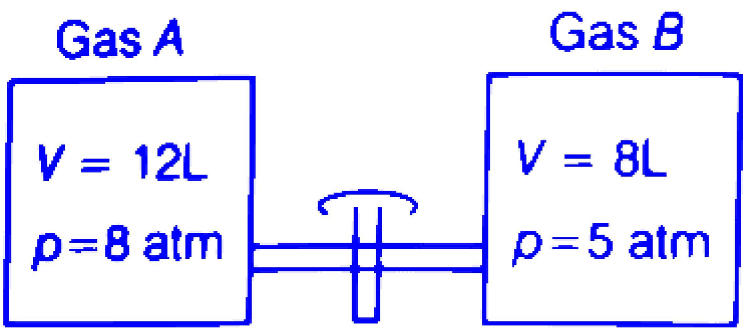
Question1

Two vessels are filled with ideal gases A and B and are connected through a pipe of zero volume as shown in figure. The stop cock is opened and the gases are allowed to mix homogeneously and the temperature is

kept constant. The partial pressures of A and B respectively (in atm) are



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Options:

A.

8.0, 5

B.

9.6, 4

C.

6.4, 4

D.

4.8, 2

Answer: D

Solution:

Using formula for partial pressure

$$p_i = \frac{n_i}{n_{\text{total}}} \times p_{\text{total}}$$

$$p_{\text{total}} = \frac{n_{\text{total}} \cdot R \cdot T}{V} = \frac{136}{20} = 6.8 \text{ atm}$$

$$p_A = \frac{96}{136} \times 6.8 = 4.8 \text{ atm}$$

$$p_B = \frac{40}{136} \times 6.8 = 2 \text{ atm}$$

Question2

The RMS velocity of dihydrogen is $\sqrt{7}$ times more than that of dinitrogen. If T_{H_2} and T_{N_2} are the temperatures of dihydrogen and dinitrogen, then the correct relationship between them is

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Options:

A.

$T_{H_2} = T_{N_2}$

B.

$T_{H_2} > T_{N_2}$

C.

$T_{H_2} = \sqrt{7} T_{N_2}$

D.

$T_{H_2} = \frac{T_{N_2}}{2}$

Answer: D

Solution:

v_{rms} is given by, $v_{\text{rms}} = \sqrt{\frac{3RT}{M}}$

Also, it is given that

$$
\begin{aligned}
v_{\text{rms}(\text{H}_2)} &= \sqrt{7} v_{\text{rms}(\text{ N}_2)} \\
\sqrt{\frac{3RT_{\text{H}_2}}{M_{\text{H}_2}}} &= \sqrt{7} \times \sqrt{\frac{3RT_{\text{N}_2}}{M_{\text{N}_2}}} \\
\sqrt{\frac{3T_{\text{H}_2}}{2}} &= \sqrt{7} \times \sqrt{\frac{3T_{\text{N}_2}}{28}} \\
\frac{T_{\text{H}_2}}{2} &= \frac{T_{\text{N}_2}}{4} \Rightarrow T_{\text{H}_2} = \frac{T_{\text{N}_2}}{2}
\end{aligned}
$$

Question3

At what temperature will the RMS velocity of sulphur dioxide molecules at 400 K be the same as the most probable velocity of oxygen molecules?

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Options:

A. 600 K

B. 200 K

C. 400 K

D. 300 K

Answer: D

Solution:

The RMS velocity is given by

$v_{\text{ms}} = \sqrt{\frac{3RT}{M}}$

According to question, v_{mms} of $\text{SO}_2 =$ Most probable speed of O_2

$$\sqrt{\frac{3RT_{SO_2}}{M_{SO_2}}} = \sqrt{\frac{2RT_{O_2}}{M_{O_2}}}$$

$$\sqrt{\frac{3T_{SO_2}}{M_{SO_2}}} = \sqrt{\frac{2T_{O_2}}{M_{O_2}}}$$

Squaring both side and substituting the values,

$\frac{3\times400}{64} = \frac{T_{O_2}\times2}{32} \Rightarrow T_{O_2} = 300 \text{ K}$

Question4

The variation of volume of an ideal gas with its number of moles (n) is obtained as a graph at 300 K and 1 atm pressure. What is the slope of the graph ?

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Options:

A. 24.6 L

B. 24.6 L mol^{-1}

C. $\frac{1}{24.6} L^{-1}$

D. $\frac{1}{24.6} L^{-1} \text{ mol}$

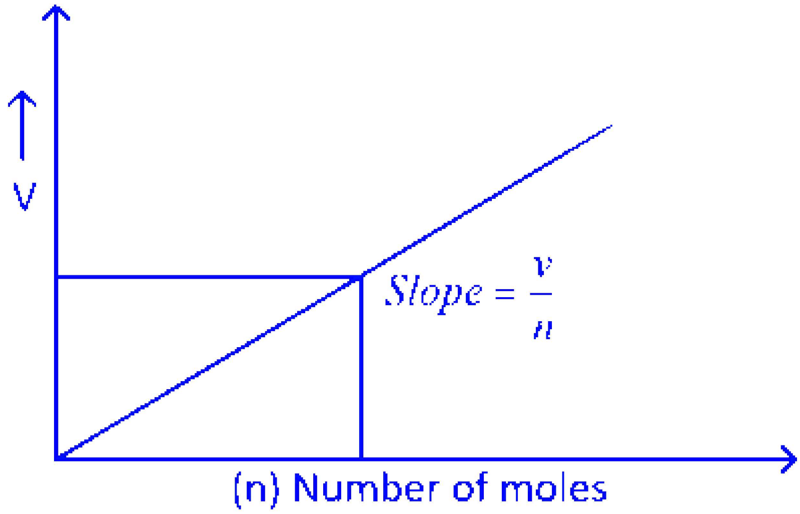
Answer: B

Solution:

From ideal gas equation,

$pV = nRT$

The plot is between volume and number of moles



Slope $= \frac{V}{n}$

From Eq. (i) $\frac{V}{n}$ is $\frac{RT}{p}$

where, $R = 0.0821$ L atm/mol/K

$T = 300$ K, $p = 1$ atm

Slope $= \frac{V}{n} = \frac{0.082 \times 300}{1}$

$\frac{V}{n} = 24.6$ L/mol

Question5

At 400 K , an ideal gas is enclosed in a $0.5 \mathrm{~m}^3$ vessel at pressure of 203 kPa . What is the change in temperature required (in K), if it occupies a volume of $0.2 \mathrm{~m}^3$ under a pressure of 304 kPa ? (Nearest interger)

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Options:

A. 240

B. 160

C. 120

D. 80

Answer: B

Solution:

Given:

Initial temperature, $T_1 = 400\,\mathrm{K}$

Initial volume, $V_1 = 0.5\,\mathrm{m}^3$

Initial pressure, $p_1 = 203\,\mathrm{kPa}$

For the final state:

Final volume, $V_2 = 0.2\,\mathrm{m}^3$

Final pressure, $p_2 = 304\,\mathrm{kPa}$

To find the change in temperature using the ideal gas law, we can use the relation:

$pV = nRT \quad$ which implies $\quad \frac{pV}{T} = \text{constant}$

By comparing the initial and final states, we have:

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

Solving for T_2:

$T_2 = \frac{p_2V_2T_1}{p_1V_1}$

Substituting the given values:

$T_2 = \frac{304\times0.2\times400}{203\times0.5} \approx 240\,\mathrm{K}$

Therefore, the change in temperature is:

Change in temperature $= T_1 - T_2 = 400 - 240 = 160\,\mathrm{K}$

Question6

If the density of a mixture of nitrogen and oxygen gases at 400 K and l atm pressure is $0.920\mathrm{gL}^{-1}$, what is the mole fraction of nitrogen in the mixture?($R = 0.082$ L atm mol^{-1} K^{-1}; assume ideal gas behaviour for oxygen and nitrogen)

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Options:

A. 0.456

B. 0.554

C. 0.432

D. 0.568

Answer: A

Solution:

Given,

$d = 0.920 \text{ g/L}$

$T = 400 \text{ K}, p = 1 \text{ atm}$

$R = 0.082 \text{ L atm/mol/K}$

Using the formula,

$pM = dRT$ or $M = \frac{dRT}{p} \quad \ldots \text{(i)}$

Substituting the value in Eq. (i),

$$M = \frac{0.920 \times 0.082 \times 400}{1}$$

$$\Rightarrow \quad M = 30.176 \quad \ldots \text{(ii)}$$

$$30.176 = \frac{28n_{N_2} + 32n_{O_2}}{n_{N_2} + n_{O_2}}$$

$$30.176n_{N_2} + 30.176n_{O_2} = 28n_{N_2} + 32n_{O_2}$$

$$2.176n_{N_2} = 1.824n_{O_2}$$

$$n_{N_2} = \frac{1.824}{2.176}n_{O_2} \quad \ldots \text{(iii)}$$

$$\chi_{N_2} = \frac{n_{N_2}}{n_{N_2} + n_{O_2}}$$

Now,

$$= \frac{\frac{1.824}{2.176}n_{O_2}}{\frac{1.824}{2.176}n_{O_2} + n_{O_2}} = 0.456$$

Question7

An open vessel containing air was heated from 27°C to 727°C. Some air was expelled. What is the fraction of air remaining in the vessel ? (Assume air as an ideal gas.)

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Options:

A. $1/10$

B. $7/10$

C. $3/10$

D. $9/10$

Answer: B

Solution:

Given,
$T_1 = 27^{\circ}\text{C or } 300\text{ K}, T_2 = 727^{\circ}\text{C or } 1000\text{ K}$

According to ideal gas equation

$pV = nRT$

or $p_1V_1 = n_1RT_1$ and $p_2V_2 = n_2RT_2$

$\frac{n_1}{n_2} = \frac{T_2}{T_1} \quad \ldots \text{(i)}$

or $\quad n_2 = \frac{T_1}{T_2} \times n_1 \quad [n_1 = 1\text{ unit}]$

$n_2 = \frac{300}{1000}$

or $\quad n_2 = \frac{3}{10}$

Fraction remains in vessel,

$1 - n_2 = 1 - \frac{3}{10} = \frac{7}{10}$

Question8

Certain volume of oxygen gas diffuses through a porous pot in 20 seconds. Same volume of another gas, X diffuses in Y seconds as

that of oxygen, then X and Y respectively are

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Options:

A. $H_2, 5$

B. He, 10

C. $CO, 30$

D. $CO_2, 40$

Answer: A

Solution:

According to Graham's law, the rate of diffusion of gases is inversely proportional to the square root of their molar masses. This relationship is represented as:

$$\frac{r_0}{r_X} = \sqrt{\frac{M_X}{M_0}}$$

where:

r_0 and r_X represent the rate of diffusion of oxygen and another gas X, respectively.

M_X and M_0 are the molar masses of gas X and oxygen, respectively.

The rate of diffusion is defined as the volume diffused divided by the time taken, which gives us:

$$\text{Rate of diffusion} = \frac{\text{Volume diffused}}{\text{Time taken}}$$

By applying this to the context of diffusion for gases, we reformulate equation (i) as:

$$\frac{\left(\frac{V_0}{20}\right)}{\left(\frac{V_X}{Y}\right)} = \sqrt{\frac{M_X}{16}}$$

Given that the volumes V_0 (oxygen) and V_X (gas X) are equal, the equation simplifies to:

$$\frac{Y}{20} = \frac{\sqrt{M_X}}{4}$$

Solving for Y and $\sqrt{M_X}$:

$$\frac{Y}{\sqrt{M_X}} = 5$$

Therefore, $Y = 5$ seconds.

$\sqrt{M_X} = 1$, which implies $M_X = 1^2 = 1$.

The molar mass of 1 indicates that the gas X is hydrogen (H_2).

Question9

At 300 K and 760 torr pressure, the density of a mixture of He and O_2 gases is $0.543 gL^{-1}$. The mass percent of oxygen approximately is $\left(R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}\right)$

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Options:

A. 33

B. 80

C. 20

D. 67

Answer: B

Solution:

To find the mass percent of oxygen in a gas mixture at 300 K and 760 torr with a density of 0.543 g/L, we can use the ideal gas law in its density form:

$p \times M = d \times R \times T$

where:

$p =$ pressure in atm (convert 760 torr to 1 atm),

$M =$ molar mass of the gas mixture,

$d =$ density (0.543 g/L),

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

$T =$ temperature (300 K).

First, calculate the molar mass M:

$M = \frac{d \times R \times T}{p} = \frac{0.543 \times 0.0821 \times 300}{1} = 13.37 \text{ g/mol}$

Assume y is the mole fraction of oxygen (O_2) in the mixture and $(1 - y)$ is the mole fraction of helium (He).

The average molecular weight of the mixture is given by:

$M = y \times M_{O_2} + (1 - y) \times M_{He}$

Substitute the known values:

$13.37 = y \times 32 + (1 - y) \times 4$

$13.37 = 32y + 4 - 4y$

$13.37 = 28y + 4$

$9.37 = 28y$

$y = \frac{9.37}{28} \approx 0.335$

The mass percent of oxygen is calculated as follows:

Mass percent of $O_2 = \frac{y \times M_{O_2} \times 100}{(1-y) \times M_{He} + y \times M_{O_2}}$

$= \frac{0.335 \times 32 \times 100}{(1-0.335) \times 4 + 0.335 \times 32}$

Simplifying:

$= \frac{10.72}{10.72 + 2.66} \times 100$

$= \frac{10.72}{13.38} \times 100 \approx 80\%$

Thus, the mass percent of oxygen in the mixture is approximately 80%.
