# CH5 – STATES OF MATTER

**Question 1. What will be the minimum pressure required to compress 500 $dm^3$ of air at 1 bar to 200 $dm^3$ at 30°C?**

**Answer:** $P_1$ = 1 bar, $P_2$ = ? $V_1$ = 500 $dm^3$, $V_2$ = 200 $dm^3$

As temperature remains constant at 30°C,

$P_1V_1 = P_2V_2$

1 bar x 500 $dm^3$ = $P_2$ x 200 $dm^3$ or $P_2$ = 500/200 bar = 2.5 bar

**Question 2. A vessel of 120 mL capacity contains a certain amount of gas at 35°C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 mL at 35°C. What would be its pressure?**

**Answer:** $V_1$ = 120 mL, $P_1$ = 1.2 bar,

$V_2$ = 180 mL, $P_2$ = ?

As temperature remains constant, $P_1V_1 = P_2V_2$

(1.2 bar) (120 mL) = P2 (180mL)

**Question 3. Using the equation of state PV = nRT, show that at a given temperature, density of a gas is proportional to the gas pressure P.**

**Answer:** According to ideal gas equation

PV = nRT or PV=nRT/V

$$n = \frac{\text{Constant Mass of gas}}{\text{Molar mass of gas}}$$

$$P = \frac{mRT}{MV} \qquad \left[\because \rho\,(density) = \frac{m}{V}\right]$$

$$P = \frac{\rho RT}{M}$$

$P \propto \rho$ at constant temperature

**Question 4. At 0°C, the density of a gaseous oxide at 2 bar is same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?**

**Answer:** Using the expression, d = MP/RT , at the same temperature and for same density,

$M_1P_1 = M_2P_2$ (as R is constant)

(Gaseous oxide) ($N_2$)

or

$M_1 \times 2 = 28 \times 5$ (Molecular mass of $N_2 = 28$ u)

or $M_1 = 70$u

Question 5. Pressure of l g of an ideal gas A at 27°C is found to be 2 bar. When 2 g of another ideal gas B is introduced in the same flask at same temperature, the pressure becomes 3 bar. Find the relationship between their molecular masses.

Answer: Suppose molecular masses of A and B are $M_A$ and $M_B$ respectively. Then their number of moles will be

$$n_A = \frac{1}{M_A}, \qquad n_B = \frac{2}{M_B}$$

$P_A = 2$ bar, $\quad P_A + P_B = 3$ bar, $\quad$ *i.e.*, $P_B = 1$ bar

Applying the relation $PV = nRT$

$$P_AV = n_ART, \qquad P_BV = n_BRT \quad \therefore \quad \frac{P_A}{P_B} = \frac{n_A}{n_B} = \frac{1/M_A}{2/M_B} = \frac{M_B}{2M_A}$$

$$\text{or} \qquad \frac{M_B}{M_A} = 2 \times \frac{P_A}{P_B} = 2 \times \frac{2}{1} = 4 \quad \text{or} \quad \boldsymbol{M_B = 4\,M_A}.$$

Question 6. The drain cleaner, Drainex contains small bits of aluminium which react with caustic soda to produce dihydrogen. What volume of dihydrogen at 20 °C and one bar will be released when 0.15g of aluminium reacts?

Answer: The chemical equation for the reaction is

$2\,Al + 2\,NaOH + H_2O \rightarrow 2\,NaAlO_2 + 3H_2$ (3 x 22400 mL At N.T.P)

2 x 27 = 54 g.

54 g of Al at N.T.P release

$H_2$ gas = 3 x 22400 0.15 g of Al at N.T.P release

$$H_2 \text{ gas} = \frac{3 \times 22400 \times 0.15}{54} = 186.7 \text{ mL}$$

N.T.P condition.

$V_1 = 186.7$ mL $\qquad V_2 = ?$

$P_1 = 1.013$ bar $\qquad P_2 = 1$ bar

$T_1 = 0 + 273 = 273$ K $\qquad T_2 = 20 + 273 = 293$ K

According to Gas equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{or} \quad V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

$$V_2 = \frac{1.013 \text{ bar} \times 186.7 \text{ mL} \times 293 \text{K}}{1 \text{ bar} \times 273 \text{K}}$$

$$= \mathbf{203 \text{ mL}}$$

Question 7. What will be the pressure exerted by a mixture of 3.2g of methane and 4.4g of carbon dioxide contained in a 9 $dm^3$ flask at 27 °C?

Answer:

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

$$p_{CH_4} = \left(\frac{3.2}{16}\text{mol}\right)\frac{0.0821 \text{ dm}^3\text{atm K}^{-1}\text{mol}^{-1} \times 300 \text{ K}}{9 \text{ dm}^3} = 0.55 \text{ atm}$$

$$p_{CO_2} = \left(\frac{4.4}{44}\text{mol}\right)\frac{0.0821 \text{ dm}^3\text{atm K}^{-1}\text{mol}^{-1} \times 300 \text{ K}}{9 \text{ dm}^3} = 0.27 \text{ atm}$$

$$p_{total} = 0.55 + 0.27 = \mathbf{0.82 \text{ atm}}$$

In terms of SI units, $R = 8.314\ p\ m^3\ K^{-1}\ mol^{-1}$, $V = 9 \times 10^{-3}\ m^3$

$$P = 5.543 \times 10^4 \text{ Pa} + 2.771 \times 10^4 \text{ Pa} = \mathbf{8.314 \times 10^4 \text{ Pa.}}$$

Question 8. What will be the pressure of the gas mixture when 0.5 L of $H_2$ at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in all vessel at 27 °C?

Answer: Calculation of partial pressure of $H_2$ in 1L vessel $P_1$= 0.8 bar, $P_2$= ? $V_1$= 0.5 L , $V_2$ = 1.0 L

As temperature remains constant, $P_1V_1 = P_2V_2$

(0.8 bar) (0.5 L) = $P_2$ (1.0 L) or $P_2$ = 0.40 bar, i.e., $PH_2$ = 0.40 bar

Calculation of partial pressure of 02 in 1 L vessel

$P_1' V_1 = P_2'V_2'$

(0.7 bar) (2.0 L) = $P_2$ (1L) or $P_2'$ = 1.4 bar, i.e.,$Po_2$= 1.4 bar

Total pressure =$P_{Hz}$ + $P_{Q2}$ = 0.4 bar + 1.4 bar = 1.8 bar

Question 9. Density of a gas is found to be 5.46 g/$dm^3$ at 27 °C and at 2 bar pressure. What will be its density at STP?

Answer:

$d = \frac{MP}{RT}$. For the same gas at different temperatures and pressures, $\frac{d_1}{d_2} = \frac{P_1}{T_1} \times \frac{T_2}{P_2}$.

Here, $d_1 = 5.46 \text{ g dm}^{-3}$, $T_1 = 27\,°C = 300$ K, $P_1 = 2$ bar.
At STP, $d_2 = ?$, $T_2 = 0\,°C = 273$ K, $P_2 = 1$ bar

$$\therefore \quad \frac{5.46 \text{g dm}^{-3}}{d_2} = \frac{2 \text{ bar}}{300 \text{ K}} \times \frac{273 \text{ K}}{1 \text{ bar}} \text{ or } d_2 = \mathbf{3 \text{ g dm}^{-3}}$$

Question 10. 34.05 mL of phosphorus vapour weighs 0.0625 g at 546°C and 1.0 bar pressure. What is the molar mass of phosphorus?

Answer:

Step I. Calculation of volume at 0°C and 1 bar pressure

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \text{ i.e., } \frac{1 \times 34.05}{546 + 273} = \frac{1 \times V_2}{273} \text{ or } V_2 = 11.35 \text{ mL}$$

11.35 mL of vapour at 0°C and 1 bar pressure weigh = 0.0625 g

$\therefore$ 22700 mL of vapour at 0°C and 1 bar pressure will weigh

$$= \frac{0.0625}{11.35} \times 22700 = 125 \text{ g}$$

$\therefore$ Molar mass = **125 g mol$^{-1}$**

Alternatively, using

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

$$PV = nRT, \text{ i.e., } n = \frac{PV}{RT} - \frac{1.0 \text{ bar} \times (34.05 \times 10^{-3} \text{ dm}^3)}{0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ Mol}^{-1} \times 819 \text{ K}}$$

$$= 5 \times 10^{-4} \text{ mol}$$

$$\therefore \quad \text{Mass of 1 mole} = \frac{0.0625}{5 \times 10^{-4}} g = 125 \text{ g}$$

$\therefore$ Molar mass = **125 g mol$^{-1}$**

Question 11. A student forgot to add the reaction mixture to the round bottomed flask at 27 °C but instead, he/she placed the flask on the flame. After a lapse of time, he realized his mistake, and using a pyrometer, he found the temperature of the flask was 477 °C. What fraction of air would have been expelled out?

Answer:

Suppose volume of vessel = $V$ $cm^3$

*i.e.,* volume of air in the flask at 27°C = $V$ $cm^3$.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}, \quad i.e., \quad \frac{V}{300} = \frac{V_2}{750} \quad \text{or} \quad V_2 = 2.5\ V$$

$\therefore$ Volume expelled = $2.5\ V - V = 1.5\ V$

$\therefore$ Fraction of air expelled = $\frac{1.5\ V}{2.5\ V} = \frac{3}{5}$

Question 12.Calculate the temperature of 4.0 moles of a gas occupying 5 $dm^3$ at 3.32 bar (R = 0.083 bar $dm^3$ $K^{-1}$ $mol^{-1}$)

Answer:

$$PV = nRT \quad \text{or} \quad T = \frac{PV}{nR} = \frac{3.32\ \text{bar} \times 5\ dm^3}{4.0\ \text{mol} \times 0.083\ \text{bar}\ dm^3\ K^{-1} mol^{-1}} = \mathbf{50K}$$

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

Answer: Molecular mass of $N_2$ = 28g

28 g of $N_2$ has No. of molecules = 6.022 x $10^{23}$ 1.4 g of

$N_2$ has No. of molecules = 6.022 x $10^{23}$ x 1.4 g/28 g

= 3.011 x $10^{22}$ molecules.

Atomic No. of Nitrogen (N) = 7

1 molecule of $N_2$ has electrons = 7 x 2 = 14

3.011 x $10^{22}$ molecules of $N_2$ have electrons

= 14 x 3.011 x $10^{22}$

= 4.215 x $10^{23}$ electrons.

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

Answer:

Time taken to distribute $10^{10}$grains = 1s

Time taken to distribute = $6.022 \times 10^{23}$ grains

$$= \frac{1s \times 6.022 \times 10^{23}\ \text{grains}}{10^{10}\ \text{grains}}$$

$$= \frac{6.022 \times 10^{13}}{60 \times 60 \times 24 \times 365} = \mathbf{1.9 \times 10^6\ yr.}$$

Question 15. Calculate the total pressure in a mixture of 8g of oxygen and 4g of hydrogen confined in a vessel of l $dm^3$ at 27°C. R = 0.083 bar $dm^3$ $K^{-1}$ $mol^{-1}$.

Answer:

$$\text{Molar mass of } O_2 = 32 \text{ g mol}^{-1} \quad \therefore \quad 8 \text{ g } O_2 = \frac{8}{32} \text{ mol} = 0.25 \text{ mol}$$

$$\text{Molar mass of } H_2 = 2 \text{ g mol}^{-1} \quad \therefore \quad 4 \text{ g } H_2 = \frac{4}{2} = 2 \text{ mol}$$

$\therefore$ Total number of moles ($n$) = 2 + 0.25 = 2.25

$$V = 1 \text{ dm}^3, \quad T = 27°C = 300 \text{ K}, R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$$

$$PV = nRT$$

or

$$P = \frac{nRT}{V} = \frac{(2.25 \text{ mol})(0.083 \text{ bar dm}^3\text{K}^{-1}\text{mol}^{-1})(300 \text{ K})}{1 \text{ dm}^3}$$

$$= \mathbf{56.025 \text{ bar}}$$

Question 16. Pay load is defined as the difference between the mass of the displaced air and the mass of the balloon. Calculate the pay load when a balloon of radius 10 m, mass 100 kg is filled with helium at 1.66 bar at 27°C (Density of air = 1.2 kg $m^{-3}$ and R = 0.083 bar $dm^3$ $K^{-1}$ $mol^{-1}$).

Answer:

Radius of the balloon = 10 m

$$\therefore \text{ Volume of the balloon } = \frac{4}{3}\pi r^3 = \frac{4}{3} \times \frac{22}{7} \times (10 \text{ m})^3 = 4190.5 \text{ m}^3$$

Volume of He filled at 1.66 bar and 27 °C = 4190.5 $m^3$

Calculation of mass of He

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

or

$$w = \frac{MPV}{RT} = \frac{(4 \times 10^{-3} \text{ kg mol}^{-1})(1.66 \text{ bar})(4190.5 \times 10^3 \text{ dm}^3)}{(0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1})(300 \text{ K})}$$

$$= 1117.5 \text{ kg}$$

Total mass of the balloon alongwith He = 100 + 1117.5 = 1217.5 kg

Maximum mass of the air that can be displaced by balloon to go up = Volume × Density

$$= 4190.5 \text{ m}^3 \times 1.2 \text{ kg m}^{-3} = 5028.6 \text{ kg}$$

$$\therefore \quad \text{Pay load} = 5028.6 - 1217.5 \text{ kg} = \mathbf{3811.1 \text{ kg}}$$

Question 17. Calculate the volume occupied by 8.8 g of $CO_2$ at 31.1 °C and 1 bar pressure. R = 0.083 bar $LK^{-1}$ $mol^{-1}$

Answer:

$$\text{No. of moles of } CO_2\ (n) = \frac{\text{Mass of } CO_2}{\text{Molar mass}} = \frac{8.8\ g}{44 g\ mol^{-1}} = 0.2\ mol$$

$$\text{Pressure of } CO_2\ (P) = 1\ bar$$
$$R = 0.083\ bar\ LK^{-1}\ mol^{-1}$$
$$\text{Temperature } (T) = 273 + 31.1 = 304.1\ K$$
$$\text{Since from gas eq. } PV = nRT$$
$$V = \frac{nRT}{P} = \frac{0.2 \times 0.083 \times 304.1}{1\ bar}$$
$$= \mathbf{5.048\ L}$$

Question 18. 2.9 g of a gas at 95°C occupied the same volume as 0.184 g of hydrogen at 17°C at the same pressure. What is the molar mass of the gas ?

Answer:

As $P_1 = P_2$ and $V_1 = V_2$,

$\therefore$ $P_1V_2 = P_2V_2$, *i.e.*, $n_1RT_1 = n_2\ RT_2$ $\therefore$ $n_1T_1 = n_2T_2$

or $\dfrac{w_1}{M_1}T_1 = \dfrac{w_2}{M_2}T_2$

$$\frac{2.9}{M_x} \times (95 + 273) = \frac{0.184}{2} \times (17 + 273) \quad \text{or} \quad M_x = \frac{2.9 \times 368 \times 2}{0.184 \times 290} = \mathbf{40\ g\ mol^{-1}}$$

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

Answer: As the mixture $H_2$ and $O_2$ contains 20% by weight of dihydrogen, therefore, if $H_2$ = 20g, then $O_2$ = 80g

$$n_{H_2} = \frac{20}{2} = 10 \text{ moles}, \quad n_{O_2} = \frac{80}{32} = 2.5 \text{ moles}$$

$$p_{H_2} = \frac{n_{H_2}}{n_{H_2} + n_{O_2}} \times P_{total} = \frac{10}{10 + 2.5} \times 1\ bar = \mathbf{0.8\ bar}$$

Question 20. What would be the SI unit for the quantity $PV^2T^2/n$?

Answer:

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

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

Answer: At -273°C, volume of the gas becomes equal to zero, i.e., the gas ceases to exist.

Question 22. Critical temperature for $CO_2$ and $CH_4$ are 31.1°C and -81.9°C respectively. Which of these has stronger intermolecular forces and why?

Answer: Higher the critical temperature, more easily the gas can be liquefied, i.e., greater are the intermolecular forces of attraction. Hence, $Co_2$ has stronger intermolecular forces than $CH_4$.

Question 23. Explain the physical significance of vander Waals parameters.

Answer: 'a' is a pleasure of the magnitude of the intermolecular forces of attraction, while b is a measure of the effective size of the gas molecules.

MORE QUESTIONS SOLVED

I. Very Short Answer Type Questions

Question 1. What is the value of the gas constant in SI units?

Answer: 8.314 $JK^{-1}$ $mol^{-1}$.

Question 2. Define boiling point of a liquid.

Answer: The temperature at which the vapour pressure of a liquid is equal to external pressure is called boiling point of liquid.

Question 3. What is SI unit of (i) Viscosity (ii) Surface tension?

Answer: (i) Unit of viscosity is $Nsm^{-2}$

(ii) Unit of surface tension is $Nm^{-1}$

Question 4. What is the effect of temperature on (i) surface tension and (ii) Viscosity?

Answer: (i) Surface tension decreases with increase of temperature.

(ii) Viscosity decreases with increase of temperature.

Question 5. What is the unit of coefficient of viscosity?

Ans. Poise.

Question 6. What do you understand by laminar flow of a liquid?

Answer: The type of flow in which there is regular gradation of velocity in passing from one layer to the next is called laminar flow.

Question 7. What do you mean by compressibility factor?

Answer: The deviation from ideal behaviour can be measured in terms of compressibility factor Z.

$Z = PV/nRT$

Question 8. What is Boyle Temperature?

Answer: The temperature at which a real gas obeys ideal gas law over an appreciable range of pressure, is called Boyle temperature or Boyle point.

Question 9. What is meant by elastic collision ?

Answer: Collision in which there is no loss of kinetic energy but there is transfer of energy, is called elastic collision.

Question 10. Define critical temperature of gas.

Answer: The temperature above which a gas cannot be liquefied.

Question 11. What are real gases ?

Answer: A gas which can deviate from ideal gas behaviour at higher pressure and lower temperature, is called a real gas.

Question 12. Define an ideal gas.

Answer: A gas that follows Boyle's law, Charles' law and Avogadro law strictly, is called an ideal gas.



Question 13. Name four properties of gases.

Answer:

- Gases, have no definite shape and no definite volume.
- There is no force of attraction existing between the molecules of gases.
- Gases are highly compressible.
- Gases'can mix evenly and can spread in whole space.

Question 14. State Dalton's law of partial pressure.

Answer: Daltons' Law states that, total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of individual gases.

Question 15. What do you mean by aqueous tension?

Answer: Pressure exerted by saturated water vapour is called aqueous tension.

Question 16. Give mathematical expression for ideal gas equation.

Answer: PV = nRT

Where R is called Gas constant.

Question 17. Write van der Waals equation for n moles of a gas.

Answer:

$$\left[P + \frac{an^2}{V^2}\right](V - nb) = nRT$$

Where 'a' and 'V are van der waals constants.

Question 18. How is compressibility factor expressed in terms of molar volume of the real gas and that of the ideal gas?

Answer:

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

Question 19. Why liquids diffuse slowly as compared to gases?

Answer: In liquids, the molecules are more compact in comparison to gases.

Question 20. What is the effect of temperatures on the vapour pressure of a liquid?

Answer: Vapour pressure increases with rise in temperature.

Question 21. Why falling liquid drops are spherical?

Answer: Because of the property of surface tension, liquid tends to minimise its area.

II. Short Answer Type Questions

Question 1. A weather balloon has a volume of 175 $dm^3$ when filled with hydrogen gas at a pressure of 1.0 bar. Calculate the volume of the balloon when it rises to a height where the atmospheric pressure is 0.8 bar. Assume that temperature is constant.

Answer:

From the given data: $P_1 = 1$ bar, $P_2 = 0.8$ bar
$V_1 = 175\ dm^3$, $V_2 = ?$

Since the temperature is constant, Boyle's Law can be applied

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

or

$$V_2 = \frac{(1\ \text{bar}) \times (175\ \text{dm}^3)}{(0.8\ \text{bar})} = \mathbf{218.75\ dm^3}$$

Question 2. A certain amount of a gas at 27°C and 1 bar pressure occupies a volume of 25 $m^3$. If the pressure is kept constant and the temperature is raised to 77°C, what will be the volume of the gas?

Answer: From the available data: $V_1 = 25\ m^3$, $T_1 = 27 + 273 = 300$ K
$V_2 = ?$ $T_2 = 77 + 273 = 350$ K

Since the pressure of the gas is constant, Charles' law is applicable

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} \quad \text{or} \quad V_2 = \frac{V_1 \times T_2}{T_1}$$

$$V_2 = \frac{(25\ \text{m}^3) \times (350)\ \text{K}}{(300\ \text{K})} = \mathbf{29.17\ m^3.}$$

Question 3. A flask was heated from 27°C to 227°C at constant pressure. Calculate the volume of the flask if 0.1 $dm^3$ of air measured at 227°C was expelled from the flask.

Answer: Let the volume of the flask = V $dm^3$ (after expelling the air)
$V_1 = V\ dm^3$, $T_1 = 27 + 273 = 300K$
$VT_2 = (V + 0.1)\ dm^3$, $T_2 = 227 + 273 = 500$ K

Since the pressure of the gas is constant, Charles' law is applicable.

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} \quad \text{or} \quad V_1 = \frac{T_1}{T_2} \times V_2$$

or

$$V = \frac{(300\ \text{K}) \times (V + 0.1\ \text{dm}^3)}{(500\ \text{K})} \quad \text{or} \quad V = \frac{3\,(V + 0.1)}{5}\ \text{dm}^3$$

or

$$5V - 3V = 0.3\ \text{dm}^3 \quad \text{or} \quad 2V = 0.3\ \text{dm}^3$$

$\therefore$

$$V = \frac{0.3\ \text{dm}^3}{2} = \mathbf{0.15\ dm^3}$$

Question 4. A gas occupying a volume of 100 litres is at 20°C under a pressure of 2 bar. What temperature will it have when it is placed in an evacuated chamber of volume 175 litres? The pressure of the gas in the chamber is one-third of its initial pressure.

Answer:

From the available data: $V_1 = 100$ L, $V_2 = 175$ L

$P_1 = 2$ bar, $P_2 = 2 \times 1/3 = 2/3$ bar

$T_1 = 20 + 273 = 293$ K, $T_2 = ?$

According to Gas equation, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ or $T_2 = \frac{P_2V_2T_1}{P_1V_1}$

By substituting the values, $T_2 = \frac{(2/3 \text{ bar}) \times (175 \text{ L}) \times (293 \text{ K})}{(2 \text{ bar}) \times (100 \text{ L})}$

$= 170.9 \text{ K} = 170.9 - 273.0 = \mathbf{-102.1°C}$

Question 5. At 25°C and 760 mm ofHg pressure a gas occupies 600 mL volume. What will be its pressure at a height where temperature is 10°C and volume of the gas is 640 mL?

Answer:

$P_1 = 760$ mm Hg, $V_1 = 600$ mL

$T_1 = 25 + 273 = 298$ K

$V_2 = 640$ mL and $T_2 = 10 + 273 = 283$ K

According to combined gas law,

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

$$\Rightarrow \quad p_2 = \frac{p_1V_1T_2}{T_1V_2}$$

$$\Rightarrow \quad p_2 = \frac{(760 \text{ mm Hg}) \times (600 \text{ mL}) \times (283 \text{ K})}{(640 \text{ mL}) \times (298 \text{ K})}$$

$$= \mathbf{676.6 \text{ mm Hg}}$$

Question 6. A 34.0 $dm^3$ cylinder contains 212 g of oxygen gas at 21°C. What mass of oxygen must be released to reduce the pressure in the cylinder to 1.24 bar?

Answer:

**Step I.** Calculation of no. of moles of $O_2$ left in cylinder.

$P = 1.24$ bar, $V = 34\ dm^3$

$T = (21 + 273) = 294$ K, $R = 0.083\ dm^3\ bar\ K^{-1}\ mol^{-1}$

According to ideal gas equation,

$$PV = nRT$$

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

$$n = \frac{(1.24\ bar) \times (34\ dm^3)}{(0.083\ dm^3 bar\ K^{-1}\ mol^{-1}) \times (294\ K)}$$

$$= \mathbf{1.727\ mol}$$

**Step II.** Calculation of mass of oxygen released.

Mass of $O_2$ left in the cylinder $= n \times M = (1.727\ mol) \times (32 g\ mol^{-1}) = 55.26$ g

Mass of $O_2$ initially present $= 212$ g

Mass of $O_2$ released $= (212 - 55.26) =$ **156.74 g**

Question 7. The values of the van der Waal's constants for a gas are a = 4.10 $dm^6$ bar $mol^{-2}$ and b = 0.035 $dm^3$ bar $mol^{-1}$. Calculate the values of the critical temperature and critical pressure for the gas.

Answer:

(*i*) Calculation of critical temperature ($T_c$).

$a = 4.10\ dm^6\ bar\ mol^{-2}$

$b = 0.035\ dm^3\ mol^{-1}$

$R = 0.0821\ dm^3\ bar\ mol^{-1}\ K^{-1}$.

Now, critical temperature, $T_c = \frac{8a}{27Rb}$

Substituting the values, $T_c = \frac{8 \times 4.10}{27 \times 0.0821 \times 0.035}$

(*ii*) Calculation of critical pressure ($P_c$)

$$P_c = \frac{a}{27b^2} = \frac{4.10}{27 \times 0.035 \times 0.035} = \mathbf{123.96\ bar}$$

Question 8. The pressure of a mixture of $H_2$ and $N_2$ in a container is 1200 torr. The partial pressure of nitrogen in the mixture is 300 torr. What is the ratio of $H_2$ and $N_2$ molecules in the mixture?

Answer:

Total pressure of mixture = 1200 torr

Partial pressure of $N_2$ ($P_{N_2}$) = 300 torr

Partial pressure of $H_2$ ($P_{H_2}$) = 1200 – 300 = 900 torr

According to ideal gas equation,

$$PV = nRT$$

$$P_{H_2} = \frac{n_{H_2}RT}{V} = 900 \text{ torr} \quad ...(i)$$

$$P_{N_2} = \frac{n_{N_2}RT}{V} = 300 \text{ torr} \quad ...(ii)$$

Divide (*i*) by (*ii*), $\frac{P_{H_2}}{P_{N_2}} = \frac{n_{H_2}}{n_{N_2}} = \frac{900}{300} = \frac{3}{1} \therefore n_{H_2} : n_{N_2} :: 3:1$

Question 9. (a) What do you mean by'Surface Tension'of a liquid?
(b) Explain the factors which can affect the surface tension of a liquid.

Answer: (a) Surface tension: It is defined as the force acting per unit length perpendicular to the line drawn on the surface. It's unit is Nnr1.

(b) Surface tension of a liquid depends upon following factors.

(i) Temperature: Surface tension decreases with rise in temperature. As the temperature of the liquid increases, the average kinetic energy of the molecules increases. Thus, there is a decrease in intermolecular force of attraction which decreases the surface tension.

(ii) Nature of the liquid: Greater the magnitude of intermolecular forces of attraction in the liquid, greater will be the value of surface tension.

Question 10. A neon-dioxygen mixture contains 70.6 g dioxygen and 167.5g neon. If pressure of the mixture of gases in the cylinder is 25 bar. What is the partial pressure of dioxygen and neon in the mixture?

Answer:

$$\text{Number of moles of dioxygen} = \frac{70.6\text{ g}}{32\text{g mol}^{-1}} = 2.21\text{ mol}$$

$$\text{Number of moles of neon} = \frac{167.5\text{ g}}{20\text{ g mol}^{-1}} = 8.375\text{ mol}$$

$$\text{Mole fraction of dioxygen} = \frac{2.21}{2.21+8.375} = \frac{2.21}{10.585} = 0.21$$

$$\text{Mole fraction of neon} = \frac{8.375}{2.21+8.375} = 0.79$$

Alternatively,

mole fraction of neon $= 1 - 0.21 = 0.79$

Partial pressure of a gas = mole fraction × total pressure

$\Rightarrow$ Partial pressure of oxygen $= 0.21 \times (25 \text{ bar})$ $=$ **5.25 bar**

Partial Pressure of neon $= 0.79 \times (25 \text{ bar})$ $=$ **19.75 bar**

III. Long Answer Type Questions

Question 1. State and explain Boyle's law. Represent the law graphically.

Answer: It states that, the pressure of a fixed mass of a gas is inversely proportional to its ' volume if temperature is kept constant.

$$P \propto \frac{1}{V}$$
$$PV = \text{constant } (n \text{ and } T \text{ are constant})$$
$$P_1V_1 = P_2V_2.$$

**Graphical representation:**

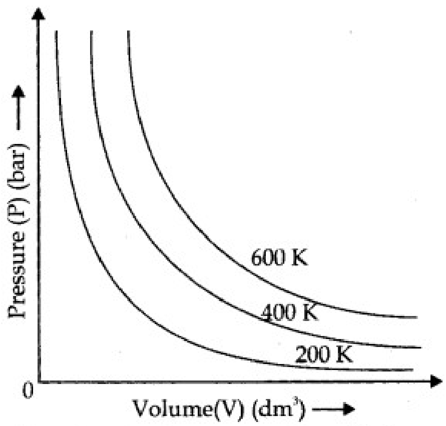


**Fig.** *Graph of pressure, P vs. Volume, V of a gas at different temperatures.*

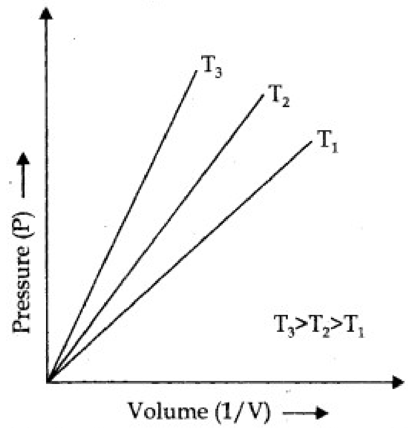


**Fig.** *Graph of pressure of a gas, P vs. 1/V*

Question 2. Give an expression for the van der Wools equation. Give the significance of the constants used in the equation. What are their units?
Answer:

$$\left(p + \frac{n^2a}{V^2}\right)(V - nb) = nRT$$

When n is the no. of moles present and 'a' and 'V are known as van der Waals constants.
Significance of van der Waals constants
van der Waals constant 'a' : 'a' is related to the magnitude of the attractive forces among the molecules of a particular gas. Greater the value of V, more will be the attractive forces.
Unit of 'a' = $L^2\ mol^{-2}$
van der Waals Constant 'b': 'b' determines the volume occupied by the gas molecules which depends upon size of molecule.
Unit of 'b' = $L\ mol^{-1}$.
Question 3. What are ideal and real gases? Out of $CO_2$ and $NH_3$ gases, which is expected to show more deviation from the ideal gas behaviour?
Answer: Ideal Gas: A gas that follows Boyle's law, Charles' law and Avogadro

law strictly, is called an ideal gas. It is assumed that intermolecular forces are not present between the molecules of an ideal gas.
Real Gases: Gases which deviate from ideal gas behaviour are known as real gases. $NH_3$ is expected to show more deviation. Since $NH_3$ is polar in nature and it can be liquified easily.
Question 4. State and explain Dalton's law of partial pressures. Can we apply Dalton's law of partial pressures to a mixture of carbon monoxide and oxygen?
Answer: Dalton's law of partial pressure: When two or more non-reacting gases are enclosed in a vessel, the total pressure of the gaseous mixture is equal to the sum of the partial pressures that each gas will exert when enclosed separately in the same vessel at constant temperature.
$P = P_1 + P_2 + P_3$
Where, P is the total pressure of the three gases A, B, and C enclosed in a container. $P_1$ , $P_2$ and $P_3$ are the partial pressures of the three gases when enclosed separately in the same vessel at a given temperature one by one.
No, the law cannot be applied. Carbon monoxide and oxygen readily combine to form carbon dioxide. The law can be applied only to the non-reacting gases.
IV. Multiple Choice Questions
Question 1. For one mole of a gas, the ideal gas equation is
(a) PV = 1/2RT (b) PV = RT (c) PV = 3/2RT (d) PV = 5/2 RT
Question 2. The average kinetic energy of the gas molecule is
(a) inversely proportional to its absolute temperature
(b) directly proportional to its absolute temperature
(c) equal to the square of its absolute temperature
(d) All of the above
Question 3. Which of the following is the correct mathematical relation for Charles law at constant pressure?
(a) $V \alpha T$ (b) $V \alpha t$. (c) $V \alpha 1/T$ (d) all of above
Question 4. At constant temperature, the pressure of the gas is reduced to one-third, the volume

(a) reduce to one-third (b) increases by three times
(c) remaining the same (d) cannot be predicted
Question 5. With rise in temperature, the surface tension of a liquid
(a) decreases (b) increases
(c) remaining the same (d) none of the above
Question 6. Viscosity of a liquid is a measure of
(a) repulsive forces between the liquid molecules
(b) frictional resistance
(c) intermolecular forces between the molecules
(d)none of the above
Question 7. The cleansing action of soaps and detergents is due to
(a) internal friction (b) high hydrogen bonding
(c) viscosity (d) surface tensions
Question 8. In van der Waals equation of state for a non-ideal gas the net force of attaction among the molecules is given by
(a) $\frac{an^2}{V^2}$ (b) $P+\frac{an^2}{V^2}$ (c) $P-\frac{an^2}{V^2}$ (d) $-\frac{an^2}{V^2}$
Question 9. The compressibility factor, z for an ideal gas is
(a) zero (b) less than one (c) greater than one (d) equal to one
Question 10. Which of the following gases will have the lowest rate of diffusion?
(a) $H_2$ (b) $N_2$ (C) $F_2$ (d) $O_2$
Answer: 1. (b) 2. (b) 3. (a) 4. (b) 5. (a)
6. (b) 7. (d) 8. (a) 9. (d) 10. (c)
V. HOTS Questions
Question 1. (a) Why aerated water bottles kept under water during summer?
(b) Which property of liquid is responsible for spherical shape of drop?
(c) Why is moist air lighter than dry air?
(d) Define aqueous tension.
(e) What are units of a and b which are van der Waals constants?
Answer: (a) To reduce temperature, so as to reduce pressure, otherwise

bottle may burst.

(b) Surface Tension.

(c) Moist air has water vapours which lowers vapour density, so it is lighter.

(d) It is pressure of water vapours at given temperature.

(e) Unit of a is $L^2\ mol^{-2}$, b is $L\ mol^{-1}$.

Question 2. Why does sharpened edge becomes smooth on heating up to melting point?

Answer: On heating the glass, it melts and take up rounded shape at edges which has minimum surface area b/c of surface tension.

Question 3. Arrange the following in order of increasing density:

$d = \frac{PM}{RT}$, $O_2$ at 25°C, 2 atm. $O_2$ at 0°C, 2 atm. $O_2$ at 273°C, 1 atm.

Answer:

$d = \frac{PM}{RT}$, R and M are constant, so $d$ depends upon $\frac{P}{T}$. So at 25°C, 1 atm, $\frac{P}{T} = \frac{1}{298}$. At 273° C, 1 atm, $\frac{P}{T} = \frac{1}{546}$. Hence, increasing order of density will be:

$O_2$ at 273° C, 1 atm < $O_2$ at 25° C, 2 atm.

Question 4. An $O_2$ cylinder has 10 $LO_2$ at 200 atm. If patient takes 0.50 ml of $O_2$ at 1 atm in one breath 37°C, how many breaths are possible?

Answer:

$P_1 = 200$ atm, $V_1 = 10L$

$P_2 = 1$ atm, $V_2 = ?$

$P_1V_1 = P_2V_2 \Rightarrow 200 \times 10 = 1 \times V_2$ or $V_2 = 2000$ L.

$$\text{No. of breathes} = \frac{\text{Total Volume}}{\text{Volume for 1 breath}} = \frac{2000L}{0.5 \times 10^{-3}L} = 4\times10^6$$

* * * * * * * * * * * * * * * * * * * * * *