

- Q1.** Explain the significance of van der Waals parameters.
- Q2.** What would be SI units of a quantity pV^2T^2/n ?
- Q3.** An iron cylinder contains helium gas at a pressure of 250 kPa at 300 K. The cylinder can withstand a pressure of 1×10^6 Pa. The room in which cylinder is placed catches fire. Predict whether the cylinder will blow up before it melts or not. (Melting point of iron = 1800 K).
- Q4.** A gas occupies 60.50 L at 25°C and 1 atm. At what temperature the same gas will occupy 30.25 L at 1.0 atm?
- Q5.** The volume of a given amount of a gas at 100 kPa and 100°C is 10.0 dm³, what will be its volume at 200°C?
- Q6.** A metal tank contains helium gas maintained under a pressure of 2000 kPa and at a temperature of 27°C. If the tank is in a building that catches fire, will the tank blow up (burst) before it melts. The melting point of metal is 1500°C, and the maximum pressure that the metal tank can withstand is 7000 kPa.
Method: The metal tank containing the gas will burst if the pressure exceeds 7000 kPa at a temperature lower than 1500°C. If the temperature becomes 1500°C below a pressure of 7000 kPa then the metal will melt. Thus, we have to calculate the temperature at which the gas pressure is 7000 kPa.
- Q7.** A flask containing air was heated from 27°C to 477°C at a constant atmospheric pressure. Calculate the fraction of the air that would have been expelled out from the open flask.
- Q8.** A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 0.27 L volume; upto what volume can the balloon be expanded by filling H₂?
- Q9.** 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?
- Q10.** In term of Charles's law explain why –273°C is the lowest possible temperature?
- Q11.** 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?
- Q12.** What will be the minimum pressure required to compress 500 dm³ of air at 1 bar to 200 dm³ at 30°C?

S1. van der Waals parameter 'a' is a measure of the magnitude of intermolecular forces. The van der Waals parameter 'b' which is also called *co-volume* is a measure of effective size of the gas molecules.

S2. Putting SI units for different parameters, we get

$$pV^2 T^2/n = \frac{(\text{Nm}^{-2})(\text{m}^3)^2(\text{K})^2}{\text{mol}} = \text{Nm}^4 \text{K}^2 \text{mol}^{-1}.$$

S3. Let us suppose that T_2 is the temperature when the bursting pressure is $P_2 = 1 \times 10^6$ Pa i.e., $P_2 = 1 \times 10^3$ kPa. The relation $P_1/P_2 = T_1/T_2$ gives

$$T_2 = T_1 \frac{P_2}{P_1} = 300 \text{ K} \times \frac{1 \times 10^3 \text{ kPa}}{250 \text{ kPa}} = 1200 \text{ K} < 1800 \text{ K (m. pt)}$$

The temperature for the bursting pressure of iron cylinder is less than the melting point of iron. Therefore, cylinder will blow up before it melts.

S4. Data given:

$$V_1 = 60.50 \text{ L} \qquad V_2 = 30.25 \text{ L}$$

$$T_1 = 25 + 273 = 298 \text{ K} \qquad T_2 = ?$$

$$\therefore T_2 = \frac{T_1 \times V_2}{V_1} = \frac{298 \text{ K} \times 30.25 \text{ L}}{60.50 \text{ L}} = 149 \text{ K}$$

Remark: This temperature corresponds to -124°C

S5. Data given:

$$V_1 = 10.0 \text{ dm}^3 \qquad T_1 = 273 + 100^\circ\text{C} = 373 \text{ K}$$

$$V_2 = ? \qquad T_2 = 273 + 200^\circ\text{C} = 473 \text{ K}$$

Charles' law gives:

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{10.0 \text{ dm}^3 \times 473 \text{ K}}{373 \text{ K}} = 12.7 \text{ dm}^3$$

Comments: The volume of the gas has increased on increasing the temperature from 100°C to 200°C . Though the temperature is doubled, but the volume is not. Why? Because the volume is directly proportional to the absolute temperature and not to the degree celsius.

S6. Since the pressure of the gas increases with increase of temperature as the volume is kept constant, we shall use the relation $P_1/P_2 = T_1/T_2$, that is,

$$T_2 = \frac{T_1 \times P_2}{P_1} = \frac{300 \text{ K} \times 7000 \text{ kPa}}{2000 \text{ kPa}} = 1050 \text{ K}$$

Thus at a temperature of 1050 K (777°C) the maximum pressure that the tank can withstand is reached. Beyond this, if the temperature is increased the tank would burst before it melts.

S7.

$T_1 = 27^\circ\text{C} = 300\text{ K}$	Initial volume of air = V_1
$T_2 = 477^\circ\text{C} = 750\text{ K}$	Final volume of air = V_2

On heating the volume of air increases. Since volume of flask is fixed, the increased volume of air will be expelled.

$$\begin{aligned} \text{Change in volume} &= \text{Volume of air expelled} \\ &= V_2 - V_1 \end{aligned}$$

$$\text{Fraction of air expelled} = (V_2 - V_1) / V_2 = 1 - \frac{V_1}{V_2}$$

According to Charle's law

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

$$\begin{aligned} \text{Fraction of air expelled} &= 1 - \frac{V_1}{V_2} = 1 - \frac{T_1}{T_2} = \frac{T_2 - T_1}{T_2} \\ &= \frac{(750\text{ K} - 300\text{ K})}{750\text{ K}} = \frac{3}{5} = 0.6 \end{aligned}$$

S8. Here,

$P_1 = 1\text{ bar}$	$P_2 = 0.2\text{ bar}$
$V_1 = 0.27\text{ L}$	$V_2 = ?$

From Boyle's law equation:

$$P_1 V_1 = P_2 V_2$$

$$\therefore V_2 = \frac{P_1 V_1}{P_2} = \frac{1 \times 0.27}{0.2} = 1.35\text{ L}$$

Since the balloon bursts at 0.2 bar pressure. Hence, the volume of balloon should remain less than **1.35 L**.

S9. The higher the value of critical temperature of the gas is, the more easy will be its liquefaction which in turn reflects the higher magnitude or stronger intermolecular forces. Since critical temperature of CO_2 (31.1°C) is higher than that of CH_4 (-81.9°C). It means that interparticle forces in CO_2 are stronger than those in methane.

S10. Theoretical considerations based on Charle's law suggest that at -273°C the gases should occupy zero volume. In other words, the gases should cease to exist. If the temperature falls below -273°C , the volume of the gas according to the law should become $-ve$ which is not possible. This indirectly means that -273°C is the lowest possible temperature that can be achieved theoretically.

S11. From the given data:

$P_1 = 1.2\text{ bar}$	$P_2 = ?$
$V_1 = 120\text{ mL}$	$V_2 = 180\text{ mL}$

Temperature remains constant

According to Boyle's law.

$$P_1V_1 = P_2V_2 \quad \text{or} \quad P_2 = \frac{P_1V_1}{V_2}$$

$$P_2 = \frac{1.2(\text{bar}) \times 120(\text{mL})}{180(\text{mL})} = \mathbf{0.8 \text{ bar.}}$$

S12. From the given data:

$$P_1 = 1 \text{ bar}$$

$$P_2 = ?$$

$$V_1 = 500 \text{ dm}^3$$

$$V_2 = 200 \text{ dm}^3$$

Temperature remains constant

According to Boyle's law.

$$P_1V_1 = P_2V_2 \quad \text{or} \quad P_2 = \frac{P_1V_1}{V_2}$$

$$P_2 = \frac{1(\text{bar}) \times 500(\text{dm}^3)}{200(\text{dm}^3)} = \mathbf{2.5 \text{ bar.}}$$

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- Q1. Calculate the total number of electrons in 1.4 g of nitrogen gas.
- Q2. The rate of counting of apples by a high speed machine is 10^{10} apples per second. Calculate the time taken to count one Avogadro number of apples.
- Q3. A given mass of gas occupies a volume of 24 L at 0°C and 760 mm Hg. Calculate its volume at 273°C and 1200 mm Hg.
- Q4. At 40°C and 1 atm a given mass of a gas occupies a volume 3.13 dm^3 . Calculate the pressure when the same amount of gas occupies 0.213 cm^3 at -60°C .
- Q5. A known mass of a gas occupies a volume of 4.1 L at 400 K and 4 bar. Calculate its volume at STP.
- Q6. How will you reduce the volume of a gas to STP value?
- Q7. At 23°C a volume of 17 cm^3 of nitrogen is collected over water when the total pressure is 515 mm Hg. Calculate the volume of dry gas at 273 K and 760 mm Hg. At 23°C the vapour pressure of water is 25 mm Hg.
- Q8. There is an air bubble with a radius of 0.50 cm at the bottom of a lake where the temperature is 7°C and the pressure is 280 kPa. It rises to the surface where the temperature is 27°C and the pressure is 100 kPa. What is the radius of the air bubble when it reaches the surface?
- Q9. A gas 'X' occupies 10 litres at 27°C and 0.7 bar pressure. Calculate the mass of gas 'X' if its relative molecular mass is 60.
- Q10. A plant virus is found to consist of uniform cylindrical particles of 150 Å in diameter and 5000 Å long. The specific volume of the virus is $0.75\text{ cm}^3/\text{g}$. If the virus is considered to be a single particle, find its molar mass.

S1. One N_2 molecule has 14 electrons.

\therefore 1 mol N_2 gas will have 14 mol of electrons.

It is given that: $m = 1.4 \text{ g}$

Molar mass of $N_2 = 28 \text{ g mol}^{-1} = M$

$$n = \frac{m}{M} = \frac{1.4 \text{ g}}{28 \text{ g mol}^{-1}} = 0.05 \text{ mol}$$

$$\begin{aligned} \text{Number of electron} &= \text{Number of electrons in 1 mol } N_2 \times \text{mol of } N_2 \times \text{Avogadro number} \\ &= 14 \times 0.05 \times 6.02 \times 10^{23} \\ &= 4.214 \times 10^{23} \text{ electrons} \end{aligned}$$

S2.

$$\begin{aligned} \text{Time} &= \frac{\text{Number of apples}}{\text{Rate of counting}} = \frac{6.02 \times 10^{23} \text{ apples}}{10^{10} \text{ apples/s}} \\ &= 6.02 \times 10^{13} \text{ s} = \frac{6.02 \times 10^{13}}{60 \times 60 \times 24 \times 365} \text{ year} = 1.909 \times 10^6 \text{ years} \\ &= 1.909 \text{ million years} \end{aligned}$$

S3. The given problem is concerned with the changes in temperature, pressure and volume of a given amount of gas. Therefore, we shall use the combined gas equation $P_1V_1/T_1 = P_2V_2/T_2$.

Initial state: $V_1 = 24 \text{ L},$ $P_1 = 760 \text{ mm Hg},$
 $T_1 = 0^\circ\text{C} + 273 = 273 \text{ K}$

Final state: $V_2 = ?$ $P_2 = 1200 \text{ mm Hg},$
 $T_2 = 273^\circ\text{C} + 273 = 546 \text{ K}$

Combined gas equation $P_1V_1/T_1 = P_2V_2/T_2$ gives

$$V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2} = \frac{760 \text{ mm Hg} \times 24 \text{ L} \times 546 \text{ K}}{273 \text{ K} \times 1200 \text{ mm Hg}} = 30.4 \text{ L}$$

S4. Initial state: $V_1 = 3.13 \text{ dm}^3,$ $P_1 = 1 \text{ atm},$ $T_1 = 40^\circ\text{C} + 273 = 313 \text{ K}$

Final state: $V_2 = 0.213 \text{ cm}^3 = 0.213 \times 10^{-3} \text{ dm}^3,$ $P_2 = ?$
 $T_2 = -60^\circ\text{C} + 273 = 213 \text{ K}$

Combined gas equation $P_1V_1/T_1 = P_2V_2/T_2$ gives

$$P_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{V_2} = \frac{1 \text{ atm} \times 3.13 \text{ dm}^3 \times 213 \text{ K}}{313 \text{ K} \times 0.213 \times 10^{-3} \text{ dm}^3} = 10000 \text{ atm}$$

S5. The combined gas equation gives the volume at STP as:

$$V_0 = \frac{P_2 \times V_2}{T_2} \times \frac{T_0}{P_0} = \frac{4 \text{ bar} \times 4.1 \text{ L}}{400 \text{ K}} \times \frac{273 \text{ K}}{1 \text{ bar}} = 11.2 \text{ L.}$$

S6. Let us suppose that for a given amount of gas

$$T_0 = \text{standard temperature} = 273 \text{ K}$$

$$P_0 = \text{standard pressure} = 1 \text{ bar}$$

$$V_0 = \text{standard volume at STP}$$

The combined gas equation in this situation is written as

$$\frac{P_0 \times V_0}{T_0} = \frac{P_2 \times V_2}{T_2} \quad (\text{For a constant mass})$$

and

$$V_0 = \frac{P_2 \times V_2}{T_2} \times \frac{T_0}{P_0} = \frac{P_2 \times V_2}{T_2} \times \frac{273 \text{ K}}{1 \text{ bar}}$$

S7. Pressure of dry N_2 at 23°C = Total pressure – vapour pressure of water

$$= 540 \text{ mm Hg} - 25 \text{ mm Hg} = 515 \text{ mm Hg}$$

$$P_1 = 515 \text{ mm Hg}, \quad V_1 = 17 \text{ cm}^3, \quad T_1 = 23 + 273 = 296 \text{ K}$$

$$P_2 = 760 \text{ mm Hg}, \quad V_2 = ?, \quad T_2 = 273 \text{ K}$$

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

$$\frac{515 \text{ mm Hg} \times 17 \text{ cm}^3}{296 \text{ K}} = \frac{760 \text{ mm Hg} \times V_2}{273 \text{ K}}$$

$$V_2 = \frac{515 \text{ mm Hg}}{760 \text{ mm Hg}} \times 17 \text{ cm}^3 \times \frac{273 \text{ K}}{296 \text{ K}}$$

$$= 0.677 \times 17 \text{ cm}^3 \times 0.922 = 10.6 \text{ cm}^3.$$

S8. The problem may be reset as

At the bottom of the lake

$$\text{Radius} = r_1 = 0.50 \text{ cm}$$

$$\text{Volume} = V_1 = \frac{4}{3} \pi r_1^3$$

$$\text{Pressure} = P_1 = 280 \text{ kPa}$$

$$T_1 = 7^\circ\text{C} + 273 = 280 \text{ K}$$

At the surface of the lake

$$\text{Radius} = r_2$$

$$\text{Volume} = V_2 = \frac{4}{3} \pi r_2^3$$

$$\text{Pressure} = P_2 = 100 \text{ kPa}$$

$$T_2 = 27^\circ\text{C} + 273 = 300 \text{ K}$$

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

$$\therefore V_2 = \frac{P_1 V_1}{T_1} \times \frac{T_2}{P_2} = \frac{280 \text{ kPa} \times V_1 \times 300 \text{ K}}{280 \text{ K} \times 100 \text{ kPa}} = 3V_1$$

i.e., $\frac{4}{3} \pi r_2^3 = 3 \times \frac{4}{3} \pi r_1^3$

$$\therefore r_2 = (3)^{1/3} r_1 = 1.44 \times 0.50 \text{ cm} \\ = 0.72 \text{ cm.}$$

S9. Let us suppose that W is the mass of gas 'X' and V_0 is its volume at STP. It is given that the relative molecular mass of X is 60, that is, its molar mass is 60 grams.

We know that at STP the molar mass of a gas occupies 22.7 litres. Therefore, mass of one litre of the gas at STP is $60/22.7$, and the mass of V_0 litre of the same gas is $60 \times V_0/22.7$. Therefore, our first aim becomes the calculation of V_0 then W .

Data known: $P_0 = 1 \text{ bar}, T_0 = 273 \text{ K}, V_0 = ?$
 $P_2 = 0.7 \text{ bar}, T_2 = 27^\circ\text{C} + 273 = 300 \text{ K}, V_2 = 10 \text{ L}$

According to combined gas equation:

$$\frac{P_0 V_0}{T_0} = \frac{P_2 V_2}{T_2}$$

$$\frac{1 \text{ bar} \times V_0}{273 \text{ K}} = \frac{0.7 \text{ bar} \times 10 \text{ L}}{300 \text{ K}}$$

$$\therefore V_0 = \frac{0.7 \text{ bar} \times 10 \text{ L}}{300 \text{ K}} \times \frac{273}{1 \text{ bar}} = 6.37 \text{ L}$$

Thus, the mass of gas X = $W = \frac{60 \text{ g} \times 6.37 \text{ L}}{22.7 \text{ L}} = 16.8 \text{ g}$

S10. Molar mass = Mass of one particle \times Avogadro number

$$\therefore M = m \times N_A$$

$$m = \text{Volume of a particle} \times \text{density} = \text{Volume} \times \frac{1}{\text{specific volume}} \\ = \pi r^2 h \times d$$

$$= 3.14 \times \left(\frac{150 \times 10^{-8} \text{ cm}}{2} \right)^2 \times 5000 \times 10^{-8} \text{ cm} \times \frac{1}{0.75 \text{ cm}^3 / \text{g}} \\ = 1.177 \times 10^{-16} \text{ g}$$

$$\therefore M = 1.177 \times 10^{-16} \text{ g} \times 6.022 \times 10^{23} \text{ mol}^{-1} \\ = 7.09 \times 10^7 \text{ g mol}^{-1}.$$

- Q1. Calculate the Volume occupied by 8.8 g of an ideal gas (Molar mass = 44 g/mol) at 31.1°C and at a pressure of 1 bar. Take $R = 0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$.
- Q2. When a ship is sailing in Pacific ocean where temperature is 23.4°C, a balloon is filled with 2.0 L of air. What will be the volume of the balloon when the ship reaches Indian ocean, where temperature is 26.1°C.
- Q3. Using the equation of state $PV = nRT$ show that at a given temperature, the density of the gas is proportional to the gas pressure P .
- Q4. At what temperature 4.0 mol of an ideal gas occupies a volume of 5 dm³ at a pressure of 3.32 bar? $R = 0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$.
- Q5. A sample of 0.050 g of gas X occupies a volume 20.5 cm³ at 27°C and pressure $1.01 \times 10^5 \text{ N m}^{-2}$. Compute the molar mass of X.
- Q6. Compute the pressure exerted by 2.0 mol of an ideal gas enclosed in a vessel of volume 0.1 m³ and maintained at 300 K. Report the result in atm also. Take $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$.
- Q7. A sample of an ideal gas has a volume V_1 at a temperature T_1 and pressure P_1 . The state of the gas is changed to a pressure P_2 so that the temperature becomes T_2 , and volume becomes V_2 .
- (a) Use the ideal gas equation to find the relation between $P_1 V_1 / T_1$ and $P_2 V_2 / T_2$.
- (b) Calculate P_2 if $P_1 = 0.50 \text{ atm}$, $V_1 = 1.00 \text{ L}$, $T_1 = 313 \text{ K}$, $T_2 = 279 \text{ K}$, $V_2 = 0.225 \text{ cm}^3$.
- Q8. The density of a gas (X) at 0°C and 2 bar is the same as the density of nitrogen at 0°C and 5 bar. Calculate the molar mass of gas X when the molar mass of N₂ is 28 g/mol.
- Q9. At 546°C and 0.1 bar pressure a sample of 34.05 mL of an ideal gas weighs 0.0625 g. Calculate the molar mass of the gas. Take $R = 8.314 \text{ Nm mol}^{-1} \text{ K}^{-1}$.
- Q10. What will be the density of a gas at STP when it is known that at 27°C and 2 bar the density of that gas is 5.46 g/dm³.
- Q11. The density of certain gaseous oxide at 1.5 bar pressure at 10°C is same as that of dioxygen at 20°C and 4.5 bar pressure. Calculate the molar mass of gaseous oxide.
- Q12. Calculate the volume of hydrogen gas released when 0.15 g of aluminium present in a drain cleaner reacts with sodium hydroxide (caustic soda) at 20°C and 1 bar.
- Q13. Calculate the payload when a weather balloon of radius 10 m and mass 100 kg is filled with helium at 1.66 bar and at 27°C. It is given that density of air = 1.2 kg/m³ and $R = 0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$.
- Q14. At 95°C a sample of 2.9 g of gas A occupies the same volume as 0.184 g of hydrogen at 17°C at the same pressure. Calculate the molar mass of gas A when the molar mass of H₂ is 2 g/mol.

Q15. Isobutane (C_4H_{10}) undergoes combustion in oxygen according to the reaction:



When 10.00 L of isobutane is burnt at $27^\circ C$ and 1 bar pressure, what volume of CO_2 is produced at $80^\circ C$ and 1.5 bar pressure.

Q16. What mass of potassium chlorate must be decomposed to produce 2.40 L of oxygen at 1.5 bar and $25^\circ C$?

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S1. For an ideal gas $PV = nRT = mRT / M$, we have

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

It is given that:

$$m = 8.8 \text{ g}, M = 44 \text{ g/mol}$$

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

$$T = 31.1 + 273.15 = 304.25 \text{ K}$$

$$V = \frac{mRT}{MP} = \frac{8.8 \text{ g} \times 0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1} \times 304.25 \text{ K}}{44 \text{ g mol}^{-1} \times 1 \text{ bar}} = 5.05 \text{ dm}^3.$$

S2. Here,

$$T_1 = 23.4^\circ\text{C} = 23.4 + 273.5 = 296.55 \text{ K}$$

$$T_2 = 26.1^\circ\text{C} = 26.1 + 273.15 = 299.25 \text{ K}$$

$$V_1 = 2.0 \text{ L}; V_2 = ?$$

Apply Charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad V_2 = \frac{V_1 T_2}{T_1} = \frac{2.0(\text{L}) \times 299.25(\text{K})}{296.55(\text{K})} = 2.018 \text{ L.}$$

S3. According to ideal gas equation:

$$PV = nRT \quad \text{or} \quad P = \frac{nRT}{V}$$

$$n = \frac{\text{Mass of gas (m)}}{\text{Molar mass of gas (M)}}$$

$$P = \frac{mRT}{MV}$$

Now, density (d) = $\frac{m}{V}$

$$\therefore P = \frac{dRT}{M}$$

or $d \propto P$ at constant temperature.

S4. From the ideal gas equation

$$PV = nRT \text{ we get}$$

$$T = \frac{PV}{nR} = \frac{3.32 \text{ bar} \times 5 \text{ dm}^3}{4.0 \text{ mol} \times 0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1}}$$

$$= 50 \text{ K}$$

S5. Data given:

$$m = 0.050 \text{ g} = 0.050 \times 10^{-3} \text{ kg}$$

$$P = 1.01 \times 10^5 \text{ N m}^{-2}$$

$$V = 20.5 \text{ cm}^3 = 20.5 \times 10^{-6} \text{ m}^3$$

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

$$T = 27^\circ\text{C} + 273 = 300 \text{ K}$$

$$\therefore M = \frac{mRT}{PV} = \frac{0.050 \times 10^{-3} \text{ kg} \times 8.314 \text{ N m mol}^{-1} \times 300 \text{ K}}{1.01 \times 10^5 \text{ N m}^{-2} \times 20.5 \times 10^{-6} \text{ m}^3}$$

$$= 6.02 \times 10^{-2} \text{ kg mol}^{-1} = 60.2 \text{ g mol}^{-1}$$

S6. Data given:

$$n = 2.0 \text{ mol}, \quad V = 0.1 \text{ m}^3, \quad T = 300 \text{ K}$$

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

$$P = \frac{nRT}{V} = \frac{2.0 \text{ mol} \times 8.314 \text{ N m mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}}{0.1 \text{ m}^3}$$

$$= 49884 \text{ N m}^{-2} = 49884 \text{ Pa} = 0.499 \text{ bar}$$

Pressure expressed in atm

$$P = \frac{49884 \text{ Pa}}{101325 \text{ Pa atm}^{-1}} = 0.49 \text{ atm.}$$

S7. (a) The gas obeys the equation $PV = nRT$. Thus:

Initial state: $P_1V_1 = nRT_1$ or $P_1V_1/T_1 = nRT$

Final state: $P_2V_2 = nRT_2$ or $P_2V_2/T_2 = nR$

$$\therefore P_1V_1/T_1 = nR = P_2V_2/T_2$$

(b) Data given:

$$P_1 = 0.50 \text{ atm}$$

$$P_2 = ?$$

$$V_1 = 1.00 \text{ L}$$

$$V_2 = 0.225 \text{ cm}^3 = 0.225 \text{ mL} = 0.225 \times 10^{-3} \text{ L}$$

$$T_1 = 313 \text{ K}$$

$$T_2 = 279 \text{ K}$$

The equation $P_1V_1/T_1 = P_2V_2/T_2$ gives

$$P_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{V_2}$$

$$= 1.98 \times 10^3 \text{ atm}$$

S8. From the ideal gas equation $PV = nRT$, we find that density is given by

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

For gas X,

$$d(X) = \frac{M(X)P(X)}{RT(X)}$$

For as N_2 ,
$$d(N_2) = \frac{M(N_2)P(N_2)}{RT(N_2)}$$

It is given that $d(X) = d(N_2)$. Therefore,

$$\frac{M(X)P(X)}{RT(X)} = \frac{M(N_2)P(N_2)}{RT(N_2)}$$

It is also given that, $T(X) = T(N_2) = 273 \text{ K}$

$$P(X) = 2 \text{ bar}, P(N_2) = 5 \text{ bar}$$

$$\therefore M(X) = 28 \text{ g mol}^{-1} \times \frac{5 \text{ bar}}{2 \text{ bar}} = 70 \text{ g/mol}$$

S9. Formula used
$$M = \frac{mRT}{PV}$$

Data given: $m = 0.0625 \text{ g} = 0.0625 \times 10^{-3} \text{ kg}$

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

$$P = 0.1 \text{ bar} = 0.1 \times 10^5 \text{ Nm}^{-2}$$

$$V = 34.05 \text{ mL} = 34.05 \times 10^{-6} \text{ m}^3$$

$$T = 546 + 273 = 819 \text{ K}$$

$$M = \frac{(0.062 \times 10^{-3} \text{ kg})(8.314 \text{ Nm mol}^{-1} \text{ K}^{-1})(819 \text{ K})}{(0.1 \times 10^5 \text{ Nm}^{-2})(34.05 \times 10^{-6} \text{ m}^3)}$$

$$= 0.125 \text{ kg mol}^{-1} = 125 \text{ g mol}^{-1}$$

S10. For an ideal gas, density is given by

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

Under two sets of conditions the densities d_1 and d_2 are expressed as

$$d_1 = \frac{MP_1}{RT_1} \quad \text{and} \quad d_2 = \frac{MP_2}{RT_2}$$

$$\therefore \frac{d_2}{d_1} = \frac{P_2}{T_2} \times \frac{T_1}{P_1}$$

or
$$d_2 = \frac{P_2}{T_2} \times \frac{T_1}{P_1} \times d_1$$

Now, $T_1 = 300 \text{ K}, P_1 = 2 \text{ bar}, d_1 = 5.46 \text{ g/dm}^3$

At STP, $T_2 = 273 \text{ K}, P_2 = 1 \text{ bar}, d_2 = \text{to be calculated}$

Thus,

$$d_2 = \frac{1\text{bar}}{273\text{K}} \times \frac{300\text{K}}{2\text{bar}} \times 5.46\text{ g/dm}^3$$
$$= 3.0\text{ g/dm}^3$$

S11. Density of dioxygen (O₂) at 4.5 bar pressure and 10°C

$$d_{\text{O}_2} = \frac{4.5(\text{bar}) \times 32(\text{g mol}^{-1})}{R \times 293.15(\text{K})}$$

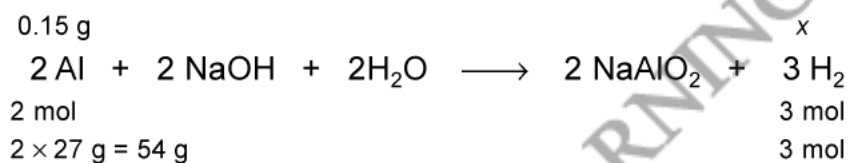
$$d_{\text{oxide}} = \frac{1.5(\text{bar}) \times M(\text{g mol}^{-1})}{R \times 283.15(\text{K})}$$

Now, $d_{\text{O}_2} = d_{\text{oxide}}$

$$\therefore \frac{4.5(\text{bar}) \times 32(\text{g mol}^{-1})}{R \times 293.15(\text{K})} = \frac{1.5(\text{bar}) \times M(\text{g mol}^{-1})}{R \times 283.15(\text{K})}$$

or $M = \frac{4.5 \times 293.15 \times 32}{283.15 \times 1.5} = 99.39\text{ g mol}^{-1}$.

S12. Reaction between aluminium and caustic soda (NaOH)



Ratio proportion: $\frac{0.15\text{ g}}{54\text{ g}} = \frac{x}{3\text{ mol}}$

$$x = \frac{0.15\text{ g}}{54\text{ g}} \times 3\text{ mol} = 8.33 \times 10^{-3}\text{ mol of H}_2$$

Volume of $8.33 \times 10^{-3}\text{ mol H}_2$ is calculated from ideal gas equation $PV = nRT$

Data given:

$$T = 20 + 273 = 293\text{ K}, \quad P = 1\text{ bar}$$

$$R = 0.083\text{ bar dm}^3\text{ mol}^{-1}, \quad n = x = 8.33 \times 10^{-3}\text{ mol}$$

$$V = \frac{nRT}{P} = \frac{8.33 \times 10^{-3}\text{ mol} \times 0.083\text{ bar dm}^3\text{ mol}^{-1}\text{ K}^{-1} \times 293\text{ K}}{1\text{ bar}}$$
$$= 0.2025\text{ dm}^3 = 202.5\text{ cm}^3.$$

S13.

Pay load = Mass of displaced air – mass of the balloon

Given:

Radius of balloon = $r = 10\text{ m}$

Mass of balloon = $m = 100\text{ kg}$

Pressure of He inside balloon = 1.66 bar = P

Temperature = 27°C = 300 K = T

Density of air = $d = 1.2 \text{ kg m}^{-3}$

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

$$\text{Volume of balloon} = \frac{4}{3} \pi r^3 = \frac{4}{3} \times 3.14 \times (10 \text{ m})^3 = 4.1866 \times 10^3 \text{ m}^3$$

Volume of air displaced = Volume of balloon

Mass of air displaced = Volume \times density

$$= 4.1805 \times 10^3 \text{ m}^3 \times 1.2 \text{ kg/m}^3$$

$$= 5.0239 \times 10^3 \text{ kg} = 5023.9 \text{ kg}$$

$$\text{Mass of He} = w = \frac{PVM}{RT}$$

$$= \frac{1.66 \text{ bar} \times 4.186 \times 10^3 \text{ m}^3 \times 4 \text{ g mol}^{-1}}{0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}} = 1116.3 \text{ kg}$$

Total mass of balloon = Mass of balloon + Mass of He

$$= (100.0 + 1116.3) \text{ kg} = 1216.3 \text{ kg}$$

Pay load = (5023.9 – 1216.3) kg

$$= \mathbf{3807.6 \text{ kg.}}$$

S14. Let molar mass of the gas be M.

$$n(\text{g}) = \frac{2.9(\text{g})}{M(\text{g mol}^{-1})}; \quad n(\text{H}_2) = \frac{0.184(\text{g})}{2(\text{g mol}^{-1})} = 0.092$$

Volume of 0.092 moles of H₂

$$= \frac{0.092 \times R \times 290}{P} = \frac{26.68 R}{P}$$

Volume of $\frac{2.9}{M}$ mole of gas

$$= \frac{2.9 \times R \times 368}{MP} = \frac{1067.2 R}{MP}$$

Now, $\frac{1067.2 R}{MP} = \frac{26.68 R}{P}$ or $M = 40 \text{ g mol}^{-1}$.

S15. Step 1: Calculation of volume of CO₂ at 27°C and 1 bar.

From the chemical equation it is clear that

2 L of C₄H₁₀ produce CO₂ = 8 L (Similar condition of T and P)

\therefore 10 L of C₄H₁₀ produce CO₂

$$= \frac{8}{2} \times 10 = 40 \text{ L (at } 27^\circ\text{C, 1 bar)}$$

Step 2: Calculation of volume of CO_2 at 80°C and 1.5 bar.

Here,

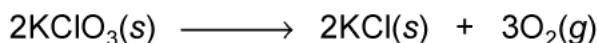
$P_1 = 1 \text{ bar}$	$P_2 = 1.5 \text{ bar}$
$V_1 = 40 \text{ L}$	$V_2 = ?$
$T_1 = 300 \text{ K}$	$T_2 = 80^\circ\text{C}$
	$= 80 + 273 = 353 \text{ K}$

Using the gas equation,

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

$$V_2 = \frac{P_1 V_1 T_2}{T_1 P_2} = \frac{1 \times 40 \times 353}{300 \times 1.5} = \mathbf{31.37 \text{ L.}}$$

S16. The decomposition of potassium chlorate is represented as



Step 1: Calculation of number of moles of oxygen

$$P = 1.5 \text{ bar}$$

$$V = 2.40 \text{ L}$$

$$T = 25^\circ\text{C} = 298 \text{ K}$$

By applying ideal gas equation,

$$n = \frac{PV}{RT} = \frac{1.5(\text{bar}) \times 2.40(\text{L})}{0.083(\text{bar K}^{-1} \text{mol}^{-1}) \times 298(\text{K})} = \mathbf{0.148 \text{ mol.}}$$

Step 2: Calculation of amount of KClO_3

From the chemical equation,

3 mol of oxygen are produced from KClO_3

$$= 2 \text{ moles}$$

0.148 mol of oxygen are produced from KClO_3

$$= \frac{2}{3} \times 0.148 \text{ mol} = 0.0986 \text{ mol}$$

Now, 1 mole of $\text{KClO}_3 = 122.5 \text{ g}$

$$0.0986 \text{ mole of } \text{KClO}_3 = 122.5 \times 0.0986 \text{ g}$$

$$= \mathbf{12.078 \text{ g.}}$$

- Q1. Calculate the pressure exerted by a mixture containing 3.2 g CH₄ and 4.4 g CO₂ in a container of 9 dm³ maintained at 27°C ($R = 8.314 \text{ Nm mol}^{-1} \text{ K}^{-1}$)
- Q2. A mixture of hydrogen and oxygen contains 20% hydrogen by mass. Calculate the partial pressure of hydrogen when the total pressure of the mixture is 1 bar.
- Q3. Calculate the final pressure of the gaseous mixture which is formed by mixing 0.5 L H₂ at 0.8 bar and 2.0 L of O₂ at 0.7 bar in a container of 1 L at 27°C.
- Q4. When 0.01 mole methane and 0.03 mol oxygen are mixed in a vessel maintained at 27°C and 101325 N m^{-2} , the combustion takes place as represented by the equation
- $$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$$
- Calculate (a) the volume of the flask (b) the total pressure of the methane and the partial pressures of various species before and after the completion of the reaction at 27°C. ($R = 8.314 \text{ N mK}^{-1} \text{ mol}^{-1}$)
- Q5. A gaseous mixture is prepared at 100 kPa and 25°C by mixing 0.2 g H₂, 2.8 g N₂ and 0.16 g of O₂. Calculate (a) the mole, (b) mole fraction and (c) the partial pressure of each gas in the mixture.
- Q6. A gaseous mixture is obtained on mixing 8 g oxygen and 4 g hydrogen in a flask of 1 dm³ at 27°C. Calculate the total pressure of the mixture. ($R = 0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$).
- Q7. Equal masses of ethane and hydrogen are mixed in any empty container at 25°C. Calculate the fraction of total pressure exerted by hydrogen gas.
- Q8. At 27°C the pressure of 1 g of an ideal gas A is 2 bar. When 2 g of another ideal gas B is introduced in the same container the pressure becomes 3 bar at the same temperature. Prove that their molecular masses are related by $4 M_A = M_B$.

S1. Given data:

$$V = 9 \text{ dm}^3 = 9 \times 10^{-3} \text{ m}^3$$

$$T = 27^\circ\text{C} + 273 = 300 \text{ K}$$

$$\text{Mass of methane} = m_{\text{CH}_4} = 3.2 \text{ g}$$

$$\text{Molar mass of methane} = 16 \text{ g mol}^{-1}$$

$$\text{Moles of methane} = n_{\text{CH}_4} = \frac{3.2 \text{ g}}{16 \text{ g mol}^{-1}} = 0.2 \text{ mol}$$

$$\text{Mass of carbon dioxide} = m_{\text{CO}_2} = 4.4 \text{ g}$$

$$\text{Molar mass of CO}_2 = 44 \text{ g mol}^{-1}$$

$$\text{Moles of CO}_2 = n_{\text{CO}_2} = \frac{4.4 \text{ g}}{44 \text{ g mol}^{-1}} = 0.1 \text{ mol}$$

For ideal gas mixture

$$P = (n_A + n_B) \frac{RT}{V}$$

$$P = \frac{(0.2 + 0.1) \text{ mol} \times 8.314 \text{ Nm K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{9 \times 10^{-3} \text{ m}^3}$$

$$P = 8.314 \times 10^4 \text{ N m}^{-2} = \mathbf{0.8314 \text{ bar.}}$$

S2. Given:

$$P_{\text{total}} = 1 \text{ bar}$$

Let total mass = 100 g

$$m_{\text{H}_2} 20\% = 20 \text{ g}$$

$$\therefore m_{\text{O}_2} = 80 \text{ g}$$

From the Dalton's law of partial pressure we have

$$P_{\text{H}_2} = x_{\text{H}_2} P_{\text{total}}$$

$$P_{\text{O}_2} = n_{\text{O}_2} P_{\text{total}}$$

$$n_{\text{H}_2} = \frac{20 \text{ g}}{2 \text{ g mol}^{-1}} = 10 \text{ mol}$$

$$n_{\text{O}_2} = \frac{80 \text{ g}}{32 \text{ g mol}^{-1}} = 2.5 \text{ mol}$$

$$x_{\text{H}_2} = \frac{n_{\text{H}_2}}{n_{\text{H}_2} + n_{\text{O}_2}} = \frac{10}{10 + 2.5} = \frac{10}{12.5} = 0.8$$

$$P_{\text{H}_2} = 0.8 \times 1 \text{ bar} \\ = \mathbf{0.8 \text{ bar.}}$$

S3. Given data:

$$\text{Volume of H}_2 = V_{\text{H}_2} = 0.5 \text{ L}$$

$$\text{Pressure of H}_2 = P_{\text{H}_2} = 0.8 \text{ bar}$$

$$\text{Volume of O}_2 = V_{\text{O}_2} = 2.0 \text{ L}$$

$$\text{Pressure of O}_2 = P_{\text{O}_2} = 0.7 \text{ bar}$$

$$\text{Final volume} = \text{Volume of container} = 1 \text{ L} = V_f$$

$$T = 27^\circ\text{C} + 273 = 300 \text{ K}$$

$$\text{Final pressure} = P_f$$

From ideal gas equation $PV = nRT$, we have

$$n_{\text{H}_2} = \frac{P_{\text{H}_2} V_{\text{H}_2}}{RT}, \quad n_{\text{O}_2} = \frac{P_{\text{O}_2} V_{\text{O}_2}}{RT}$$

$$\text{Total moles in the container} = n_{\text{total}} = n_{\text{H}_2} + n_{\text{O}_2} = \frac{(P_{\text{H}_2} V_{\text{H}_2} + P_{\text{O}_2} V_{\text{O}_2})}{RT}$$

Thus, the ideal gas equation gives

$$P_f V_f = n_{\text{total}} RT \\ = \frac{(P_{\text{H}_2} V_{\text{H}_2} + P_{\text{O}_2} V_{\text{O}_2})}{RT} \times RT$$

$$P_f = \frac{P_{\text{H}_2} V_{\text{H}_2} + P_{\text{O}_2} V_{\text{O}_2}}{V} = \frac{0.8 \text{ bar} \times 0.5 \text{ L} + 0.7 \text{ bar} \times 2.0 \text{ L}}{1 \text{ L}}$$

$$P_f = \mathbf{1.8 \text{ bar.}}$$

S4. Given data:

$$T = 27^\circ\text{C} + 273 = 300 \text{ K}, \quad P = 101325 \text{ N m}^{-2}$$

$$\text{Initial total moles} = 0.01 \text{ mol CH}_4 + 0.03 \text{ mol O}_2 = 0.04 \text{ mol}$$

(a) The volume of the flask = Volume of the gas = $V = nRT/P$

$$= \frac{0.04 \text{ mol} \times 8.314 \text{ Nm mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}}{101325 \text{ Nm}^{-2}}$$

$$= 9.84 \times 10^{-4} \text{ m}^3$$

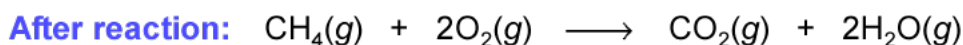
(b) **Calculation of partial pressure before reaction:**

$$\text{Mole fraction of methane} = 0.01/0.04 = 0.25$$

$$\text{Mole fraction of oxygen} = 0.03/0.04 = 0.75$$

$$\text{Partial pressure of methane} = 0.25 \times 101325 \text{ N m}^{-2} = 25331 \text{ N m}^{-2}$$

$$\text{Partial pressure of oxygen} = 0.75 \times 101325 \text{ N m}^{-2} = 75994 \text{ N m}^{-2}$$



The stoichiometry suggests that 0.01 mol CH_4 will combine with 0.02 mol O_2 . That is, after the completion of the reaction no methane will be left, and the amount of unused O_2 will be $0.03 - 0.02 = 0.01$ mol. Also 0.01 mol CO_2 and 0.02 mol H_2O (g) will be formed. Thus:

$$\text{Total moles} = 0.01 \text{ O}_2 + 0.01 \text{ CO}_2 + 0.02 \text{ H}_2\text{O} = 0.04 \text{ mol}$$

Mole fractions	Partial pressures = $x_A P$
$x(\text{CH}_4) = 0$	$p(\text{CH}_4) = 0$
$x(\text{O}_2) = 0.01/0.04 = 0.25$	$p(\text{O}_2) = 25331 \text{ N m}^{-2}$
$x(\text{CO}_2) = 0.01/0.04 = 0.25$	$p(\text{CO}_2) = 25331 \text{ N m}^{-2}$
$x(\text{H}_2\text{O}) = 0.02/0.04 = 0.50$	$p(\text{H}_2\text{O}) = 50663 \text{ N m}^{-2}$

Note: Total pressure before and after the reaction is the same (101325 N m^{-2}). However, if H_2O is in the liquid form, then the total pressure will be low, and will be due to 0.01 mol O_2 and 0.01 mol CO_2 that is, only half of the initial value ($101325/2 = 50662.5 \text{ N m}^{-2}$).

S5. (a)

$$\text{Mole} = \frac{\text{Mass}}{\text{Molar mass}} = n = \frac{m}{M}$$

$$n(\text{H}_2) = 0.2 \text{ g}/2.0 \text{ g mol}^{-1} = 0.1 \text{ mol}$$

$$n(\text{N}_2) = 2.8 \text{ g}/28.0 \text{ g mol}^{-1} = 0.1 \text{ mol}$$

$$n(\text{O}_2) = 0.16 \text{ g}/32.0 \text{ g mol}^{-1} = 0.05 \text{ mol}$$

$$\text{Total mole} = 0.1 + 0.1 + 0.05 = 0.25$$

(b) Mole fraction of A = x_A = mole of A divided by total moles

$$\therefore x(\text{H}_2) = 0.1/0.25 = 0.4, \quad x(\text{N}_2) = 0.1/0.25 = 0.4, \quad x(\text{O}_2) = 0.05/0.25 = 0.2$$

(c) Dalton's law of partial pressure gives $p_A = x_A P$. Thus

$$p(\text{H}_2) = 0.4 \times 100 \text{ kPa} = 40 \text{ kPa}$$

$$p(\text{N}_2) = 0.4 \times 100 \text{ kPa} = 40 \text{ kPa}$$

$$p(\text{O}_2) = 0.2 \times 100 \text{ kPa} = 20 \text{ kPa}$$

S6. **Given data:** Mass of $\text{O}_2 = m_{\text{O}_2} = 8 \text{ g}$

$$\text{Mass of H}_2 = m_{\text{H}_2} = 4 \text{ g}$$

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

$$\text{Temperature} = T = 27^\circ\text{C} + 273 = 300 \text{ K}$$

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

$$\text{Total pressure} = P_{\text{total}}$$

According to Dalton's law of partial pressure

$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{H}_2}$$

$$P_{\text{O}_2} = \frac{n_{\text{O}_2} RT}{V} = \frac{m_{\text{O}_2} RT}{M_{\text{O}_2} V}$$

$$P_{\text{H}_2} = \frac{m_{\text{H}_2} RT}{M_{\text{H}_2} V}$$

$$\therefore P_{\text{total}} = \left(\frac{m_{\text{O}_2}}{M_{\text{O}_2}} + \frac{m_{\text{H}_2}}{M_{\text{H}_2}} \right) \frac{RT}{V}$$

$$M_{\text{O}_2} = 32 \text{ g mol}^{-1}$$

$$M_{\text{H}_2} = 2 \text{ g mol}^{-1}$$

$$P_{\text{total}} = \left(\frac{8 \text{ g}}{32 \text{ mol}^{-1}} + \frac{4 \text{ g}}{2 \text{ mol}^{-1}} \right) \times \frac{0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}}{1 \text{ dm}^3}$$

$$= \left(\frac{1}{4} + 2 \right) \times 24.9 \text{ bar}$$

$$P_{\text{total}} = 56.025 \text{ bar}$$

S7. Let the mass of each gas in the mixture = m

$$\text{Partial pressure of H}_2 = n(\text{H}_2) \frac{RT}{V} = \frac{m RT}{2 V}$$

$$\text{Partial pressure of C}_2\text{H}_6 = n(\text{C}_2\text{H}_6) \frac{RT}{V} = \frac{m RT}{30 V}$$

$$\text{Total pressure} = P_{(\text{total})} = p(\text{H}_2) + p(\text{C}_2\text{H}_6)$$

... Dalton's law

$$= \frac{m(\text{H}_2) RT}{2 V} + \frac{m(\text{C}_2\text{H}_6) RT}{30 V}$$

$$= \frac{mRT}{V} \left(\frac{1}{2} + \frac{1}{30} \right) = \frac{16 mRT}{30 V}$$

Fraction of total pressure exerted by H₂

$$= \frac{p(\text{H}_2)}{P(\text{total})}$$

$$= \frac{m}{2} \frac{RT}{16 mRT / 30V} = \frac{30}{32} = \frac{15}{16}$$

S8. Given data: $P_{\text{initial}} = 1 \text{ bar} = P_A$ $m_A = 1 \text{ g}$
 $T_{\text{initial}} = 27 + 273 = 300 \text{ K}$
 $m_B = 2 \text{ g}, P_{\text{total}} = 3 \text{ bar}$
 $\frac{M_A}{M_B} = ?$ (to be calculated)

Ideal gas equation gives:

$$PV = nRT = \frac{m}{M}RT$$

$$M = \frac{mRT}{PV}$$

$$M_A = \frac{m_A RT}{P_A V} \quad \text{and} \quad M_B = \frac{m_B RT}{P_B V}$$

$$\therefore \frac{M_A}{M_B} = \frac{m_A P_B}{m_B P_A}$$

According to Dalton's law of partial pressure

$$P_{\text{total}} = P_A + P_B = P_{\text{initial}} + P_B$$

$$\therefore P_B = P_{\text{total}} - P_{\text{initial}} = (3 - 1) \text{ bar} = 2 \text{ bar}$$

$$\therefore \frac{M_A}{M_B} = \frac{1 \text{ g} \times 2 \text{ bar}}{2 \text{ g} \times 1 \text{ bar}} = \frac{1}{2}$$

and, $2 M_A = M_B$

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