465 \mathrm{~dm}^3$$

Question33

What is the mass of $\mathrm{KClO}_{3(\mathrm{~s})}$ required to liberate 22. $4 \mathrm{~dm}^3$ oxygen at STP during thermal decomposition?

(**Molar Mass of** $\mathrm{KClO}_{3(\mathrm{~s})}=122.5 \mathrm{~g} / \mathrm{mol}$)

MHT CET 2023 12th May Evening Shift

Options:

A. 122.5 g

B. 81.67 g

C. 10.25 g

D. 8.16 g

Answer: B

Solution:

$$\underset{[2 \text{ moles}]}{2KClO_3} \longrightarrow 2KCl + \underset{[3 \text{ moles}]}{3O_2\uparrow}$$

2 moles of $KClO_3 = 2 \times 122.5 = 245$ g

3 moles of O_2 at STP occupy $= \left(3 \times 22.4 dm^3\right)$

Thus, 245 g of potassium chlorate will liberate 67.2 dm^3 of oxygen gas.

Let 'x' gram of $KClO_3$ liberate 22.4 dm^3 of oxygen gas at S.T.P.

$\therefore \quad x = \frac{245 \times 22.4}{3 \times 22.4} = 81.67$ g

----

Question34

Equal masses of $H_{2(g)}$ and $He_{(g)}$ are enclosed in a container at constant temperature. The ratio of partial pressure of H_2 to He is

MHT CET 2023 12th May Evening Shift

Options:

A. 1 : 1

B. 1 : 2

C. 2 : 1

D. 1 : 4

Answer: C

Solution:

Let the mass be ',x'.

$n_{H_2} = \frac{xg}{2 \text{ g mol}^{-1}} = \frac{x}{2}$

$$n_{He} = \frac{xg}{4 \text{ g mol}^{-1}} = \frac{x}{4}$$

$$n_{Total} = \frac{3x}{4}$$

Now,

$$x_{H_2} = \frac{n}{n_{Total}} = \frac{x/2}{3x/4} = \frac{2}{3}$$

$$x_{H_e} = \frac{n}{n_{Total}} = \frac{x/4}{3x/4} = \frac{1}{3}$$

Now,

$$P_{H_2} = x_{H_2} \times P_{Total} = \frac{2}{3} \times P$$

$$P_{He} = x_{He} \times P_{Total} = \frac{1}{3} \times P$$

$$P_{H_2} : P_{H_e} = \frac{2/3P}{1/3P} = 2 : 1$$

Question35

What volume of $CO_{2(g)}$ at STP is obtained by complete combustion of 6 g carbon?

MHT CET 2023 12th May Morning Shift

Options:

A. 22.4 dm^3

B. $11.2 \mathrm{~dm}^3$

C. $5.6 \mathrm{~dm}^3$

D. $2.24 \mathrm{~dm}^3$

Answer: B

Solution:

$$
\begin{aligned}
& \mathrm{C}_{(\mathrm{s})}+\mathrm{O}_{2(\mathrm{~g})} \longrightarrow \mathrm{CO}_{2(\mathrm{~g})} \\
& 1 \mathrm{~mol} \mathrm{C} \equiv 1 \mathrm{~mol} \mathrm{CO}_{2(\mathrm{~g})} \\
& 6 \mathrm{~g} \mathrm{C}=0.5 \mathrm{~mol} \mathrm{C} \\
\therefore \quad & 0.5 \mathrm{~mol} \mathrm{C} \equiv 0.5 \mathrm{~mol} \mathrm{CO}_{2(\mathrm{~g})} \\
& \text{At STP, } 1 \mathrm{~mol} \mathrm{CO}_{2(\mathrm{~g})} \equiv 22.4 \mathrm{~dm}^3 \\
\therefore \quad & 0.5 \mathrm{~mol} \mathrm{CO}_{2(\mathrm{~g})}=11.2 \mathrm{~dm}^3
\end{aligned}
$$

Question36

A hot air balloon has volume of $2000 \mathrm{~dm}^3$ at $99^{\circ} \mathrm{C}$. What is the new volume if air in balloon cools to $80^{\circ} \mathrm{C}$?

MHT CET 2023 12th May Morning Shift

Options:

A. $2428.9 \mathrm{~dm}^3$

B. $2656.9 \mathrm{~dm}^3$

C. $2814.9 \mathrm{~dm}^3$

D. $1897.8 \mathrm{~dm}^3$

Answer: D

Solution:

Using Charles' law,

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

$$\therefore \quad \frac{2000 dm^3}{372\ K} = \frac{V_2}{353\ K}$$

$$\therefore \quad V_2 = \frac{2000 \times 353}{372} = 1897.8\ dm^3$$

According to Charles' law, at constant pressure, the volume of a fixed mass of a gas is directly proportional to its temperature in kelvin. Therefore, as temperature decreases, volume will decrease. Hence, only option (D) is valid.

Question37

If N_2 gas is compressed at 2 atmosphere from 9.0 L to 3.0 L at 300 K, find the final pressure at same temperature.

MHT CET 2023 11th May Evening Shift

Options:

A. 1.66 atm

B. 3.32 atm

C. 6.0 atm

D. 9.0 atm

Answer: C

Solution:

According to Boyle's law,

$P_1 V_1 = P_2 V_2$ (at constant n and T)

$$\therefore \quad P_2 = \frac{P_1 V_1}{V_2} = \frac{2 \times 9.0}{3.0} = 6.0 \text{ bar}$$

At constant temperature, the volume of a given amount of gas is inversely proportional to its pressure. Therefore, if the volume is reduced to one-third of its original value, the pressure will correspondingly increase by a factor of three. Hence, correct answer is option (C).

Question38

What is the volume in dm^3 occupied by $3 \mathrm{~mol}$ of ammonia gas at STP?

MHT CET 2023 11th May Morning Shift

Options:

A. 2.24

B. 22.4

C. 56.0

D. 67.2

Answer: D

Solution:

$$
\begin{aligned}
&\text{Number of moles of a gas (n)} \\
&=\frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}} \\
&\therefore \quad \text{Volume of ammonia gas at STP} \\
&\quad = \text{Number of moles of the gas (n)} \\
&\quad \times \text{Molar volume of the gas} \\
&= 3 \mathrm{~mol} \times 22.4 \mathrm{~dm}^3 \mathrm{~mol}^{-1} \\
&= 67.2 \mathrm{~dm}^3
\end{aligned}
$$

Question39

Which of the following formulae is used to determine compressibility factor for measurement of deviation from ideal behaviour?

MHT CET 2023 11th May Morning Shift

Options:

A. $Z = \frac{nRT}{PV}$

B. $Z = \frac{PV}{nRT}$

C. $Z = \frac{nRT}{V}$

D. $Z = \frac{nRT}{P}$

Answer: B

Solution:

The **compressibility factor** (Z) is used to measure the **deviation of a real gas from ideal behavior** and is defined as:

$$Z = \frac{PV}{nRT}$$

For an ideal gas, $Z = 1$.

✅ **Correct answer: B** — $Z = \frac{PV}{nRT}$

Question40

What is the volume of 1 mole real gas at STP $\left(V_0 = 22.4\ \mathrm{dm}^3\right)$, if compressibility factor of real gas is 1.1 at STP?

MHT CET 2023 10th May Evening Shift

Options:

A. $22.40\ \mathrm{dm}^3$

B. $23.64\ \mathrm{dm}^3$

C. $24.64\ \mathrm{dm}^3$

D. $23.50\ \mathrm{dm}^3$

Answer: C

Solution:

$$Z=\frac{V_{\text{real}}}{V_{\text{ideal}}}$$

$$V_{\text{real}} \quad = Z \times V_{\text{ideal}} \quad = 1.1 \times 22.4 = 24.64 \text{ dm}^3$$

Alternate method:

$$\mathrm{Z}=\frac{\mathrm{PV}}{\mathrm{nRT}}$$

$$1.1=\frac{1 \times \mathrm{V}}{1 \times 0.08206 \times 273}$$

$$1.1=\frac{\mathrm{V}}{0.08206 \times 273}$$

$$\therefore \quad \mathrm{V}=24.64 \mathrm{~dm}^3$$

Question41

What is new temperature of a gas when its initial volume 3 dm^3 at 300 K is doubled at constant pressure?

MHT CET 2023 10th May Morning Shift

Options:

A. 450 K

B. 600 K

C. 750 K

D. 900 K

Answer: B

Solution:

Using Charles' law,

$$\frac{\mathrm{V}_1}{\mathrm{T}_1}=\frac{\mathrm{V}_2}{\mathrm{T}_2}$$

$$\therefore \quad \frac{3 \mathrm{dm}^3}{300 \mathrm{~K}}=\frac{2 \times 3 \mathrm{dm}^3}{\mathrm{T}_2}$$

$$\therefore \quad \mathrm{T}_2=\frac{2 \times 3 \times 300}{3}=600 \mathrm{~K}$$

At constant pressure, the volume of a given amount of gas is directly proportional to its temperature in Kelvin. Therefore, if the volume doubles, the temperature will be doubled. Hence, correct answer is option (B).

Question42

Which of the following is the SI unit of coefficient of viscosity?

MHT CET 2023 9th May Evening Shift

Options:

A. $\mathrm{N\ s^{-1}\ m^{-2}}$

B. $\mathrm{N\ s\ m^{-2}}$

C. $\mathrm{N\ s\ m^{2}}$

D. $\mathrm{N\ s^{-1}\ m^{-1}}$

Answer: B

Solution:

The SI unit of coefficient of viscosity (η) is:

$$\mathrm{N\ s\ m^{-2}}$$

This is also equivalent to Pa·s (pascal–second).

✅ Correct answer: B — N s m^{-2}

Question43

At $0^{\circ}\mathrm{C}$ a gas occupies 22.4 liters. What is the temperature in Kelvin to reach the volume of 224 liters?

MHT CET 2023 9th May Morning Shift

Options:

A. 546 K

B. 273 K

C. 2730 K

D. 5460 K

Answer: C

Solution:

According to Charles' law,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \text{(at constant n and P)}$$

$$\therefore \quad \mathrm{T_2} = \frac{\mathrm{V_2 \times T_1}}{\mathrm{V_1}} = \frac{224 \times 273}{22.4} = 2730 \mathrm{~K}$$

Question44

A certain mass of a gas occupies volume of 250 mL at 2 atm. Calculate the volume of gas if pressure is increased to 2.5 atm at constant temperature.

MHT CET 2021 24th September Evening Shift

Options:

A. 352.0 mL

B. 300.0 mL

C. 200 mL

D. 443.0 mL

Answer: C

Solution:

To solve this problem, we can use Boyle's Law, which states that the pressure of a given mass of gas is inversely proportional to its volume at a constant temperature. Mathematically, Boyle's Law is represented as:

$P_1 \times V_1 = P_2 \times V_2$

Where:

- P_1 is the initial pressure
- V_1 is the initial volume
- P_2 is the final pressure
- V_2 is the final volume

Given:

- $P_1 = 2 \text{ atm}$
- $V_1 = 250 \text{ mL}$
- $P_2 = 2.5 \text{ atm}$

We need to find V_2. Rearranging Boyle's Law to solve for V_2, we get:

$V_2 = \frac{P_1 \times V_1}{P_2}$

Substituting the given values:

$V_2 = \frac{2 \text{ atm} \times 250 \text{ mL}}{2.5 \text{ atm}}$

Simplifying the expression:

$V_2 = \frac{500 \text{ mL atm}}{2.5 \text{ atm}}$

$V_2 = 200 \text{ mL}$

Therefore, the volume of the gas when the pressure is increased to 2.5 atm at constant temperature is 200 mL.

The correct answer is:

Option C: 200 mL

Question45

At $300 \mathrm{~K}, 22 \mathrm{~g}$ of CO_2 gas exerts a pressure of 5 atmosphere. What is the volume of the gas at the same temperature? $\left(\mathrm{R}=0.0821 \mathrm{~L} \text{ atm } \mathrm{K}^{-1} \mathrm{~mol}^{-1}\right)$

MHT CET 2021 24th September Morning Shift

Options:

A. 5.61 dm^3

B. 8.20 dm^3

C. 2.46 dm^3

D. 3.80 dm^3

Answer: C

Solution:

To find the volume of the CO_2 gas at a given temperature and pressure, we can use the Ideal Gas Law, which is given by:

$PV = nRT$

Where:

- P is the pressure
- V is the volume
- n is the number of moles
- R is the universal gas constant
- T is the temperature in Kelvin

Here, we are given:

- $P = 5$ atmospheres
- $T = 300$ Kelvin
- $R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}$
- Mass of $CO_2 = 22$ grams

First, we need to find the number of moles (n) of CO_2. The molar mass of CO_2 is:

Molar mass of $CO_2 = 12 \text{ g/mol} \ (for\ Carbon) + 2 \times 16 \text{ g/mol} \ (for\ Oxygen) = 44 \text{ g/mol}$

Now, we can calculate the number of moles:

$n = \frac{\text{mass of } \quad CO_2}{\text{molar mass of } \quad CO_2} = \frac{22 \text{ g}}{44 \text{ g/mol}} = 0.5 \text{ mol}$

Using the Ideal Gas Law to find the volume:

$PV = nRT$

Rearranging to solve for V:

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

Substitute the values in:

$V = \frac{(0.5 \text{ mol})(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1})(300 \text{ K})}{5 \text{ atm}}$

Simplifying the expression:

$V = \frac{(0.5 \times 0.0821 \times 300)}{5}$

$V = \frac{12.315}{5} = 2.463 \text{ L}$

Since 1 liter is equivalent to 1 cubic decimeter (dm^3), the volume in dm^3 is:

$V \approx 2.46 \text{ dm}^3$

Therefore, the correct option is:

Option C 2.46 dm^3

Question46

What is the mass of 33.6 dm^3 of methane gas at S.T.P.?

MHT CET 2021 23th September Morning Shift

Options:

A. $4.8 \times 10^{-2} \text{ kg}$

B. $3.3 \times 10^{-2} \text{ kg}$

C. $1.6 \times 10^{-2} \text{ kg}$

D. $2.4 \times 10^{-2} \text{ kg}$

Answer: D

Solution:

$$\begin{aligned} &22.4 \text{ dm}^3 \text{ of } CH_4 = 16 \text{ g at STP} \\ &\therefore 33.6 \text{ dm}^3 \text{ of } CH_4 = \frac{16 \times 33.6}{22.4} \\ &= 24 \text{ g} \quad = 2.4 \times 10^{-2} \text{ kg} \end{aligned}$$

Question47

A gas has a volume of 3.4 L at 298 K. What is the final temperature if the volume of gas increases to 6.8 L ?

MHT CET 2021 23th September Morning Shift

Options:

A. 596 K

B. 412 K

C. 298 K

D. 149 K

Answer: A

Solution:

To solve this problem, we can use Charles's Law which states that the volume of a gas is directly proportional to its temperature (in kelvin) when the pressure and the amount of gas are kept constant. The formula for Charles's Law is:

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

Where:

V_1 = Initial volume = 3.4 L

T_1 = Initial temperature = 298 K

V_2 = Final volume = 6.8 L

T_2 = Final temperature = ?

We need to solve for T_2:

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

Rearranging for T_2, we get:

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

Now, substitute the given values:

$T_2 = \frac{6.8 \text{ L} \cdot 298 \text{ K}}{3.4 \text{ L}}$

Simplifying the equation:

$T_2 = \frac{6.8 \cdot 298}{3.4}$

$T_2 = \frac{2026.4}{3.4}$

$T_2 = 596 \text{ K}$

So, the final temperature is 596 K. The correct answer is:

Option A: 596 K

Question48

A balloon contains $2.27 \mathrm{~L}$ air and has a pressure of $1.013 \times 10^5 \mathrm{~Nm}^{-2}$. The balloon rises to a certain height and expands to volume of $4540 \mathrm{~mL}$. What is the final pressure of air in balloon?

MHT CET 2021 23th September Morning Shift

Options:

A. $2.026 \times 10^2 \mathrm{~Nm}^{-2}$

B. $5.065 \times 10^4 \mathrm{~Nm}^{-2}$

C. $4.540 \times 10^4 \mathrm{~Nm}^{-2}$

D. $5.065 \times 10^{-4} \mathrm{~Nm}^{-2}$

Answer: B

Solution:

To determine the final pressure of the air in the balloon as it expands, we can use Boyle's Law. Boyle's Law states that for a fixed amount of gas at a constant temperature, the pressure of the gas is inversely proportional to its volume. Mathematically, Boyle's Law is expressed as:

$P_1 V_1=P_2 V_2$

Where:

- P_1 is the initial pressure.
- V_1 is the initial volume.
- P_2 is the final pressure.
- V_2 is the final volume.

Given data:

- Initial pressure, $P_1=1.013 \times 10^5 \mathrm{~Nm}^{-2}$
- Initial volume, $V_1=2.27 \mathrm{~L}=2.27 \times 10^3 \mathrm{~mL}$
- Final volume, $V_2=4540 \mathrm{~mL}$

We are required to find the final pressure, P_2. Using the relation:

$P_2 = \frac{P_1 V_1}{V_2}$

Substituting the given values:

$P_2 = \frac{1.013 \times 10^5 \text{ Nm}^{-2} \times 2.27 \times 10^3 \text{ mL}}{4540 \text{ mL}}$

Calculating the above expression:

$P_2 = \frac{1.013 \times 10^5 \times 2.27 \times 10^3}{4540}$

$P_2 = \frac{1.013 \times 2.27 \times 10^8}{4.54 \times 10^3}$

$P_2 = \frac{2.2971 \times 10^8}{4.54 \times 10^3}$

$P_2 = 5.065 \times 10^4 \text{ Nm}^{-2}$

Thus, the final pressure of the air in the balloon is $5.065 \times 10^4 \text{ Nm}^{-2}$.

Therefore, the correct option is:

Option B: $5.065 \times 10^4 \text{ Nm}^{-2}$

Question49

The volume of a gas at 0°C is 2 dm^3. What is its volume if temperature is decreased by 272°C ?

MHT CET 2021 22th September Evening Shift

Options:

A. $\left(\frac{3}{272}\right) \text{dm}^3$

B. $\left(\frac{2}{272}\right) \text{dm}^3$

C. $\left(\frac{4}{273}\right) \text{dm}^3$

D. $\left(\frac{2}{273}\right) \text{dm}^3$

Answer: D

Solution:

To solve the problem, we can apply Charles's Law, which describes how gases tend to expand when heated in a constant-pressure condition. Charles's Law states that the volume of a gas is directly proportional to its temperature in kelvins, which can be expressed as:

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

where:

- V_1 and T_1 are the initial volume and temperature (in kelvins) of the gas,
- V_2 and T_2 are the final volume and temperature (in kelvins) of the gas.

To proceed, let's convert all temperatures to kelvins. The initial temperature is $0^{\circ}\mathrm{C}$, which is equivalent to $273\mathrm{K}$ $(0 + 273)$. The problem states that the temperature is decreased by $272^{\circ}\mathrm{C}$, resulting in a final temperature of $-272^{\circ}\mathrm{C}$ or $1\mathrm{K}$ $(0 - 272 + 273)$.

Given that the initial volume $V_1 = 2\ \mathrm{dm}^3$ and the initial temperature $T_1 = 273\mathrm{K}$, we aim to find the final volume V_2 when $T_2 = 1\mathrm{K}$.

Substitute the given values into Charles's Law equation:

$\frac{2\ \mathrm{dm}^3}{273\mathrm{K}} = \frac{V_2}{1\mathrm{K}}$

You can solve for V_2 as follows:

$V_2 = 2\ \mathrm{dm}^3 \times \frac{1\mathrm{K}}{273\mathrm{K}} = \frac{2}{273}\ \mathrm{dm}^3$

Therefore, the volume of the gas when the temperature is decreased by $272^{\circ}\mathrm{C}$ (resulting in a final temperature of 1 Kelvin) is $\frac{2}{273}\ \mathrm{dm}^3$, which corresponds to Option D.

Question50

What is the value of critical temperature of water?

MHT CET 2021 22th September Morning Shift

Options:

A. 647 K

B. 312 K

C. 346 K

D. 493 K

Answer: A

Solution:

The critical temperature of a substance is the temperature above which it cannot exist as a liquid, regardless of the pressure applied. For water, the critical temperature holds significant importance in various scientific and industrial applications.

The critical temperature of water is:

Option A: 647 K

Thus, the correct answer is Option A (647 K).

Question51

What is the volume occupied by 16 g methane gas at STP?

MHT CET 2021 21th September Evening Shift

Options:

A. 1140 cm^3

B. 22400 cm^3

C. 214 cm^3

D. 12.4 cm^3

Answer: B

Solution:

1 mole of CH_4 = 16 g of CH_4 = 22400 cm^3 at STP

Question52

Keeping temperature constant the pressure of 11.2 dm^3 of a gas was increased from 105 kPa to 420 kPa. What is the new volume of gas?

MHT CET 2021 20th September Evening Shift

Options:

A. 1.4 dm^3

B. 7.0 dm^3

C. 5.6 dm^3

D. 2.8 dm^3

Answer: D

Solution:

$$\begin{array}{ll} V_1 = 11.2\ \mathrm{dm}^3, & P_1 = 105\ \mathrm{kPa} \\ V_2 = ?, & P_2 = 420\ \mathrm{kPa} \end{array}$$

According to Boyle's law,

$$P_1V_1 = P_2V_2$$
$$\therefore V_2 = \frac{P_1V_1}{P_2} = \frac{105\ \mathrm{kPa} \times 11.2\ \mathrm{dm}^3}{420\ \mathrm{kPa}} = 2.8\ \mathrm{dm}^3$$

Question53

A gas occupies 11.2 dm^3 at 105 kPa. What is it's volume if pressure is increased to 210 kPa?

MHT CET 2021 20th September Morning Shift

Options:

A. 22.4 dm^3

B. 33.6 dm^3

C. 5.6 dm^3

D. 16.8 dm^3

Answer: C

Solution:

$P_1 = 105$ kPa, $V_1 = 11.2$ dm^3
$P_2 = 210$ kPa, $V_2 = ?$

According to Boyle's law, $P_1V_1 = P_2V_2$

$\therefore V_2 = \frac{P_1V_1}{P_2} = \frac{105 \text{ kPa} \times 11.2 \text{ dm}^3}{210 \text{ kPa}} = 5.6 \text{ dm}^3$

Question54

The volume of oxygen required for complete combustion of 0.25 mole of methane at STP is

MHT CET 2020 16th October Evening Shift

Options:

A. 22.4 dm^3

B. 11.2 dm^3

C. 7.46 dm^3

D. 5.6 dm^3

Answer: B

Solution:

Given,

Volume of methane $= 0.25$ mol $= 5.6$ L

[Multiply mole value by molar volume constant, 22.4 L]

Combustion of methane (CH_4)

$$\underset{\text{Methane}}{CH_4} + 2O_2 \longrightarrow CO_2 + 2H_2O$$

As, we know that, volume of 1 mole of gas is 22.4 L . Amount of oxygen is required for the combustion of 22.4 L of methane $= (2 \times 22.4)\mathrm{L} = 44.8\ \mathrm{L}$

$\therefore$ Amount of oxygen required for the combustion of 5.6 L of methane

$$= \frac{44.8}{22.4} \times 5.6 = 11.2\ \mathrm{L} = 11.2\ \mathrm{dm}^3$$
$$\left[\because 1\ \mathrm{L} = 1\ \mathrm{dm}^3\right]$$

Hence, 11.2 dm^3 of oxygen required for the complete combustion of 0.25 mole of methane.

Question55

If, 2 moles of an ideal gas at 546 K has volume of 44.8 L, then what will be it's pressure? $(R = 0.082)$

MHT CET 2020 16th October Evening Shift

Options:

A. 3.129 atm

B. 1.098 atm

C. 1.998 atm

D. 2.408 atm

Answer: C

Solution:

Given,

$n = 2\ \mathrm{mol}, T = 546\ \mathrm{K}, V = 44.8\ \mathrm{L}, R = 0.082\ \mathrm{L\ atm\ mol^{-1}\ K^{-1}}$

According to ideal gas equation,

$$\Rightarrow \quad pV = nRT$$
$$\Rightarrow \quad \frac{2\ \mathrm{mol} \times 0.082\ \mathrm{L\ atm\ mol^{-1}\ K^{-1}} \times 546\ \mathrm{K}}{V}$$
$$= \frac{89.544}{44.8}\ \mathrm{atm}$$
$$= 1.998\ \mathrm{atm}$$

Question56

Which mixture is used for respiration by deep sea divers?

MHT CET 2019 3rd May Morning Shift

Options:

A. $Ne + O_2$

B. $Kr + O_2$

C. $Ar + O_2$

D. $He + O_2$

Answer: D

Solution:

Scuba divers must cope with high concentrations of dissolved gases while breathing air at high pressure, underwater. Increased pressure increases the solubility of atmospheric gases in blood. When divers come towards surface, the pressure gradually decreases. This blocks capillaries and creates a medical condition called bends. To avoid bends, the tanks are used by scuba divers that are filled with air diluted with helium (11.7% helium, 56.2% nitrogen and 32.1% oxygen). Thus, option (d) is correct.

Question57

The volume of 1 mole of any pure gas at standard temperature and pressure is always equal to

MHT CET 2019 2nd May Evening Shift

Options:

A. $0.022414\ m^3$

B. $22.414 \mathrm{~m}^3$

C. $2.2414 \mathrm{~m}^3$

D. $0.22414 \mathrm{~m}^3$

Answer: A

Solution:

At standard temperature and pressure (STP), which is defined as a temperature of 273.15 K (0°C) and a pressure of 1 atm, the volume of 1 mole of any ideal gas is approximately 22.414 liters. In cubic meters, this volume is equivalent to:

$22.414 \mathrm{~L}=0.022414 \mathrm{~m}^3$

Thus, the correct answer is:

Option A : $0.022414 \mathrm{~m}^3$

Question58

What is the density of water vapour at boiling point of water?

MHT CET 2019 2nd May Evening Shift

Options:

A. $1 \times 10^{-4} \mathrm{~g} \mathrm{~cm}^{-3}$

B. $1 \mathrm{~g} \mathrm{~cm}^{-3}$

C. $6 \times 10^{-4} \mathrm{~g} \mathrm{~cm}^{-3}$

D. $4 \times 10^{-4} \mathrm{~g} \mathrm{~cm}^{-3}$

Answer: C

Solution:

At the boiling point of water, which is 100°C or $373 \mathrm{~K}$, the density of water vapor can be calculated using the ideal gas law. The ideal gas law is given by:

$PV = nRT$

Where:

P is the pressure (at boiling point, P is approximately 1 atm or 101325 Pa)

V is the volume

n is the number of moles

R is the ideal gas constant ($8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$)

T is the temperature in Kelvin

The density ρ of a gas is related to its mass and volume:

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

Where m is the mass. Additionally, the mass of the vapor can be represented as the molar mass M times the number of moles n:

$m = nM$

Substituting $m = nM$ into the density equation and combining it with the ideal gas law, we get:

$\rho = \frac{nM}{V} = \frac{PM}{RT}$

For water vapor, the molar mass M is approximately 18.015 g/mol. Plugging in the values:

$P = 101325 \text{ Pa}$

$M = 18.015 \text{ g/mol}$

$R = 8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$

$T = 373 \text{ K}$

The calculation becomes:

$\rho = \frac{101325 \times 18.015}{8.314 \times 373} \text{ g} \cdot \text{m}^{-3}$

Simplifying, this gives:

$\rho \approx 0.598 \text{ g/m}^3 \approx 6 \times 10^{-4} \text{ g/cm}^3$

Thus, the density of water vapor at the boiling point of water is best approximated by:

Option C: $6 \times 10^{-4} \text{ g} \cdot \text{cm}^{-3}$

Question59

A cold drink bottle contains 200 mL liquid, in which CO_2 is 0.1 molar. Considering CO_2 as an ideal gas the volume of the dissolved CO_2 at S.T.P is

MHT CET 2019 2nd May Morning Shift

Options:

A. 22.4 L

B. 0.224 L

C. 2.24 L

D. 0.448 L

Answer: D

Solution:

Given,

Molarity $= 0.1\text{M}$

Volume $= 200 \text{ mL}$

$\because$ Molarity $= \frac{\text{number of moles of solute}}{\text{volume of solution} \qquad (\text{mL})} \times 1000$

$\therefore$ Number of moles $= \frac{0.1 \times 200}{1000} = 0.02$

Now,

1 mole of dissolved CO_2 occupies $= 22.4$ L volume at STP.

$\therefore$ 0.02 moles of dissolved CO_2 occupies $= 224 \times 0.02 = 0.448$ L Volume at STP.
