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- (iii) Differentiate b/w adsorption & partition chromatography.
- (iv) What are different methods used to dry crystals?
- (v) Write down the uses of chromatography.
- (vi) What are the major steps for complete quantitative determination of a sample?
- (vii) Name the solvents commonly used in crystallization.
- (viii) How the undesirable colour can be removed in the process of crystallization.
- (ix) Differentiate between qualitative and quantitative analysis.
- (x) Name the substances that can be purified by the process of sublimation.
- (xi) Write down the distribution equation for the distribution of I_2 b/w CCl_4 and H_2O ?
- (xii) What is analytical chemistry?

Q3. Answer any Eight parts from the followings.

- (i) How does Gooch Crucible increases rate of filtration?
- (ii) Why fluted filter paper is used?
- (iii) Why concentrated HCl and $KMnO_4$ solutions cannot be filtered by Gooch Crucible?
- (iv) Write the names of the major steps involved in the crystallization?
- (v) What is mother liquor?
- (vi) What is the safe and reliable method of drying crystals?
- (vii) Why crystallization is used?
- (viii) Write down the main characteristics of the solvent used for crystallization?
- (ix) Define distribution Law or partition Law?
- (x) How the mixture of naphthalene and sand is separated?
- (xi) What is ether extraction?
- (xii) What is paper chromatography? What are its types?

Q4. Answer any Six parts from the followings.

- (i) How R_f value is calculated?
- (ii) Define sublimation with an example?
- (iii) State solvent extraction and give its importance?
- (iv) How filter media is selected for filtration?
- (v) How value of distribution coefficient affects the separation of components of a mixture in chromatography?
- (vi) Why hot filtration is necessary during crystallization?
- (vii) What is chromatogram?
- (viii) Describe method to collect crystals from mother liquor?
- (ix) How iodine is dissolved in water?

Section - II (Attempt any three questions) (8x 3)=24

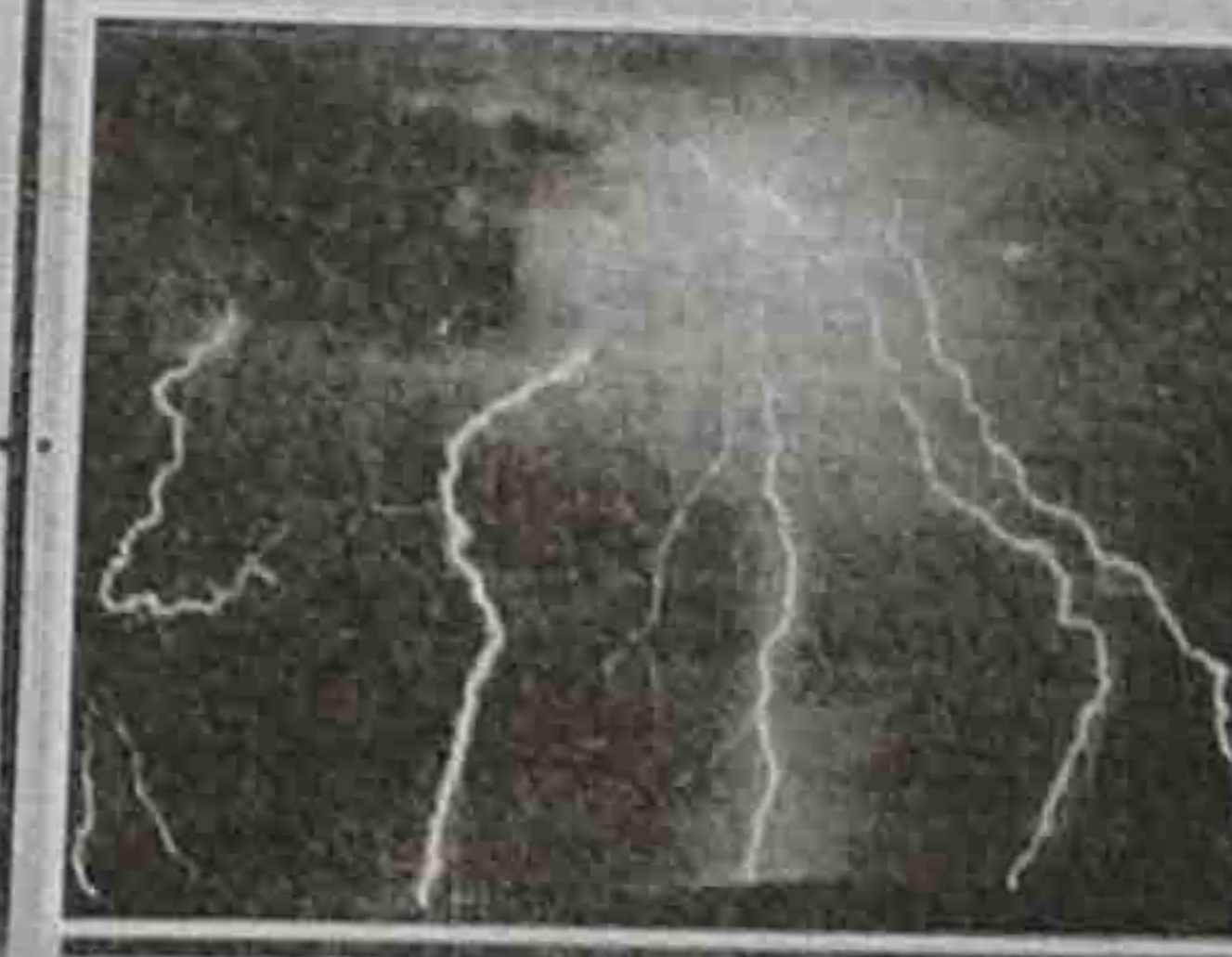
- Q5.** (a) Write down five characteristics for an ideal solvent for crystallization? (04)
- (b) What is the disadvantage of slow cooling in crystallization? (02)
- (c) What is the basic principle of crystallization? (02)
- Q6.** (a) What is distribution law? Give examples. (04)
- (b) Define the following (04)
- (i) Sublimation (04)
- (ii) Chromatography (04)
- Q7.** (a) What are the important considerations for a good filtration process. (02)
- (b) Why repeated extractions using smaller portions of solvent are more beneficial than using single extraction with larger volume of solvent? (02)
- (c) How solvent extraction is carried out? (02)
- Q8.** (a) How many different ways are there to carry out paper chromatography? Briefly describe the procedure for ascending paper chromatograph. (04)
- (b) Stem of funnel in filtration should be several inches long. Why? (02)
- (c) What type of substances cannot be filtered through Gooch crucible using filter paper? (02)
- Q9.** (a) What are the common laboratory experimental techniques? (04)
- (b) A solid organic compound is soluble in water as well as in chloroform. During its preparation, it remains in aqueous layer. Describe a method to obtain it from this layer. (04)
- (b) How will you prepare a fluted filter paper? (02)

Chapter 3

GASES



Avogadro



Lightening bolt (Plasma)



Aurora (Plasma)

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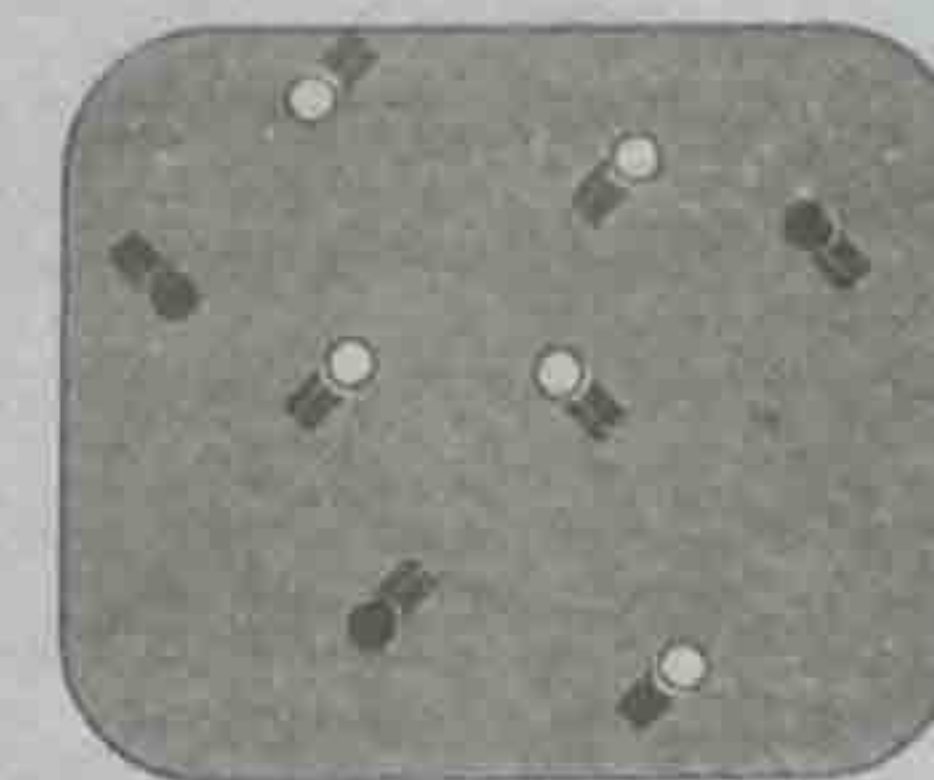
PLASMA STATE

Objective and short answer questions (exercise)

Numerical problems (exercise)

Past Papers MCQs and Short Questions

Test your skills



STATES OF MATTER

Matter exists in four states.

(i) Gas (ii) Liquid (iii) Solid (iv) Plasma

Liquids are less common than solids, gases and plasmas. It is because the liquid state of any substance can exist only within a relatively narrow range of temperature.

Properties of Gases

1. **Indefinite Volume:** Gases have indefinite volume. The volume of the gas is equal to the volume of container.
2. **Indefinite Shape:** Gases have indefinite shape. They adopt the shape of the container in which they are stored.
3. **Low density:** Gases have low densities than solids and liquids. Thus gases bubble through liquids and tend to rise up.
4. **Joule-Thomson Effect:** When gases expand suddenly, they cause cooling. This effect is known as Joule-Thomson effect. This effect is used in the liquefaction of gases on industrial scale.
5. **Effusion & Diffusion:** Gases can effuse and diffuse.
6. **Effect of Temperature:** Gases expand on heating & contract on cooling i.e., increase in temperature increases the volume of gas & vice versa. Liquids and solids do not show considerable increase in volume by increasing temperature.
7. **Effect of Pressure:** Gases expand on decreasing pressure and contract on increasing pressure.
8. **Gas Pressure:** Gas molecules are in constant random motion. During their motion, they collide with the wall of container. These collisions develop a pressure called Gas Pressure.
9. **Intermolecular Forces:** Gases have very weak intermolecular forces.

Properties of Liquids

1. **Definite volume:** Liquids have definite volume.

- Indefinite Shape:** Liquids have indefinite shape. They adopt the shape of the container in which they are stored.
- Constant motion:** In liquids, molecules are in constant motion. Evaporation and diffusion properties of liquids are due to this motion.
- Densities:** The densities of liquids are greater than gases and close to solids.
- Intermolecular forces:** The intermolecular forces among the liquid molecules are stronger than gases but weaker than solids. The melting points and boiling points of liquids depend upon the strength of these forces.
- Spaces:** The spaces among the liquid molecules are negligible just like solids.
- Kinetic energy:** Liquid molecules have K.E. Liquids can be converted into solids by cooling i.e. by decreasing the K.E. of the molecules.
- Collisions:** Molecules of liquids collide with one another and exchange energy.
- Diffusion:** Liquids can diffuse into other liquids. However, the rate of diffusion is smaller than that of gases.

Properties of solids

- Definite shape:** Solids have definite shape.
- Definite volume:** Solids have definite volume.
- Spaces:** Molecules of solids are very close to each other. They are tightly packed.
- Incompressible:** Solids are incompressible due to tight packing.
- Intermolecular forces:** Solids have strongest intermolecular forces.
- Vibrational motion:** Solid particles only show vibrational motion.
- Diffusion:** Solids have negligible rate of diffusion.

Pressure and Units of Pressure

Pressure is the force per unit area. i.e. $P = F/A$

- The SI unit of pressure is Nm^{-2} (Pascal). $1\text{Pa} = 1\text{Nm}^{-2}$.
- The unit pound per square inch (psi) is most commonly used in engineering work.
- The unit millibar (mb) is commonly used by meteorologists.
- In chemical work, pressure is defined in terms of standard atmospheric pressure.

The pressure of air that can support 760 mmHg column at sea level, is called one atmosphere. It is the force exerted by 760mm or 76cm long column of mercury on an area of 1cm^2 at 0°C .

It is the average pressure of atmosphere at sea level. Its symbol is '*atm*'. So, one atmosphere can support 760 mm column of mercury at sea level.

Interconversion of various pressure units

$$1\text{ atm} = 760\text{ torr} = 760\text{ mm of Hg} = 101325\text{ Nm}^{-2} = 101325\text{ Pa} = 101.325\text{ kPa} \\ (\text{kilopascal}) = 14.7\text{ pound inch}^{-2} = 1013.25\text{ millibars.}$$

GAS LAWS

All gases behave uniformly. When pressure and temperature are changed, the volume of gases is changed. The gas laws describe this uniform behaviour of gases.

The relationships between volume of gases and the external conditions like temperature and pressure are called the gas laws.

Exercise Q.4. (a):

What is Boyle's law of gases? Give its experimental verification.



Boyle's Law

Robert Boyle gave this law in 1662. It states,

At constant temperature, volume of a given mass of gas is inversely proportional to the pressure exerted on it.

Mathematically $V \propto \frac{1}{P}$ (When 'T' and 'n' are constants)

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

$$\text{or } PV = k \quad (1) \quad (\text{When 'T' and 'n' are constants})$$

Eq (1) gives another statement of Boyle's law

At constant temperature, the product of pressure and volume of a given mass of a gas is always constant.

If the pressure of a given mass of gas is ' P_1 ' and volume ' V_1 ', then by changing pressure to ' P_2 ', volume is changed to ' V_2 ', such that

$$\text{So } P_1 V_1 = P_2 V_2$$

Experimental Verification of Boyle's Law

- Consider a gas in a cylinder fitted with a movable piston at 25°C . The cylinder is fitted with a manometer to read the pressure of the gas directly.
- Let the initial volume of gas enclosed in cylinder is 1 dm^3 and pressure is 2 atm.

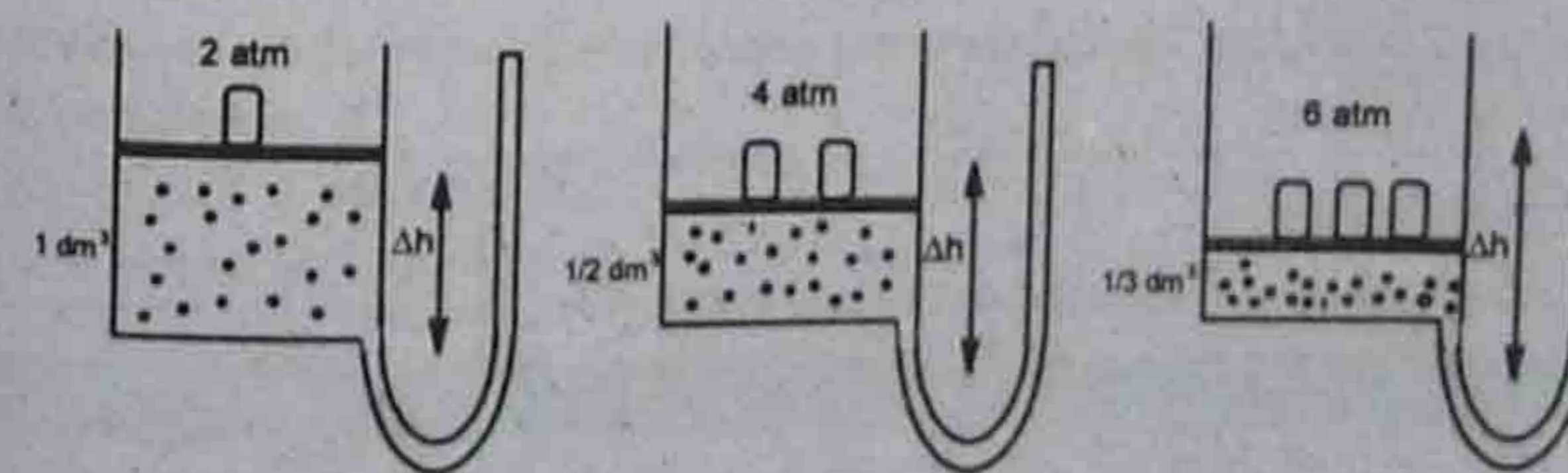


Figure: Verification of Boyle's law

- Then according to Boyle's law

$$P_1 V_1 = 2 \times 1 = 2 \text{ atm dm}^3 \text{ atm} = k$$

- When pressure is doubled on the gas, its volume becomes half.

$$\text{Thus } P = 4 \text{ atm and } V = \frac{1}{2} \text{ dm}^3$$

Therefore

$$P_2 V_2 = 4 \times \frac{1}{2} = 2 \text{ atm dm}^3 = k$$

- Similarly if pressure is tripled (i.e. 6 atm) on the gas, its volume is reduced to $\frac{1}{3}$.

Therefore

$$P_3 V_3 = 6 \times \frac{1}{3} = 2 \text{ atm dm}^3 = k$$

Since product of P and V is always constant. Hence Boyle's law is verified.

Example 1:

A gas having a volume of 10 dm^3 is enclosed in a vessel at 0°C and the pressure is 2.5 atmospheres. This gas is allowed to expand until the new pressure is 2 atmospheres. What will be the new volume of this gas, if the temperature is maintained at 273 K ?

Solution:

$$V_1 = 10 \text{ dm}^3 \quad V_2 = ?$$

$$P_1 = 2.5 \text{ atm} \quad P_2 = 2 \text{ atm}$$

Since the temperature is constant ($T_1 = T_2$), thus according to Boyle's law

$$P_1 V_1 = P_2 V_2$$

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

$$V_2 = \frac{2.5 \text{ atm} \times 10 \text{ dm}^3}{2 \text{ atm}} = 12.5 \text{ dm}^3$$

Graphical explanation of Boyle's law

Exercise Q.4. (b):

What are isotherms? What happens to the positions of isotherms when they are plotted at high temperature for a particular gas? (Lahore Board, 2014)

Graph between V and P (Isotherms)

- The plot of volume of a gas on Y-axis and pressure on X-axis at constant temperature (e.g. at 0°C) gives a curve called isotherm.
- If this graph is plotted at higher constant temperature (e.g., at 25°C), then the curve goes away from both axis. It is

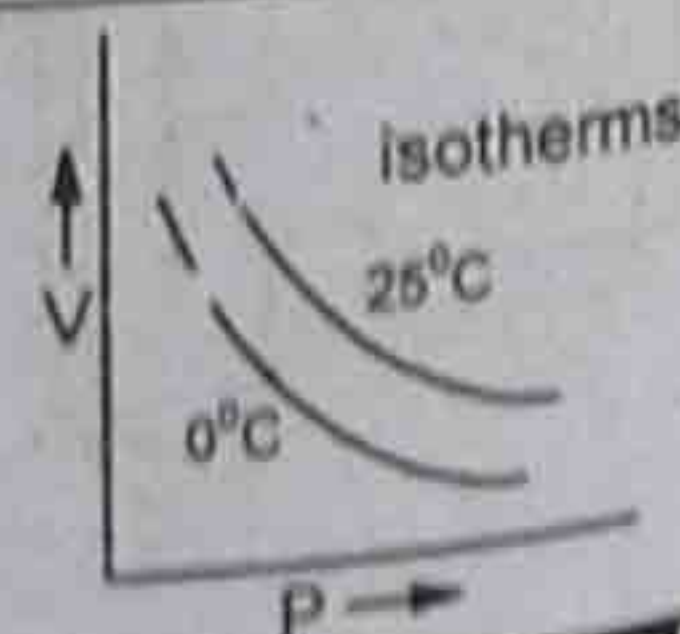


Figure: Isotherms of a gas at different temperatures

because, at high temperature, the volume of gases is increased.

Similarly, if curve is plotted at further higher temperature, then the curve further goes away from the two axes.



Exercise Q.4. (b):

Why do we get a straight line when pressures exerted on a gas are plotted against inverse of volumes? This straight line changes its position in the graph by varying the temperature. Justify it.

Graph between P and $1/V$

- The plot of pressure (P) on Y-axis against inverse volume ($1/V$) of a gas, on X-axis gives a straight line. It is because increase in pressure decreases the volume hence inverse of volume increases. Thus, ' P ' is directly proportional to $1/V$.

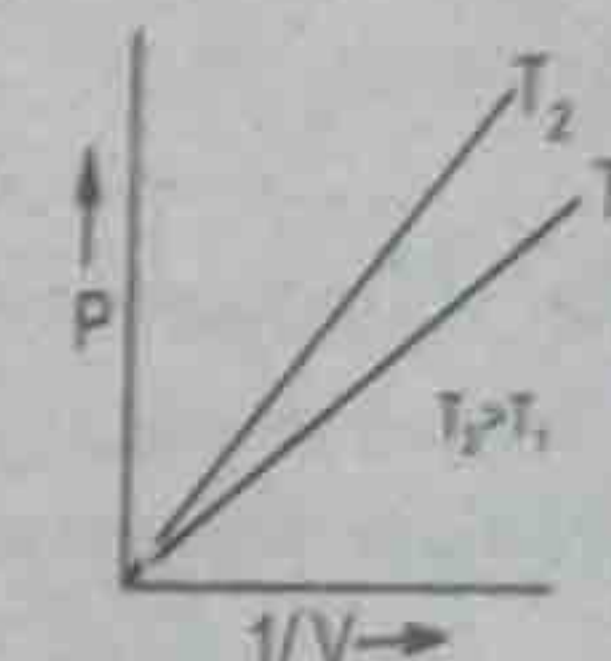


Figure: A plot between P and $1/V$

- At higher temperature straight line becomes closer to Y-axis. It is because at higher temperature volume of a gas is increased, therefore, $1/V$ is decreased.
- The straight line passes through the origin because when the pressure is close to zero then volume is so high that $1/V$ is very close to zero.

Graph between PV and P

- A plot of ' P ' on X-axis against ' PV ' on Y-axis will give a straight line parallel to X-axis. The straight line shows that ' k ' constant quantity.
- At higher temperature, volume of gas is increased. But new value of PV also remains constant. Thus, a straight line parallel X-axis is again obtained at higher temperature.
- This straight line helps us to understand the non-ideal behaviour of gases.

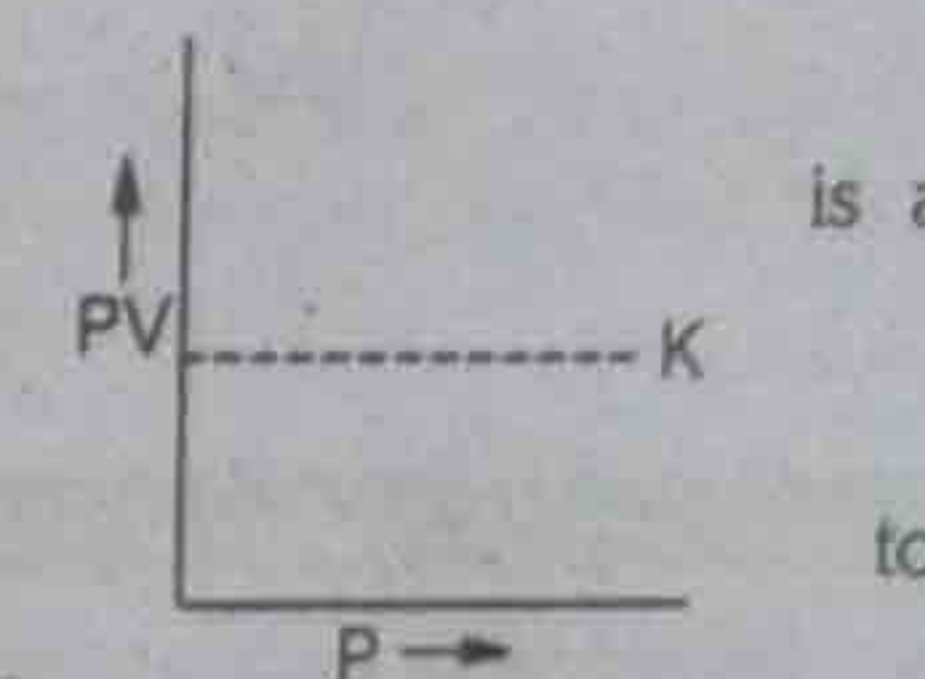


Figure: A plot between pressure and product of PV

Boyle's law is only applicable to ideal gases.



Exercise Q5 (a):

What is the Charles's law? Which scale of temperature is used to verify that $V/T = k$ (pressure and number of moles are constant)?

CHARLES'S LAW

Charles, a French scientist, gave this law in 1787.

It states

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

Mathematically

$$V \propto T \quad (\text{at constant 'P' and 'n'})$$

$$V = kT$$

$$\frac{V}{T} = k \quad (1)$$

where 'T' is the temperature in Kelvin

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

Eq. (1) also shows that

At constant pressure, the ratio between the volume of a given mass of a gas and its absolute temperature is always constant.

Graph:

The graph between 'V' of a given mass of gas and 'T' will be a straight line. It is because, 'V' is directly proportional to the absolute temperature.

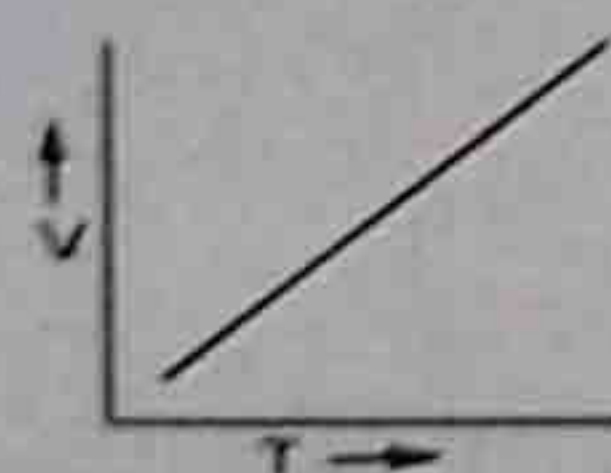


Figure: Graph between V and T

Experimental Verification of Charles's Law:

- Consider a gas, enclosed in a cylinder fitted with a moveable piston. At temperature T_1 , the volume of gas is V_1 .
- By increasing the temperature to ' T_2 ', its volume is increased to ' V_2 '. The pressure on the piston is kept constant.

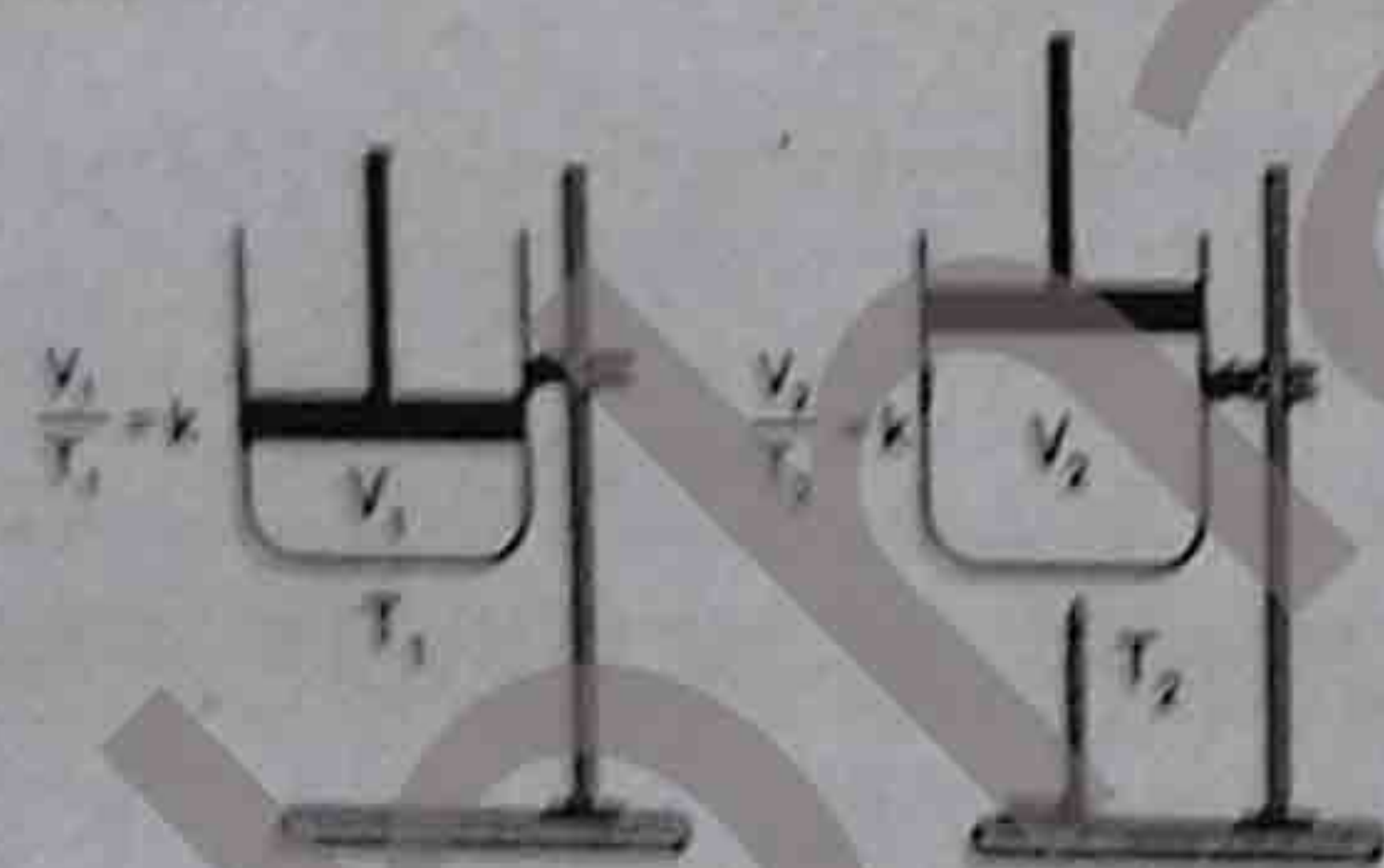


Figure: Verification of Charles's law

- It is observed that ratio between volume of a given mass of gas and temperature remains constant at constant pressure.

i.e;

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

Verification of Temperature Scale for Charles's Law

The mathematical form of Charles's law is

$$\frac{V}{T} = K$$

This equation only holds if temperature is on Kelvin scale.

Example:

Consider the following data for a gas when the temperature is on Kelvin scale.

V (dm ³)	1092	846	746
T (K)	546	423	373
V/T (dm ³ K ⁻¹)	2	2	2

Hence value of $\frac{V}{T}$ is always constant.

However, if temperature is taken on Celsius scale, then the above relation is not satisfied. e.g. Consider the following data for above gas when the temperature is on centigrade scale.

V (dm ³)	1092	846	746
T (°C)	273	150	100
V/T (dm ³ °C ⁻¹)	4	5.64	7.46

Hence, value of $\frac{V}{T}$ does not remain constant.

Thus, Charles's law is only obeyed if temperature is on Kelvin scale.

Example 2:

250 cm³ of hydrogen is cooled from 127°C to -27°C by maintaining the pressure constant. Calculate the new volume of the gas at low temperature.

Solution

$$V_1 = 250 \text{ cm}^3$$

$$V_2 = ?$$

$$T_1 = 127^\circ\text{C} + 273 = 400 \text{ K} \quad T_2 = -27^\circ\text{C} + 273 = 246 \text{ K}$$

According to Charles's law

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

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

$$V_2 = \frac{250 \times 246}{400} = 153.8 \text{ cm}^3 = \boxed{0.1538 \text{ dm}^3}$$

Exercise Q.6. (a):

What is Kelvin scale of temperature? Plot a graph for one mole of a real gas to prove that a gas becomes liquid, earlier than -273.16°C . (Graphical Explanation)

Absolute Zero

The hypothetical temperature at which the volume of all gases becomes zero is called absolute zero.

$$\text{Its value is } -273.16^\circ\text{C} = 0 \text{ K}$$

For routine calculations the value of absolute zero is taken as -273°C

Mathematical Explanation: Quantitative statement of Charles's law

At constant pressure, the volume of a given mass of an ideal gas increases or decreases by $1/273$ of its original volume at 0°C for every 1°C rise or fall in temperature respectively.

$$\text{Thus Volume at } T^\circ\text{C} = V_T = V_0 \left(\frac{273 + T}{273} \right) \quad (1)$$

where V_T = Volume at $T^\circ\text{C}$, T = Temperature in $^\circ\text{C}$

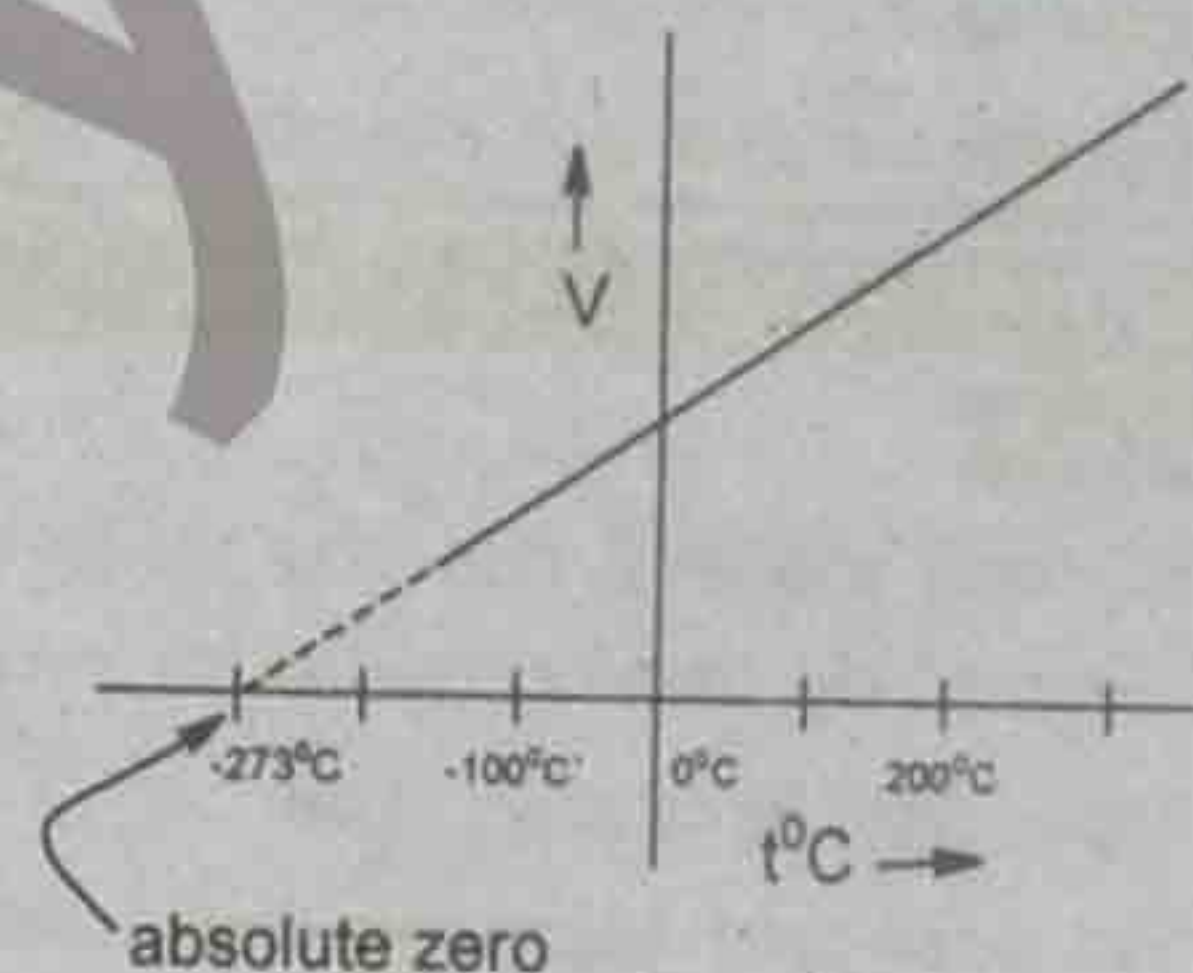
Therefore at -273°C

$$V_{-273} = V_0 \left(\frac{273 - 273}{273} \right) = 0$$

- The temperature -273°C is called Absolute Zero of Kelvin scale. Thus volume of a gas becomes zero at absolute zero.
- Absolute zero can never be achieved. It is considered as the lowest temperature. Its value is independent of the nature of gas.
- Charles's law is not obeyed when the temperature is on Centigrade scale. For this reason Kelvin scale was developed with $0 \text{ K} = -273^\circ\text{C}$

Graphical Explanation

- According to Charles's law, when a graph is plotted between ' V ' and ' T ' for a gas, a straight line is obtained
- This line intersects the temperature axis at -273°C , which is considered as the lowest temperature.
- This temperature is achieved if the substance remains in the gaseous form. But all gases liquefy before reaching this temperature.
- Thus, the lines of all the gases are extrapolated. They meet ' T ' axis at -273°C at which the volume of all gases becomes zero. However, it can never happen for a real gas. Thus, -273°C is taken as zero of Kelvin scale and it is called absolute zero.
- Generally, greater the mass of the gas greater will be the slope of the line. It is because greater number of moles of gas occupies more volume.



Fig(3.7) The graph between volume and temperature ($^\circ\text{C}$)

SCALES OF THERMOMETRY

The branch of science which deals with the measurement of temperature is called thermometry.

There are three measuring scales of thermometry.

(i) Centigrade or Celsius Scale (ii) Fahrenheit scale (iii) Kelvin Scale

(i) Centigrade or Celsius Scale ($^\circ\text{C}$)

The temperature measured on this scales is represented by $^\circ\text{C}$.

On this scale freezing point of water is marked as 0°C and boiling point as 100°C . The distance between these two ends is divided into 100 equal parts. Each part is equal to 1°C .

(ii) Fahrenheit Scale: ($^\circ\text{F}$)

The temperature measured on this scale is represented by $^\circ\text{F}$.

On this scale, the freezing point of water is marked as 32°F and boiling point as 212°F . The distance between these two ends is divided into 180 equal parts. Each part is equal to 1°F .

(iii) Kelvin scale (K)

The temperature measured on this scale is represented by K.

Interconversions of different scales of thermometry

$$\text{Fahrenheit to Celsius} \quad ^\circ\text{C} = \frac{5}{9} (^\circ\text{F} - 32)$$

$$\text{Celsius to Fahrenheit} \quad ^\circ\text{F} = \frac{9}{5} ^\circ\text{C} + 32$$

$$\text{Kelvin to Celsius and vice versa} \quad \text{K} = ^\circ\text{C} + 273.16$$

Exercise Q.7. (a):

What is the general gas equation? Derive it in various forms?

GENERAL GAS EQUATION**Derivation**

For a given mass of gas, the three variables i.e.; pressure, temperature and volume can be inter-related by one equation known as "General Gas Equation" or "Ideal Gas Equation."

According to Boyle's law

"At constant temperature, volume of a given mass of a gas is inversely proportional to the pressure exerted on it" i.e;

$$V \propto \frac{1}{P} \quad (\text{at fixed } T \text{ \& } n) \quad (1)$$

According to Charles' law

"At constant pressure, the volume of a given mass of a gas is directly proportional to the absolute temperature." i.e;

$$V \propto T \quad (\text{At fixed } P \text{ \& } n) \quad (2)$$

According to Avogadro's law

"The volume of a gas is directly proportional to the mass (moles) of the gas at constant temperature and pressure i.e;

$$V \propto n \quad (\text{At fixed } T \text{ \& } P) \quad (3)$$

Combining eq. (1), (2), (3),

$$\text{So } V \propto \frac{nT}{P}$$

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

$$\text{or } PV = nRT \quad (4)$$

where 'R' is constant known as universal gas constant.

Eq. (4) is known as general gas equation.

For 1 mole of a gas

$$PV = RT \quad \text{or} \quad \frac{PV}{T} = R$$

If P, V & T are changed for a gas from P_1, V_1 & T_1 to P_2, V_2 & T_2 .

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

**Reduction of General Gas Equation to Boyle's law, Charles's law and Avogadro's law.**

$$\text{As } PV = nRT$$

If 'T' and 'n' are constant then $PV = k$ (Boyle's Law)

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

If 'P' and 'n' are constant then $V = kT$ (Charles's Law)

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

If 'P' and 'T' are constant then $V = kn$ (Avogadro's Law)

**Exercise Q7 (b):**

Can we determine the molecular mass of an unknown gas if we know the pressure, temperature and volume along with the mass of the gas.

Molecular Mass of Gases

The general gas equation is

$$PV = nRT \quad (1)$$

Since

$$\text{Number of moles } n = \frac{\text{Given mass}}{\text{Molecular mass}} = \frac{m}{M}$$

So equation (1) becomes

$$PV = \frac{m}{M} \times RT \quad (2)$$

$$\text{or } M = \frac{mRT}{PV} \quad (3)$$

Thus knowing, pressure (P), volume (V), temperature (T) and mass (m) of a gas, its molecular mass (M) can be calculated.

**Exercise Q7 (c):**

How do you justify from general gas equation that increase in temperature or decrease of pressure decreases the density of the gas?

Density of Gases

The general gas equation is

$$PV = nRT \quad (1)$$

Since $n = \frac{\text{Given mass}}{\text{Molecular mass}} = \frac{m}{M}$

So eq (4) becomes

$$PV = \frac{m}{M} \times RT \quad (2)$$

or $PM = \frac{m}{V} \times RT \quad (3)$

Since, $d = \frac{m}{V}$ so eq. (3) becomes

$$PM = dRT$$

$$d = \frac{PM}{RT} \quad (4)$$

Hence

- Knowing, pressure (P), temperature (T) and molecular mass (M) of gas, its density (d) can be calculated.
- The equation shows that density of an ideal gas is directly proportional to its molecular mass and pressure and inversely proportional to the temperature. So, increase in 'M' and 'P', increases the density while increase in 'T' decreases the density due to increase in volume.

IDEAL GAS CONSTANT 'R'

- Numerical value of R is totally independent of the nature of the gas.
- It depends on the units of 'P' and 'V'.
- The value of 'R' is calculated for one mole of a gas at STP using Avogadro's Law.

According to Avogadro's law, At STP one mole of a gas occupies 22.414 dm³.

Non-SI units

At STP

$$n = 1.0 \text{ mole} \quad T = 0^\circ\text{C} = 273\text{K} \quad P = 1.0 \text{ atm}$$

$$V = 22.414 \text{ dm}^3 \quad R = ?$$

$$R = \frac{PV}{nT} = \frac{1 \times 22.414}{1 \times 273}$$

$$R = 0.0821 \text{ dm}^3 \text{ atm mol}^{-1} \text{ K}^{-1}$$

or $R = 62.4 \text{ dm}^3 \text{ mm of Hg mol}^{-1} \text{ K}^{-1}$

or $R = 62.4 \text{ dm}^3 \text{ torr mol}^{-1} \text{ K}^{-1}$

or $R = 62400 \text{ cm}^3 \text{ torr mol}^{-1} \text{ K}^{-1}$

SI units

At STP

$$P = 1 \text{ atm} = 101325 \text{ Nm}^{-2}$$

$$T = 273.16 \text{ K}$$

$$n = 1.0 \text{ mole}$$

$$V = 22.414 \text{ dm}^3 = 0.022414 \text{ m}^3$$

$$R = ?$$

$$R = \frac{PV}{nT} = \frac{101325 \times 0.022414}{1 \times 273.16}$$

$$R = 8.3143 \text{ Nm mol}^{-1} \text{ K}^{-1} \quad \text{or} \quad R = 8.3143 \text{ J mol}^{-1} \text{ K}^{-1}$$

Since $1 \text{ cal} = 4.18 \text{ J}$

So $R = \frac{8.3143}{4.18} \text{ cal K}^{-1} \text{ mol}^{-1}$

$$R = 1.987 \text{ cal mol}^{-1} \text{ K}^{-1}$$

Thus unit of 'R' are thus expressed in terms of "energy mol⁻¹ K⁻¹"

Physical Significance of value of 'R'

- The value of 'R' in SI unit is 8.3143 J mol⁻¹ K⁻¹. It shows that if one mole of an ideal gas is present at 273 K and 1 atmospheric pressure and its temperature is increased by 1K, then it will absorb 8.3143 J of energy.
- The value of 'R' in non-SI unit is 0.0821 dm³ atm mol⁻¹ K⁻¹. It shows that if one mole of an ideal gas is present at 273 K and 1 atmospheric pressure and its temperature is increased by 1K, then it will absorb 0.0821 dm³ atm of energy. (dm³ atm is the unit of energy)

Example 3

A sample of nitrogen gas is enclosed in a vessel of volume 380 cm³ at 120°C pressure of 101325 Nm⁻². This gas is transferred to a 10 dm³ flask and cooled to 27°C. Calculate the pressure in Nm⁻² exerted by the gas at 27°C. (Gujranwala Board, 2008; Multan Board, 2012)

Solution:

$$V_1 = 380 \text{ cm}^3 = 0.38 \text{ dm}^3 \quad V_2 = 10 \text{ dm}^3$$

$$T_1 = 120^\circ\text{C} + 273 \text{ K} = 393 \text{ K} \quad T_2 = 27^\circ\text{C} + 273 \text{ K} = 300 \text{ K}$$

$$P_1 = 101325 \text{ Nm}^{-2} \quad P_2 = ?$$

According to general gas equation

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

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

$$P_2 = \frac{101325 \times 0.38 \times 300}{393 \times 10} = \boxed{2939.2 \text{ Nm}^{-2}}$$

Example 4

Calculate the density of CH_4 (g) at 0°C and 1 atmospheric pressure. What will happen to the density if (a) temperature is increased to 27°C , (b) the pressure increased to 2 atmospheres at 0°C .

OR Calculate the density of methane at STP. (Lahore Board, 2010) OR Calculate the density of methane at 0°C and 760 mm Hg pressure. (D.G. Khan Board, 2011)

Solution:

$$T = 0^\circ\text{C} + 273 \text{ K} = 273 \text{ K}$$

$$P = 1 \text{ atm}$$

$$\text{Mol. mass of } \text{CH}_4 \quad M = 16 \text{ g mol}^{-1}$$

$$R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$$

$$d = ?$$

Density is given by the formula

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

$$d = \frac{1 \times 16}{0.0821 \times 273}$$

$$d = \boxed{0.7139 \text{ g dm}^{-3}}$$

(a) Density at 27°C

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

Density is given by the formula

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

$$d = \frac{1 \times 16}{0.0821 \times 300}$$

$$d = \boxed{0.649 \text{ dm}^{-3}}$$

(b) Density at 2 atmospheric pressure and 0°C

$$T = 0 + 273 = 273 \text{ K}$$

$$P = 2 \text{ atm}$$

Density is given by the formula

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

$$d = \frac{2 \times 16}{0.0821 \times 273}$$

$$d = \boxed{1.427 \text{ dm}^{-3}}$$

Example 5

Calculate the mass of 1 dm^3 of NH_3 gas at 30°C and 1000 mm Hg pressure, considering that NH_3 is behaving ideally. (Multan Board, 2012; Lahore Board, 2014)

Solution:

$$P = 1000 \text{ mm Hg} = \frac{1000}{760} \text{ atm} = 1.316 \text{ atm}$$

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

$$T = 30^\circ\text{C} + 273 \text{ K} = 303 \text{ K}$$

$$\text{Molecular Mass of } \text{NH}_3 = M = 17 \text{ g mol}^{-1}$$

$$\text{Mass of } \text{NH}_3 = m = ?$$

$$R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$$

Mass of the gas is given by

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

$$\text{or } m = \frac{PVM}{RT}$$

$$= \frac{1.316 \times 1 \times 17}{0.0821 \times 300} = \boxed{0.908 \text{ g}}$$

Exercise Q.9. (a):

What is Avogadro's law of gases?

AVOGADRO'S LAW

It states

"Equal volumes of all the ^{ideal} gases at the same temperature and pressure contain equal number of molecules (or moles) "

So, number of moles of a gas is directly proportional to its volume.

$$\text{i.e.; } V \propto n$$

- It shows that same number of moles of all gases occupies same volume at same temperature & pressure. The volume occupied by one mole of any gas at STP is called molar volume and is equal to 22.414 dm³.
- One mole of any gas contains constant number of molecules. This number is called Avogadro's number. Its value is 6.02 × 10²³

Examples and Explanation:

$$1 \text{ mole of O}_2 = 32 \text{ g O}_2 = 22.414 \text{ dm}^3 \text{ at STP} = 6.02 \times 10^{23} \text{ molecules of O}_2$$

$$1 \text{ mole of H}_2 = 2.016 \text{ g H}_2 = 22.414 \text{ dm}^3 \text{ at STP} = 6.02 \times 10^{23} \text{ molecules of H}_2$$

So,

1 mole of O₂ and H₂ occupies same volume at STP, although O₂ is 16 times heavier than H₂.

It is because molecules in gases are widely separated from one another and have large empty spaces. The distance between two molecules is approximately 300 times the diameter of the molecule. Therefore, size and masses of gas molecules do not affect their volume. Hence, equal moles of all gases occupy same volume at STP.

Facts

- One dm³ of every gas at STP will have molecules = $\frac{6.02 \times 10^{23}}{22.414} = 2.68 \times 10^{22}$ molecules. If temperature and pressure are changed equally for all gases, each gas will have same 2.68×10^{22} molecules.
- One dm³ of H₂ at STP weighs 0.0899 g. (since $\frac{2.016}{22.414} = 0.0899\text{g}$) and one dm³ of O₂ at STP weighs 1.4384 g (since $\frac{32}{22.414} = 1.4384\text{g}$) but they contain equal number of molecules (i.e. 2.68×10^{22}).

DALTON'S LAW OF PARTIAL PRESSURE

Dalton's law states,

The total pressure exerted by a mixture of non-reacting gases is equal to the sum of the partial pressures of all the gases present in the mixture.

Mathematical Expression

Let partial pressures of different gases in a mixture are P₁, P₂, P₃ ..., then according to Dalton's law, total pressure of this mixture is given by

$$P_t = P_1 + P_2 + P_3 + \dots$$

Partial Pressure:

The pressure of each gas in a mixture of gases is called partial pressure of that gas.

Example and Explanation

Consider four cylinders of same volume each.

Three gases H₂, CH₄ and O₂ are enclosed separately in three cylinders at the same temperature.

Let Pressure of H₂ is 400 torr, pressure of CH₄ is 500 torr and pressure of O₂ is 100 torr. Let the three gases are transferred to the fourth cylinder at the same temperature, then according to Dalton's Law total pressure of the mixture of gases will be

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{CH}_4} + P_{\text{O}_2}$$

$$P_{\text{total}} = 400 + 500 + 100 = 1000 \text{ torr}$$

- Since there are no attractive among the molecules of these gases in a mixture, therefore, every gas exerts its own individual pressure. Hence, total pressure is the sum of individual pressure of these gases.
- Due to independent motion of molecules, general gas equation can be applied to each gas separately.

Thus

$$P_{\text{H}_2} V = n_{\text{H}_2} RT \text{ or } P_{\text{H}_2} = n_{\text{H}_2} \frac{RT}{V}$$

$$P_{\text{CH}_4} V = n_{\text{CH}_4} RT \text{ or } P_{\text{CH}_4} = n_{\text{CH}_4} \frac{RT}{V}$$

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

R, T and 'V' are constants for all gases in a mixture

$$\text{Therefore } P_{\text{H}_2} \propto n_{\text{H}_2}$$

$$P_{\text{CH}_4} \propto n_{\text{CH}_4}$$

$$P_{\text{O}_2} \propto n_{\text{O}_2}$$

Hence, pressure of each gas is directly proportional to its number of moles.

- The total pressure of mixture of gases will be

$$P_t = P_{H_2} + P_{CH_4} + P_{O_2}$$

$$P_t = n_{H_2} \frac{RT}{V} + n_{CH_4} \frac{RT}{V} + n_{O_2} \frac{RT}{V}$$

$$P_t = (n_{H_2} + n_{CH_4} + n_{O_2}) \frac{RT}{V}$$

$$P_t = n_t \frac{RT}{V}$$

Where $n_t = n_{H_2} + n_{CH_4} + n_{O_2}$ = total number of all the gases

$$\text{or } P_t V = n_t RT$$

This equation shows that the total pressure of the mixture of gases is directly proportional to the total number of moles of the gases.

Exercise Q10 (b):

Derive an equation to find out the partial pressure of a gas knowing the individual moles of component gases and the total pressure of the mixture.

Calculation of Partial Pressure of a Gas.

In a mixture of gases, partial pressure of any gas can be calculated if mass or moles of the gas, total pressure and total number of moles of the gases are known.

Consider two gases A and B forming a mixture of gases.

Let total pressure of the mixture is P_t and number of moles n_t

while Partial pressure of gas A is P_A and number of moles n_A

and Partial pressure of gas B is P_B and number of moles n_B

Then we can write

$$P_t V = n_t RT \quad (1)$$

$$P_A V = n_A RT \quad (2)$$

$$P_B V = n_B RT \quad (3)$$

Divide eq. (2) by (1)

MOLE FRACTION (X)

It is the number of moles of a substance divided by total number of moles of all the substances present in the mixture.

It is denoted by X

(see Ch#9, Solutions) for more details.

$$\frac{P_A V}{P_t V} = \frac{n_A RT}{n_t RT}$$

$$\frac{P_A}{P_t} = \frac{n_A}{n_t}$$

$$P_A = \frac{n_A}{n_t} \times P_t$$

or

$$P_A = X_A \times P_t \quad (4)$$

similarly

$$P_B = X_B \times P_t \quad (5)$$

where, X_A and X_B are the mole fractions of gas A and B respectively. Thus, partial pressure of a gas is equal to its mole fraction multiplied by the total pressure.

These equations can be used to determine the partial pressure of the gases in a mixture.

Generally for i^{th} gas in a mixture of gases, we can write $P_i = X_i \times P_t$

Example 6

There is a mixture of hydrogen, helium and methane occupying a vessel of volume 13 dm³ at 37°C and pressure of 1 atmosphere. The mass of H₂ and He are 0.8 and 0.12g respectively. Calculate the partial pressure in mm Hg of each gas in the mixture.

Solution:

Volume of the mixture of gases = $V = 13 \text{ dm}^3$

Temperature of the mixture = $T = 37^\circ\text{C} + 273 = 310 \text{ K}$

Pressure of the mixture = $P_t = 1 \text{ atm}$

Total no. of moles = $n_t = ?$

$$R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$$

Total number of moles is given by

$$P_t V = n_t RT$$

$$n_t = \frac{P_t V}{RT}$$

$$n_t = \frac{1 \times 13}{0.0821 \times 310} = 0.51 \text{ moles}$$

So, the total number of moles of H₂, He and CH₄ = 0.51 moles

Mass of H₂ = 0.8 g

Molar Mass of H₂ = 2.016 g mol⁻¹

$$\text{Number of moles of H}_2 = \frac{0.8 \text{ g}}{2.016 \text{ g mol}^{-1}} = 0.396 \text{ moles}$$

$$\text{Mass of He} = 0.12 \text{ g}$$

$$\text{Molar Mass of He} = 4 \text{ g mol}^{-1}$$

$$\text{Number of moles of He} = \frac{0.12 \text{ g}}{4 \text{ g mol}^{-1}} = 0.03 \text{ moles}$$

$$\begin{aligned} \text{No. of moles of CH}_4 &= \text{Total moles} - (\text{mole of H}_2 + \text{moles of He}) \\ &= 0.51 - (0.396 + 0.03) \\ &= 0.084 \text{ moles} \end{aligned}$$

Mole Fractions can be calculated as

$$x_{\text{H}_2} = \frac{\text{No. of moles of H}_2}{\text{Total Number of moles}} = \frac{0.396}{0.51} = 0.776$$

$$x_{\text{He}} = \frac{\text{No. of moles of He}}{\text{Total Number of moles}} = \frac{0.03}{0.51} = 0.058$$

$$x_{\text{CH}_4} = \frac{\text{No. of moles of CH}_4}{\text{Total Number of moles}} = \frac{0.084}{0.51} = 0.164$$

Partial Pressures can be calculated as

$$\begin{aligned} P_{\text{H}_2} &= x_{\text{H}_2} P \\ &= 0.776 \times 1.00 = 0.776 \text{ atm} \\ &= 0.776 \times 760 = \boxed{589.76 \text{ mm Hg}} \end{aligned}$$

$$\begin{aligned} P_{\text{He}} &= x_{\text{He}} P \\ &= 0.058 \times 1.00 = 0.058 \text{ atm} \\ &= 0.058 \times 760 = \boxed{44.08 \text{ mm Hg}} \end{aligned}$$

$$\begin{aligned} P_{\text{CH}_4} &= x_{\text{CH}_4} P \\ &= 0.164 \times 1.00 = 0.164 \text{ atm} \\ &= 0.164 \times 760 = \boxed{124.64 \text{ mm Hg}} \end{aligned}$$

Exercise Q10 (c):

Explain that the process of respiration obeys the Dalton's law partial pressure.

Applications of Dalton's law of Partial Pressure

1. Respiration

Respiration process depends on the difference in partial pressures.

Partial pressure of oxygen in air (159 torr) is greater than in the lungs (116 torr). Therefore, oxygen moves from air into the lungs. The partial pressure of CO₂ in lungs is greater than in the air. Therefore, it moves out from the lungs into the air.

2. Breathing At High Altitude

At sea level, the partial pressure of oxygen in air is 159 torr and breathing is easier. At high altitude, the partial pressure of oxygen is low. It makes breathing difficult. That's why pilots use pressurized cabin during flying.

3. Breathing In Deep Sea

Deep-sea divers cannot breathe with normal air in their tanks. Instead, they use a mixture of an inert gas and O₂ in which partial pressure of oxygen is adjusted around the required limits. Actually, in sea after 100 feet depth, the diver experiences 3 atm pressure. Thus, normal air cannot be breathed in depth of sea. Moreover, the pressure of N₂ increases in depth of sea and it diffuses in the blood.

4. Collection Of Gases Over Water

Gases are usually collected over water. During this process vapours of H₂O are mixed with the gas.

The total pressure (P_{moist}) of this mixture (H₂O vapours + gas) will be

$$P_{\text{moist}} = P_{\text{g}} + P_{\text{water vapours}}$$

The partial pressure of vapours in gases is called aqueous tension.

$$\text{Thus, } P_{\text{moist}} = P_{\text{g}} + \text{aqueous tension}$$

$$\text{or } P_{\text{g}} = P_{\text{moist}} - \text{aqueous tension}$$

DIFFUSION AND EFFUSION

Diffusion

The spontaneous intermixing of molecules of different gases, due to collision, at a given temperature and pressure is called gaseous diffusion.

Why gases diffuse?

When different gases are mixed together, they wish to have same pressure everywhere. Thus, they go on mixing to form homogenous mixture until their partial pressures become equal everywhere.

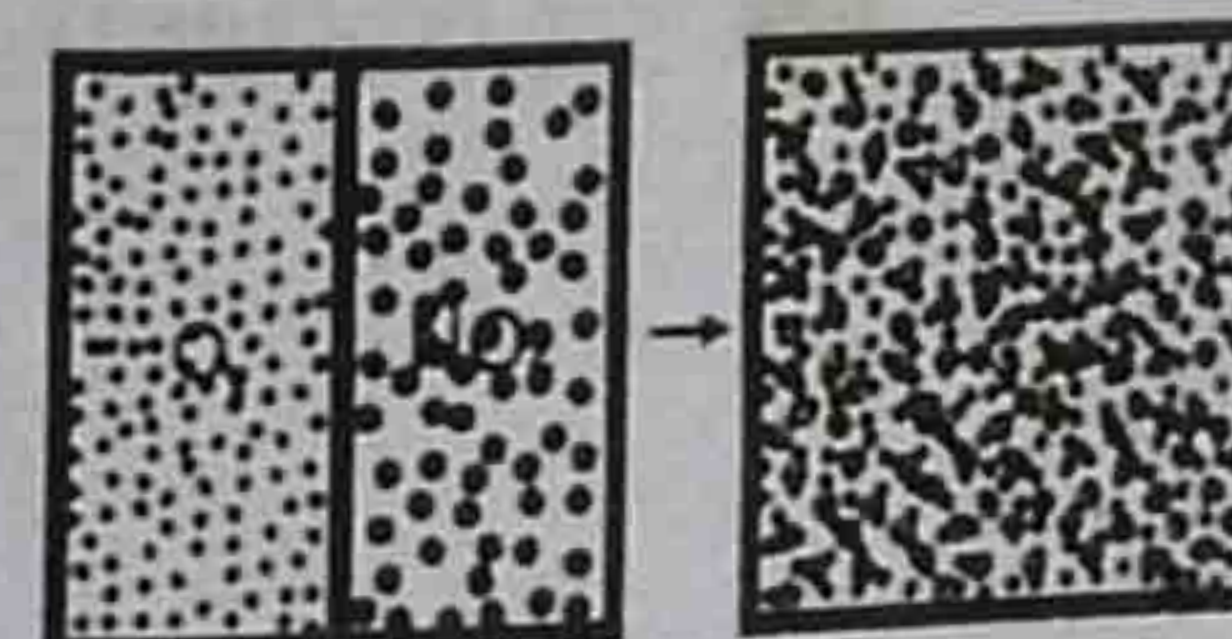


Figure: Diffusion

Examples:

- The smell of rose or a scent spreads due to diffusion.
- The mixing of NO_2 (a brown gas) and O_2 (a colourless gas) is also due to gaseous diffusion as shown in the fig. The diffusion occurs by random motion and collision.

Effusion

It is the escape of gas molecules one by one, without collisions through a hole of molecular size, into a region of low pressure.

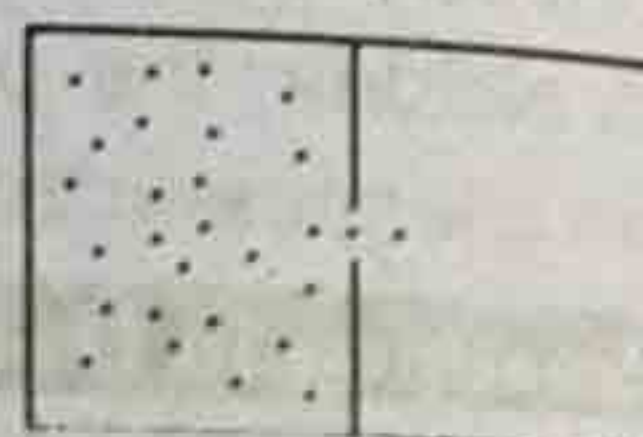


Figure: Effusion

Why gases effuse?

The escape of gas molecules is not due to collision. It is due to their tendency to escape one by one. Gas molecules are habitual in colliding with the walls of vessel.

When a molecule comes in front of a hole, it enters into the other portion of the vessel. This escape is called effusion.

Exercise Q10 (d):

How do you differentiate between diffusion and effusion? Explain Graham's law of diffusion.

Difference between Diffusion and Effusion of gases

Diffusion	Effusion
1 The spontaneous mixing of molecules of different gases is called diffusion.	1 It is the escape of gas molecules, one by one, without collision, through a hole of molecular size, into a region of low pressure.
2 In this, gas molecules move from an area of its higher concentration to an area of its lower concentration.	2 It is due to the colliding habit of molecules with the wall of the container. During collision, when molecules come in front of a hole, they escape.
3 Gases diffuse by random motion and collision.	3 Gases effuse without collision.
4 Diffusion takes place in all directions.	4 Effusion takes place through a hole of molecular size.
5 Example: The smell of rose or a scent spreads due to diffusion	5 Example: Escape of gas molecules from punctured tyre.

Graham's Law of Diffusion

It was given by an English Scientist, Thomas Graham (1805 – 1869). The Graham's law states

At constant temperature and pressure, the rate of diffusion (or effusion) of a gas is inversely proportional to the square root of its density or Molecular mass.

Mathematical Expression

$$r \propto \frac{1}{\sqrt{d}} \quad \text{or} \quad r \propto \frac{1}{\sqrt{M}}$$

Thus, lighter gases diffuse more rapidly than heavier gases.

Example:

NH_3 gas (molecular mass 17 g mol^{-1}) diffuses more rapidly than HCl gas (molecular mass = 36.5 g mol^{-1})

Explanation

Consider two gases A & B.

For gas 'A' let its rate of diffusion is r_1 , density ' d_1 ', and molecular mass ' M_1 '.

For gas 'B' let its rate of diffusion is r_2 , density ' d_2 ', and molecular mass ' M_2 '.

Then according to Graham's law

For gas 'A'

$$r_1 \propto \frac{1}{\sqrt{d_1}} \quad \text{or} \quad r_1 = \frac{K}{\sqrt{d_1}} \quad (1)$$

For gas B

$$r_2 \propto \frac{1}{\sqrt{d_2}} \quad \text{or} \quad r_2 = \frac{K}{\sqrt{d_2}} \quad (2)$$

Constant K is same for all gases at same temperature and pressure.

Dividing (1) by (2), we get

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$$

Similarly

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Where M_1 and M_2 are molecular masses of gas A and gas B respectively.

Demonstration or Verification of Graham's law**Experiment**

- Take a 100 cm. long glass tube, opened at both ends.
- Plug a cotton piece, soaked in NH_3 solution at one end and a cotton piece, soaked in HCl solution at other end.
- The vapours of both these gases escape into the glass tube simultaneously.
- The vapours of both NH_3 & HCl are invisible. But when they meet with each other, they produce white fumes of NH_4Cl .



- Determine the distance travelled by NH_3 and HCl .

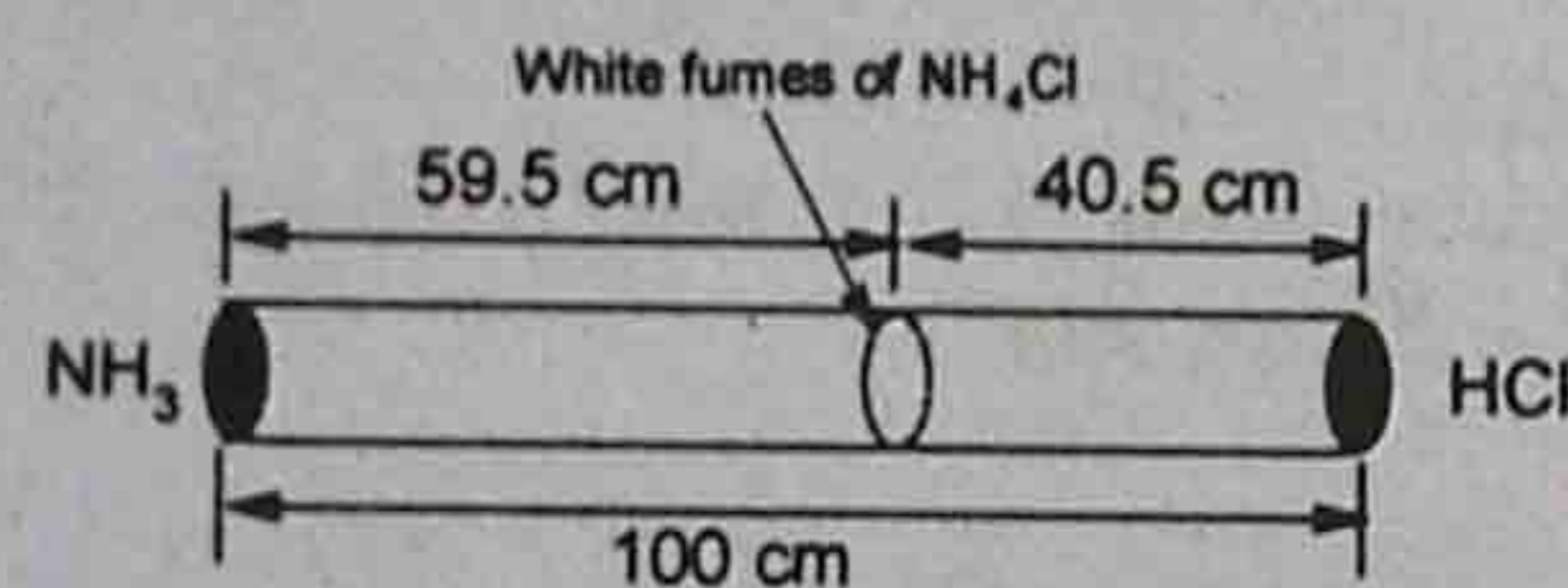


Figure: Verification of Graham's law of diffusion

Calculations

Molecular mass of $\text{NH}_3 = 17 \text{ g mol}^{-1}$

Molecular mass of $\text{HCl} = 36.5 \text{ g mol}^{-1}$

According to Graham's Law

$$\frac{r_{\text{NH}_3}}{r_{\text{HCl}}} = \sqrt{\frac{M_{\text{HCl}}}{M_{\text{NH}_3}}}$$

$$\text{Thus } \frac{59.5}{40.5} = \sqrt{\frac{36.5}{17}}$$

$$1.46 = 1.46$$

Hence, Graham's law is verified

Example 7

250 cm^3 of the sample of hydrogen effuses four times as rapidly as 250 cm^3 of an unknown gas. Calculate the molar mass of unknown gas.

(Cranwala Board, 2014; Sargodha Board, 2014)

Solution:

Rate of effusion unknown gas $= r_x = 1$

Rate of effusion of H_2 gas $= r_{\text{H}_2} = 4$

Molar mass of H_2 gas $= M_{\text{H}_2} = 2 \text{ g mol}^{-1}$

Molar mass of unknown gas $= M_x = ?$

According to Graham's law

$$\frac{r_{\text{H}_2}}{r_x} = \sqrt{\frac{M_x}{M_{\text{H}_2}}}$$

Taking square on both sides

$$\left(\frac{4}{1}\right)^2 = \left(\sqrt{\frac{M_x}{2}}\right)^2$$

$$\frac{M_x}{2} = \frac{16}{1}$$

$$M_x = 16 \times 2 = \boxed{32 \text{ g mol}^{-1}}$$

KINETIC MOLECULAR THEORY OF GASES

To explain the physical behaviour of gases, a theory has been proposed known as Kinetic Molecular Theory of gases.

- Kinetic Molecular Theory was proposed by Bernoulli (1738).
- This theory was used by Clausius (1857) to derive the kinetic equation. He explained all the gas laws with this equation.
- It was further developed by Maxwell, Boltzmann and van der waal.
- Maxwell gave law of distribution of velocities
- Boltzmann studied the distribution of energies among the gas molecules.

The main points of this theory are

- All gases consist of very small particles called molecules. Gases like He, Ne, Ar have monatomic molecules.
- The molecules of gases are widely separated from one another, therefore, these have large empty spaces between them.
- The actual volume of the gas molecules is negligible as compared to the total volume of the gas.
- There are no attractive forces among the gas molecules. Therefore every gas molecule behaves independently.
- The average K.E. of gas molecules is directly proportional to the absolute temperature i.e. $\text{K.E.} \propto T$

- At the same temperature, molecules of every gas have same average kinetic energy.
- Gas molecules are in constant random motion. They change their direction only when they collide with one another or with the walls of container.
- The collisions of gas molecules with one another and with the walls of container are perfectly elastic.
- The pressure of a gas is due to collisions of gas molecules with the walls of container.
- The motion given by force of gravity is negligible as compare to the motion given by collisions.

KINETIC EQUATION FOR AN IDEAL GAS

Kinetic Equation

R.J. Clausius derived the relationship for the pressure of an ideal gas known as kinetic equation. It is given as

$$PV = \frac{1}{3} mN\bar{c}^2$$

where

P = Pressure of gas, V = volume of gas, m = mass of one molecule of gas

N = number of molecules of gas in the container, \bar{c}^2 = mean square velocity

Mean Square Velocity

Definition

The mean of squares of all the possible velocities is called mean square velocity.

If there are n_1 molecules of a gas with velocity c_1 , n_2 with velocity c_2 and so on, then the mean square velocity is given by

$$\text{Mean square velocity} = \bar{c}^2 = \frac{n_1 c_1^2 + n_2 c_2^2 + n_3 c_3^2 + \dots}{n_1 + n_2 + n_3 + \dots}$$

Where $n_1 + n_2 + n_3 + \dots = N$ = total number of molecules.

Why mean square velocity is used?

Under the given conditions, molecules of a gas do not have same velocities. Actually, different velocities are distributed among the molecules. Therefore, mean square velocity of the molecules is taken.

Root mean square velocity

Definition

The square root of mean square velocity is called root mean square velocity denoted by c_{rms} . The value of c_{rms} has been calculated by using kinetic eq. It is given by

$$c_{rms} = \sqrt{\frac{3RT}{M}}$$

where, M = molecular mass of gas T = Absolute Temperature

This equation shows that higher the temperature greater the velocities of the gas molecules.

EXPLANATION OF GAS LAWS ON THE BASIS OF KINETIC MOLECULAR THEORY

1. Boyle's Law

Kinetic eq. for an ideal gas is

$$PV = \frac{1}{3} mN\bar{c}^2 \quad (1)$$

According to kinetic molecular theory of gases, the kinetic energy of gas molecules i.e., $\frac{1}{2} mN\bar{c}^2$ is directly proportional to the absolute temperature, 'T'.

$$\frac{1}{2} mN\bar{c}^2 \propto T$$

$$\frac{1}{2} mN\bar{c}^2 = K T \quad (2)$$

Multiply and divide R.H.S. of eq. (1) by 2, we get

$$PV = \frac{2}{2} \times \frac{1}{3} mN\bar{c}^2$$

or

$$PV = \frac{2}{3} \times \left(\frac{1}{2} mN\bar{c}^2 \right) \quad (3)$$

Putting eq. (2) in eq. (3)

$$PV = \frac{2}{3} KT$$

If T = constant

Then $PV = k'$

This is Boyle's law. It shows that at constant temperature, the volume of a given mass of a gas is inversely proportional to the pressure exerted on it.

2. Charles's Law

Kinetic eq. for an ideal gas is

$$PV = \frac{1}{3} mN\bar{c}^2$$

According to kinetic molecular theory of gases, the average kinetic energy of gas molecules i.e., $\frac{1}{2} mN\bar{c}^2$ is directly proportional to the absolute temperature, 'T'.

$$\frac{1}{2} mN\bar{c}^2 \propto T$$

$$\frac{1}{2} mN\bar{c}^2 = K T \quad (2)$$

Multiply and divide R.H.S. of eq. (1) 2, we get

$$PV = \frac{2}{2} \times \frac{1}{3} mN\bar{c}^2$$

$$PV = \frac{2}{3} \times \left(\frac{1}{2} mN\bar{c}^2 \right) \quad (3)$$

Putting eq. (2) in eq. (3)

$$PV = \frac{2}{3} KT$$

Rearranging, we get

$$V = \frac{2K}{3P} \times T$$

If P = constant then

$$V \propto T$$

$$V \propto T$$

This is Charles's law. It shows that at constant pressure, the volume of a given mass of gas is directly proportional to the absolute temperature.

3. Avogadro's Law

Consider equal volume of two gases under the same conditions of temperature and pressure.

Let For gas 1, mass is m_1 , velocity c_1 and number of molecules N_1 ,

and For gas 2, mass is m_2 , velocity c_2 and number of molecules N_2

Kinetic equations for the two gases will be

$$PV = \frac{1}{3} m_1 N_1 \bar{c}_1^2 \quad \text{for gas 1}$$

$$PV = \frac{1}{3} m_2 N_2 \bar{c}_2^2 \quad \text{for gas 2}$$

Since P & V are same for both gases

Therefore

$$\frac{1}{3} m_1 N_1 \bar{c}_1^2 = \frac{1}{3} m_2 N_2 \bar{c}_2^2$$

or

$$m_1 N_1 \bar{c}_1^2 = m_2 N_2 \bar{c}_2^2 \quad (1)$$

For both gases, at the same temperature, the average K.E. per molecule is also same.

$$\text{Thus } (K.E.)_1 = (K.E.)_2$$

$$\frac{1}{2} m_1 \bar{c}_1^2 = \frac{1}{2} m_2 \bar{c}_2^2$$

or

$$m_1 \bar{c}_1^2 = m_2 \bar{c}_2^2 \quad (2)$$

Dividing eq (1) by (2), we get

$$\frac{m_1 N_1 \bar{c}_1^2}{m_1 \bar{c}_1^2} = \frac{m_2 N_2 \bar{c}_2^2}{m_2 \bar{c}_2^2}$$

$$N_1 = N_2$$

Thus, equal volume of both gases under the same conditions of temperature and pressure, contain equal number of molecules, which is Avogadro's law.

4. Graham's Law of Diffusion

Kinetic equations for the two gases will be

$$PV = \frac{1}{3} mN\bar{c}^2 \quad (1)$$

For 1 mole of gas $N = N_A$ and $N_A m = M$ = Molar mass of gas, therefore

$$PV = \frac{1}{3} M \bar{c}^2$$

$$\bar{c}^2 = \frac{3PV}{M}$$

Taking square root

$$\sqrt{\bar{c}^2} = \sqrt{\frac{3PV}{M}}$$

$$\sqrt{\bar{c}^2} = \sqrt{\frac{3P}{M/V}} = \sqrt{\frac{3P}{d}} \quad \text{since } M/V = d = \text{density}$$

Root mean square velocity is actually the rate of diffusion of gas. Therefore, at constant pressure

$$r = \sqrt{\frac{3P}{d}}$$

$$r \propto \sqrt{\frac{1}{d}}$$

It shows that at constant temperature and pressure, the rate of diffusion (or effusion) of a gas is inversely proportional to the square root of its density, which is the Graham's law.

KINETIC INTERPRETATION OF TEMPERATURE

Consider the kinetic eq

$$PV = \frac{1}{3} mN \bar{c}^2 \quad (1)$$

Where

m = mass of one molecule of gas N = number of molecules

\bar{c}^2 = mean square velocity P = pressure of gas

V = volume

For one molecule the K.E. due to translational motion is given as

$$E_k = \frac{1}{2} m \bar{c}^2 \quad (2)$$

where

E_k = average translational K.E. of molecules

Multiply and divide R.H.S. of eq. (1) by 2, we get

$$PV = \frac{2}{3} \times \frac{1}{2} mN \bar{c}^2$$

$$PV = \frac{2}{3} N \times \left(\frac{1}{2} m \bar{c}^2 \right) \quad (3)$$

Put eq (2) in (3)

$$PV = \frac{2}{3} \times N E_k$$

Let $N = N_A$ = Avogadro's Number

$$PV = \frac{2}{3} \times N_A E_k \quad (4)$$

According to general gas equation for 1 mole of a gas

$$PV = RT \quad (5)$$

Compare (4) and (5)

$$\frac{2}{3} \times N_A E_k = RT$$

$$\text{or } E_k = \frac{3R}{2N_A} \times T$$

Thus there is a direct relationship between translational K.E. and absolute temperature.

Conclusions

- The temperature of a gas is directly proportional to the average translational K.E. of its molecules.
- Thus, a change in temperature changes the motion of molecules of a gas.
- The process of heat transfer from hot body to cold body is due to collisions of molecules with each other. During collisions, hot molecules transfer their K.E. to colder molecules until the average translational K.E. of the two bodies become equal. Thus, the T of two bodies becomes equal.
- When $T = 0$ then $E_k = 0$ i.e. motion of molecules stops, which is not possible. This temperature is called absolute zero. It can never be achieved. The lowest temperature attained so far is 10^{-8} K.
- In gases and liquids, temperature is the measure of average translational K.E. but in solids, it is the measure of only vibrational K.E.

LIQUEFACTION OF GASES

General Principle of Liquefaction

The conversions of gases into liquids require high pressure and low temperature

- High P brings the molecules close to each other
- Low temperature decreases the K.E. of molecules. Thus, attractive forces are developed among them and gas is converted into liquid.

Critical Temperature

The highest temperature at which a substance can exist as a liquid is called critical temperature.

It is denoted by T_c

Critical Pressure

The minimum pressure required to liquefy the gas at the critical temperature is called the Critical pressure

It is denoted by P_c

Critical Volume

The volume occupied by one mole of gas at T_c and P_c is called Critical Volume.

It is denoted by V_c

Example:

- For CO_2 $T_c = 31.1^\circ\text{C}$, $P_c = 72.9 \text{ atm}$, $V_c = 95.65 \text{ cm}^3 \text{ mol}^{-1}$

Importance of critical temperature

- For every gas, there is certain temperature above which a gas cannot be liquefied. This temperature is called critical temperature.

Examples: T_c for O_2 is 154.4 K (-118.75°C). Similarly, T_c for CO_2 is 31.1°C . So, these gases can be liquefied only below their T_c . Hence, these must be cooled below their T_c and then pressure is applied to liquefy them.

Effect of Polarizability on critical temperature

- Non polar gases have low polarizability and have a very low T_c e.g. T_c of Ar is 150.9 K (-122.26°C).
- Polar gases have high polarizability and comparatively high T_c e.g. T_c of NH_3 is 405.6 K (132.44°C).
- Thus, polar gases are easily converted into liquids.

Table 3.2 Critical Temperatures and Critical Pressures of some Substances

Substance	Critical Temperature T_c (K)	Critical Pressure P_c (atm)
Water vapours, H_2O	647.6 (374.44°C)	217.0
Ammonia, NH_3	405.6 (132.44°C)	111.5
Freon-12, CCl_2F_2	384.7 (111.54°C)	39.6
Carbon dioxide, CO_2	304.3 (31.142°C)	73.0
Oxygen, O_2	154.4 (-118.75°C)	49.7
Argon, Ar	150.9 (-122.26°C)	48
Nitrogen, N_2	126.1 (-147.06°C)	33.5

Methods for Liquefaction of Gases

Various methods are used for the liquefaction of gases. These methods are generally based upon Joule Thomson effect.

Joule Thomson Effect

When a compressed gas is allowed to expand suddenly, it produces cooling. This is called Joule - Thomson effect.

Reason: In a compressed gas, molecules are very close to each other and have attractive forces. When a gas is expanded suddenly molecules move away from each other. This process requires energy, which is obtained from the gas itself, hence it is cooled.

Lind's Method

- It is based upon Joule Thomson effect.
- Lind liquefied air by this process.
- The compressed air (about at 200 atm) is passed through a water-cooled pipe where the heat of compression is removed.
- This compressed air is then passed through a spiral tube having a jet at the end. When the gas comes out of jet into low pressure area (1 atm), it suddenly expands and is cooled due to Joule-Thomson effect.
- The cooled air moves up, cools the incoming gas of the Jet, and then again enters into the compression pump, to be compressed again.
- By repeating compression and expansion again and again, air is liquefied.
- All gases except H_2 and He can be liquefied by this process.

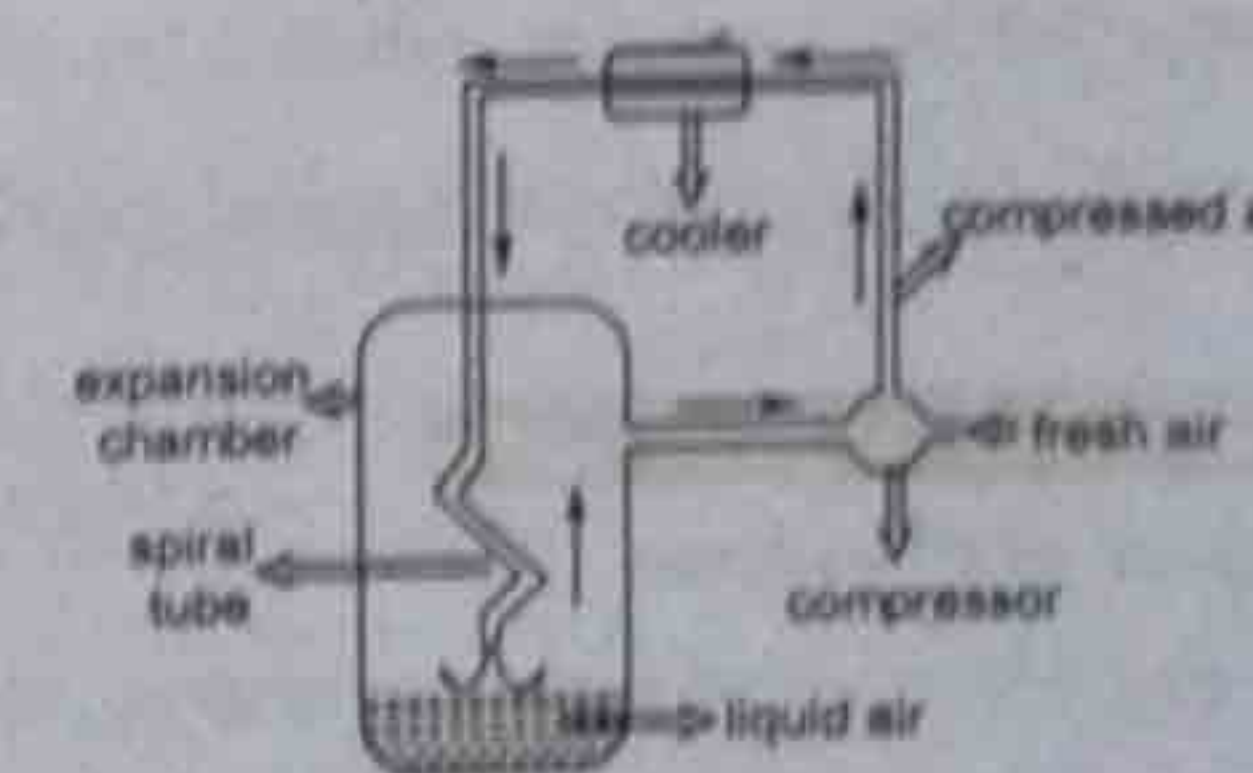


Figure: Lind's method for the liquefaction of air

Exercise Q.13. (a):

Gases show non-ideal behaviour at low temperature and high pressure. Explain this with the help of a graph.

NON-IDEAL BEHAVIOUR OF GASES

Definitions:

The gas which obeys gas laws under all conditions of temperature and pressure is called an "Ideal Gas."

While

The gas which does not obey gas laws under all conditions of temperature and pressure is called a "Real Gas" or a "Non-Ideal Gas"

Compressibility factor:

The effect of temperature and pressure on behaviour of gases can be studied in terms of a compressibility factor, Such that for 1 mole of a gas

$$PV = nRT$$

$$\text{or } \frac{PV}{nRT} = \text{compressibility factor}$$

Behaviour of Ideal gas

Since for an ideal gas,

$$PV = nRT.$$

$$\text{or } \frac{PV}{nRT} = 1.$$

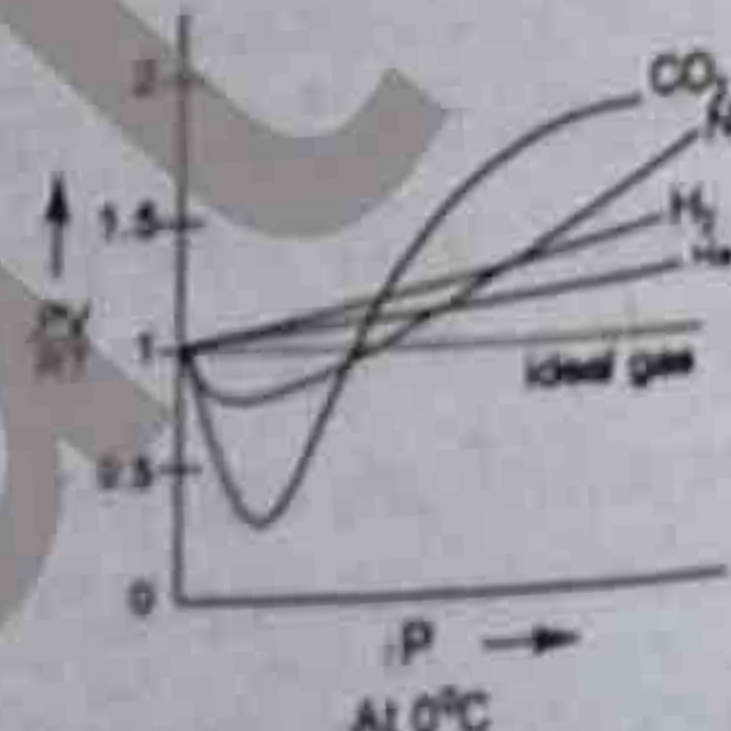
Thus value of compressibility factor is '1', for an ideal gas at all temperature and pressure.

Thus, for an ideal gas, when a graph is plotted between $\frac{PV}{nRT}$ on Y-axis and P on X-axis, a straight line is obtained parallel to pressure axis.

Behaviour of Real gases

Effect of Pressure:

- For He, at low pressure, the curve starts somewhat along the ideal line. However, at very high P, value of $\frac{PV}{nRT}$ has increased from the expected value and He show deviation from ideal behaviour.
- For H₂, deviation starts even at low pressure as compare to He



Fig(3.12a) Non-ideal behaviour of gases at 0°C

- For N₂, $\frac{PV}{nRT}$ first decreases below ideal line and then increases
- CO₂ also shows unusual behaviour.

Thus, it shows that deviations depend upon the nature of gas.

Effect of Temperature:

At high temperature, the graph of these real gases come closer to the ideal line. Thus, the gases become ideal.

Hence

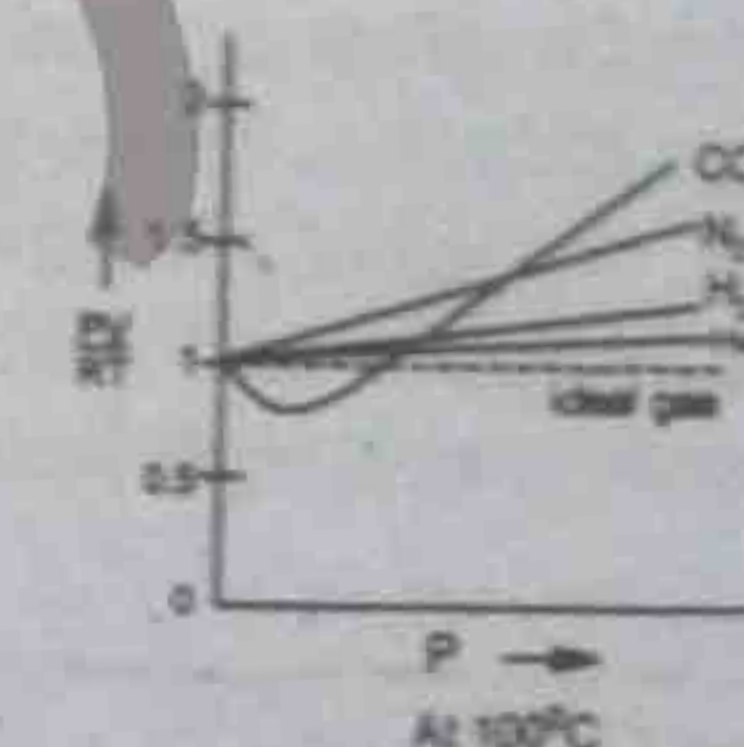


Figure: Non-ideal behaviour of gases at 100°C

Conditions for Ideal and Non-Ideal behaviour of gases

- Gases are non-ideal at high pressure and low temperature.
- Gases are ideal at low pressure and high temperature.

Ideal Gas		Non-ideal gas	
1	It obeys gas laws under all conditions of temperature and pressure	1	It does not obey gas laws under conditions of temperature and pressure
2	There are no attractive or repulsive forces among the gas particles	2	There are attractive and repulsive forces among the gas particles particularly at low T & high P.
3	Actual volume of gas particles is negligible as compared to the total volume of the gas.	3	Actual volume of gas particles is not negligible as compared to the total volume of gas particularly at low 'T' & high 'P'.
4	For an ideal gas $\frac{PV}{RT} = \text{constant}$	4	For an ideal gas $\frac{PV}{RT} = \text{constant}$
5	It is an imaginary gas and does not exist in nature.	5	It is a real gas and exists in nature. e.g. SO ₂ , NH ₃ , etc.

Exercise Q.13. (b):

Do you think that some of the postulates of kinetic molecular theory of gases are faulty? Point out these postulates.

Causes for Deviations from Ideality

Van der Waal (1873) pointed out that two ideal assumptions in the kinetic molecular theory are responsible for these deviations.

- The actual volume of the gas particles is negligible as compared to the total volume of gas.

(ii) There are no attractive forces among the gas particles.

At high pressure and low temperature, gas molecules come closer to each other and develop attractive forces among them. Moreover, at high pressure volume of a gas does not remain negligible. So, the behaviour of real gases is not according to kinetic molecular theory of gases and $PV=nRT$. Thus gases show deviations from ideal behaviour at low temperature and high pressure.

Exercise Q.14 (a):

Derive van der Waal's equation for real gases.

van der Waal's Equation for Real Gases

van der Waal made volume and pressure correction to remove the defects of ideal gas model, and gave an equation of state for real gases

Volume Correction

Van der Waal pointed out that gas molecules have definite volume. Although volume of gas molecules is very small as compare to vessel but it is not negligible. Thus, when pressure is increased on gas molecules, they oppose it. So, if V_{vessel} is the total volume of gas and ' b ' is the volume of gas molecules per mole (excluded volume), then the volume available for compression will be $(V_{\text{vessel}} - b)$ and not V .

$$\text{Hence } V_{\text{free}} = V_{\text{vessel}} - b \quad (1)$$

Thus V_{free} is the volume available to gas molecules.

Where constant ' b ' is the characteristic of each gas. Its value is

$$b = 4V_m \quad \text{where } V_m \text{ is the molar volume}$$

' b ' is the effective volume or excluded volume. It is the volume occupied by gas molecules in highly compressed state but not in liquid state.

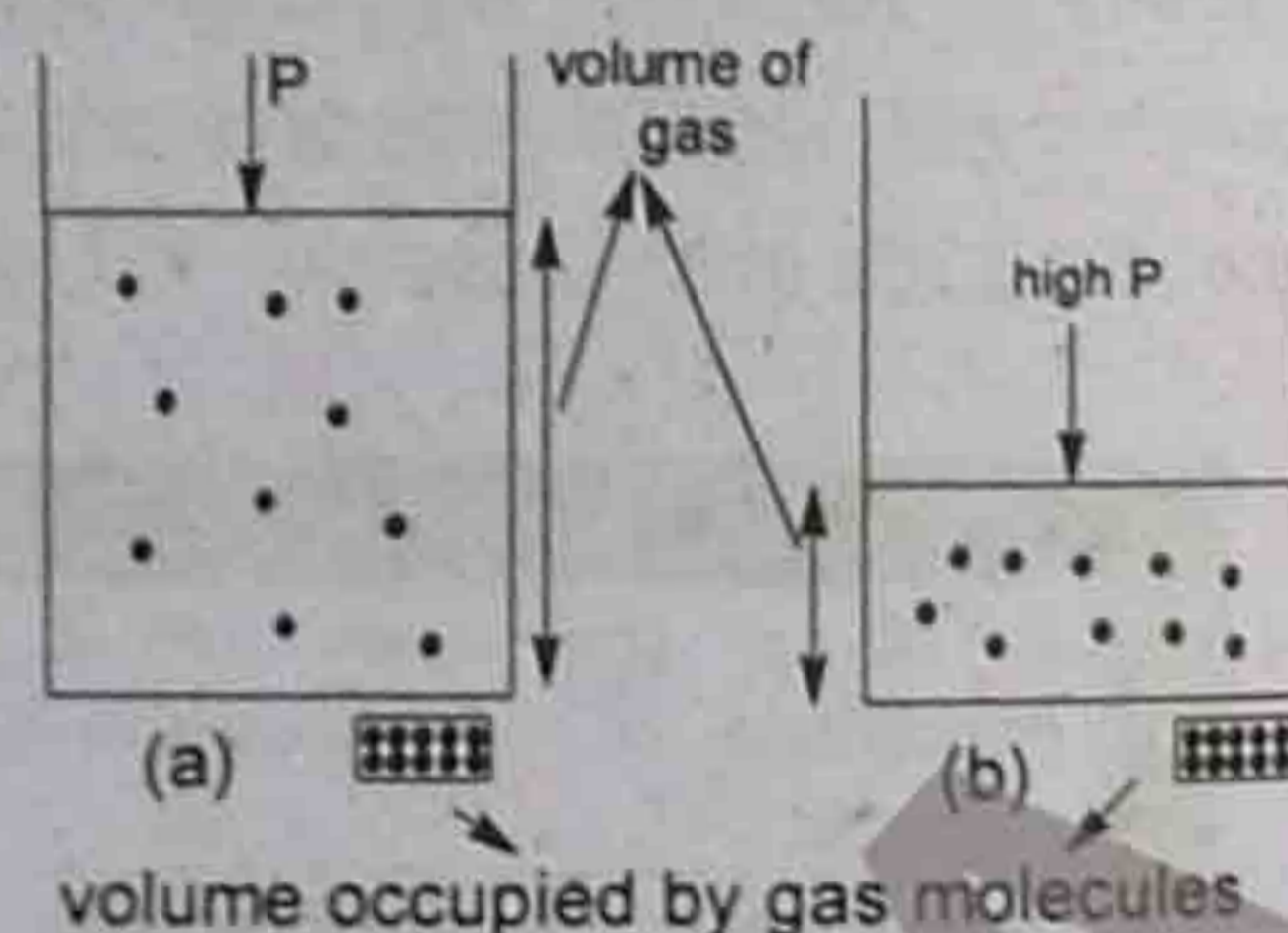


Figure:
(a) A gas at low pressure when actual volume is negligible
(b) A gas at high pressure when actual volume is not negligible

Pressure Correction

A molecule in the interior of gas has no unbalanced force. It is because it is attracted equally from all sides. However, a molecule just striking the wall of the vessel is more attracted inward. Therefore, it will not strike the wall with full force and hence observed pressure ' P ' will be less than the ideal pressure ' P_i ' by an amount P' .

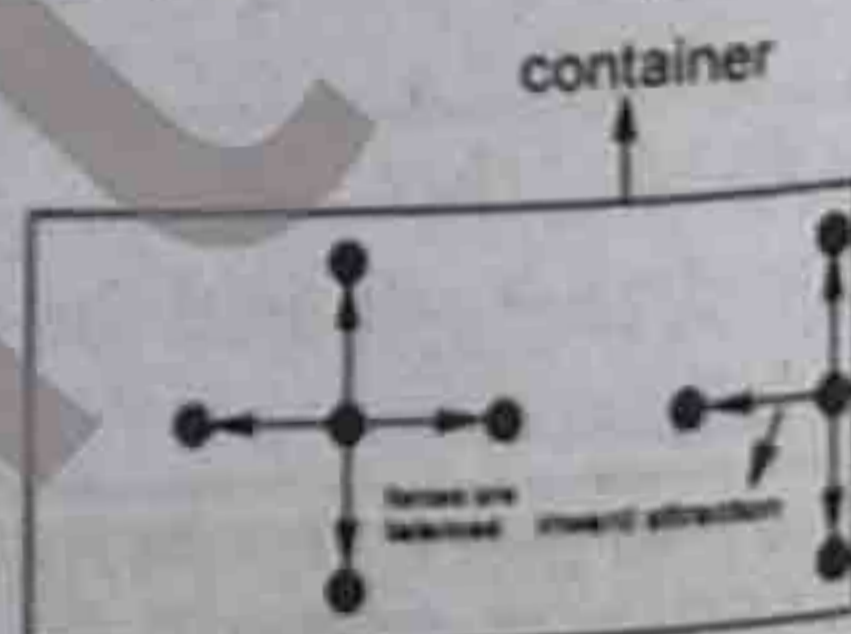


Figure: Forces of attraction and pressure correction.

$$\text{i.e. } P = P_i - P'$$

$$\text{or } P_i = P + P' \quad (2)$$

Consider two types of molecules 'A' and 'B'. Let molecule of type 'A' strikes the wall then it will be attracted inward by molecules of type 'B'. Therefore, P' depends upon the forces of attractions between 'A' and 'B'. Hence net force of attraction is proportional to the concentrations of A type and B type molecules.

$$\text{i.e. } P' \propto C_A \cdot C_B \quad (3)$$

Let ' n ' is number of moles of A and B present in a total volume ' V ', then concentrations of A and B will be given by

$$C_A = \frac{n}{V} \quad \text{and} \quad C_B = \frac{n}{V} \quad (4)$$

Put eq (4) in eq (3)

$$P' \propto \frac{n}{V} \cdot \frac{n}{V}$$

$$P' \propto \frac{n^2}{V^2}$$

$$P' = \frac{a n^2}{V^2}$$

For 1 mole of a gas

$$P' = \frac{a}{V^2} \quad (5)$$

Where a = co-efficient of attraction or attraction per unit volume

For gases with strong intermolecular forces value of ' a ' is high.

Put eq (5) in eq (2)

$$P_i = P + \frac{a}{V^2} \quad (6)$$

Thus, ideal gas equation for 1 mole will become

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT$$

This equation is called van der Waal's equation of state. This equation is applicable to real gases.

$$\text{For 'n' mole of a gas} \quad \left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

Examples:

- For H_2 gas $b = 0.0266 \text{ dm}^3 \text{ mol}^{-1}$. It means 1 mole of H_2 gas (2.016 g) occupies 0.0266 dm^3 at closet approach in gaseous state.
- For H_2 gas ' a ' is the least (0.245) due to its non-polar character and small size.

Units of van der Waals constants, ' a ' and ' b '

- SI units

Units of b : $\text{m}^3 \text{ mol}^{-1}$ Units of a :

$$a = \frac{P'V^2}{n^2} = \frac{Nm^{-2}(\text{m}^3)^2}{\text{mol}^2} = \frac{Nm^{-2}m^6}{\text{mol}^2} = Nm^4 \text{ mol}^{-2}$$

- Non-SI units

Units of b : $\text{dm}^3 \text{ mol}^{-1}$ Units of a :

$$a = \frac{P'V^2}{n^2} = \frac{\text{atm}(\text{dm}^3)^2}{\text{mol}^2} = \text{atm dm}^6 \text{ mol}^{-2}$$

- The value of ' a ' and ' b ' can be determined by noting P , V and T under two different conditions.

Table 3.3 van der Waal's constants for some common gases		
Gas	' a ' (atm dm ⁶ mol ⁻²)	' b ' (dm ³ mole ⁻¹)
Hydrogen	0.245	0.0266
Oxygen	1.360	0.0318
Nitrogen	1.390	0.0391
Carbon dioxide	3.590	0.0428
Ammonia	4.170	0.0371
Sulphur dioxide	6.170	0.0564
Chlorine	6.439	0.0562

Example 8:

One mole of methane gas is maintained at 300 K. Its volume is 250 cm^3 . Calculate the pressure exerted by the gas under the following conditions.

(i) When the gas is ideal

(ii) When the gas is non-ideal

$$a = 2.253 \text{ atm dm}^6 \text{ mol}^{-2}, b = 0.0428 \text{ dm}^3 \text{ mol}^{-1}$$

(Sargodha Board, 2012; Gujranwala Board, 2012; Lahore Board, 2014)

Solution:

(i) When the gas is ideal, general gas equation is applied

$$P = ?$$

$$V = 250 \text{ cm}^3 = 0.25 \text{ dm}^3$$

$$n = 1 \text{ mole}$$

$$T = 300 \text{ K}$$

$$R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$$

According to general gas eq.

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

$$P = \frac{1 \times 0.0821 \times 300}{0.25}$$

$$P = 98.5 \text{ atm}$$

(ii) When the gas is behaving as non-ideal, we should use the van der Waal's equation.

$$P = ?$$

$$n = 1 \text{ mole}$$

$$R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mole}^{-1}$$

$$V = 0.25 \text{ dm}^3$$

$$T = 300 \text{ K}$$

$$a = 2.253 \text{ atm dm}^6 \text{ mol}^{-2}$$

$$b = 0.0428 \text{ dm}^3 \text{ mol}^{-1}$$

According to van der Waal's equation

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

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

$$\text{or } P = \frac{nRT}{(V - nb)} - \frac{an^2}{V^2}$$

Putting values

$$P = \frac{1 \times 0.0821 \times 300}{0.25 - 1(0.0428)} - \frac{2.253 \times 1^2}{(0.25)^2}$$

$$P = \frac{24.63}{0.207} - \frac{2.253}{0.0625}$$

$$P = 118.99 - 36.05$$

$$P = \boxed{82.94 \text{ atm}}$$

Thus pressure lessened from ideal behaviour = $98.5 - 82.94 = 15.56 \text{ atm}$

PLASMA STATE

A mixture of neutral particles, positive ions and negative electrons is called Plasma.

Formation of Plasma

- On heating a solid, it is converted into liquid. On further heating, the liquid is converted into vapours. Thus, the phase of matter changes from solid to liquid and then liquid to vapours.
- Now if vapours are further heated some of them lose electrons and positive ions are formed. Hence a mixture of neutral particles, positive ions and negative electrons are produced. This is called Plasma.
- The ionization is produced by high temperature or by radiations.



FIGURE: Plasma

Occurrence of Plasma or Where Plasma Found

- About 99% of the universe is made up of Plasma.
- It is present everywhere in sun and stars
- The sun is a 1.5 million km ball of plasma. It is heated by nuclear fusion.
- It is the most abundant form of matter. It is the stuff of stars
- It is present in everything from sun to quarks (quark is the smallest particle of universe)
- Majority of the matter in inter-stellar space is plasma
- All the shining stars are plasma
- On earth, it is very limited. It is found in lightning bolts, flames, auroras and fluorescent lights, neon signs etc.
- When an electric current is passed through neon gas, it produces both plasma and light.

Characteristics of Plasma

- It consists of a significant number of charge particles. Thus it respond to both electric and magnetic fields
- Motion of particles in the plasma produces fields and electric current within plasma density. It refers to the density of charged particles.
- Plasma has a complex set of interactions. It is a unique fascinating and complex state of matter.
- It is macroscopically neutral. Although it contains ions and electrons but their number is equal.

Natural and Artificial Plasma

Artificial Plasma

- It is produced by using electrical charges on a gas e.g. in neon signs
- Plasma at low T is hard to maintain. It is because outside a vacuum, low T plasma reacts rapidly with any molecule. Thus, it is both useful and hard to use.

Natural Plasma

- It only exists at high temperature or low temperature vacuums. It does not break down or react rapidly
- These are extremely hot (over 20,000°C minimum)
- They have so much energy, that they can vaporize any material.

Applications

- Since plasma can respond to both electric and magnetic fields, it can have many uses.
- A fluorescent bulb is different from regular light bulbs. It consists of a long tube filled with gas. When electricity is passed through gas, it charges up the gas. The charging and exciting of gas creates glowing plasma inside the bulb.
 - Neon signs are glass tube filled with gas. When electricity is passed through the tube, it charges the gas and creates glowing plasma inside the tube. The colour of plasma depends upon the gas used.
 - They are used for plasma processing of semiconductors, sterilization of some medical products, lamps, lasers, diamond coated films, high power microwave sources and pulsed power switches.
 - They are helpful for generation of electricity from fusion pollution control and removal of hazardous chemicals. Thus, it helps to clean up the environment.
 - Plasma light up our offices, homes. It helps in working of computers, electronic equipment.
 - It drives lasers and particle accelerators.
 - It can be used to pasteurize food.
 - It is used to make corrosion resistant tools.

Future Horizons

Scientists are working to use plasma effectively

For effective use

- it should have low energy
- It should survive for sufficient time without reacting and degeneration.

The applications of magnetic field involve the use of plasma. The magnetic fields create low energy plasma, which create metastable state of molecules. These will then react with another molecule with right energy. Thus these molecules can survive long enough to react with desired molecules. These metastable particles are selective in their reactivity. Thus, these can be used to solve problems like radioactive contamination.

Scientists are working on mixture of gases to use as metastable agents on plutonium and uranium.

OBJECTIVE AND SHORT ANSWER QUESTIONS (Exercise)**Q.1. Multiple-choice questions.**

- (i) Pressure remaining constant, at which temperature the volume of a gas will become twice of what it is at 0°C .
 (a) 546°C (b) 200°C (c) 546 K (d) 273 K

(Bahawalpur Board, 2009) 2011 B (D.G. Khan Board, 2012) (Sargodha Board, 2012) (Gujranwala board, 2010, 2013, 2014)

- (ii) Number of molecules in one litre of water is close to:

- (a) $\frac{6.02}{22.4} \times 10^{23}$ (b) $\frac{12.04}{22.4} \times 10^{23}$ (c) $\frac{18}{22.4} \times 10^{23}$ (d) $55.6 \times 6.02 \times 10^{23}$

(D.G. Khan Board, 2010) (D.G. Khan Board, 2012) (Rawalpindi board, 2012) (Lahore board, 2013) (Gujranwala board, 2012)

- (iii) Which of the following will have the same number of molecules at STP?

- (a) 280 cm^3 of CO_2 and 280 cm^3 of N_2O (b) 11.2 cm^3 of O_2 and 32 g of O_2
 (c) 44 g of CO_2 and 11.2 dm^3 of CO (d) 28 g of N_2 and 5.6 dm^3 of Oxygen

(Faisalabad Board, 2008)

- (iv) If absolute temperature of a gas is doubled and the pressure is reduced to one half, the volume of the gas will

- (a) remain unchanged (b) increase four times
 (c) reduce to $\frac{1}{4}$ (d) be doubled

(Rawalpindi Board, 2010) (Multan Board, 2011, 2012)

- (v) How should the conditions be changed to prevent the volume of a given gas from expanding when its mass is increased?

- (a) Temperature is lowered and pressure is increased
 (b) Temperature is increased and pressure is lowered
 (c) Temperature and pressure both are lowered
 (d) Temperature and pressure both are increased

- (vi) The molar volume of CO_2 is maximum at

- (a) STP (b) 127°C and 1 atm
 (c) 0°C and 2 atm (d) 273°C and 2 atm

(D.G. Khan Board, 2009) (Rawalpindi Board, 2009) (Lahore Board, 2010) (Sargodha Board, 2010, 2012, 2013, 2014)
 (Gujranwala board, 2014)

- (vii) The order of the rate of diffusion of gases NH_3 , SO_2 , Cl_2 and CO_2 is:

- (a) $\text{NH}_3 > \text{SO}_2 > \text{Cl}_2 > \text{CO}_2$ (b) $\text{NH}_3 > \text{CO}_2 > \text{SO}_2 > \text{Cl}_2$
 (c) $\text{Cl}_2 > \text{SO}_2 > \text{CO}_2 > \text{NH}_3$ (d) $\text{NH}_3 > \text{CO}_2 > \text{Cl}_2 > \text{SO}_2$

(Faisalabad Board, 2007, 2008) (D.G. Khan Board, 2012) (Gujranwala board, 2009)

- (viii) Equal masses of methane and oxygen are mixed in an empty container at 25°C . The fraction of total pressure exerted by oxygen is

- (a) $\frac{1}{3}$ (b) $\frac{8}{9}$ (c) $\frac{1}{9}$ (d) $\frac{16}{17}$

(Multan Board, 2010) (Rawalpindi board, 2012) (Gujranwala board, 2012)

(ix) Gases deviate from ideal behaviour at high pressure. Which of the following is correct for non-ideality?

- (a) At high pressure, the gas molecules move in one direction only
- (b) At high pressure, the collision between the gas molecules are increased manifold
- (c) At high pressure, the volume of the gas becomes insignificant
- (d) At high pressure, the intermolecular attraction become significant

(x) The deviation of a gas from ideal behaviour is maximum at

- (a) -10°C and 5.0 atm
- (b) -10°C and 2.0 atm
- (c) 100°C and 2 atm
- (d) 0°C and 2 atm

(Lahore Board, 2008; D.G. Khan Board, 2012; Multan Board, 2012)

(xi) A real gas obeying van der Waals equation will resemble ideal gas if

- (a) both a and b are large
- (b) both a and b are small
- (c) a is small and b is large
- (d) a is large and b is small

(Bahawalpur Board, 2010)

ANSWERS TO MULTIPLE CHOICE QUESTIONS

(i) Ans: (c)

Let $V_1 = 1$ $T_1 = 0^{\circ}\text{C} + 273 = 273\text{ K}$
To double the volume
 $V_2 = 2$ $T_2 = ?$
According to Charles's law,
 $T_2 = \frac{V_2 \times T_1}{V_1} = \frac{2 \times 273}{1} = 546\text{ K}$

(iii) Ans: (a)

Avogadro's law states
"Equal volumes of all gases at the same temperature and pressure contain equal number of molecules"

Since CH_4 and NH_3 have same volume at STP, therefore, they will have same number of molecules.

(iv) Ans: (b)

Increase in mass of gas increases the volume of gas. Therefore, to keep volume constant, pressure must be increased or temperature must be decreased or both.

(v) Ans: (d)

According to Graham's law, rate of diffusion of a gas is inversely proportional to the square root of its molecular mass.

$$\text{i.e. } r \propto \frac{1}{\sqrt{M}}$$

Since Molecular masses of NH_3 , SO_2 , Cl_2 and CO_2 are

(ii) Ans: (d)

Mass of 1 dm^3 of water = $1\text{ kg} = 1000\text{ g}$
Molecular mass of water = 18 g mol^{-1}

$$\text{Moles of water} = \frac{1000}{18} = 55.55\text{ moles}$$

1 mole of water contains 6.02×10^{23} molecules
55.55 moles will contain $55.55 \times 6.02 \times 10^{23}$ molecules

(vi) Ans: (b)

According to gas laws, volume is directly proportional to T , and inversely proportional to P

$$\text{i.e. } V \propto \frac{T}{P}$$

Thus if T is doubled and P is reduced to one-half, then volume will increase four times.

$$V \propto \frac{2}{1/2} = 4$$

(vii) Ans: (b)

Since $V \propto \frac{T}{P}$

In (c) and (d), pressure is high, therefore, volume will be less.

In (a) and (b), pressure is low, but in (b) temperature is also high. Hence molar volume of CO_2 will be maximum in (b).

(viii) Ans: (a)

According to Dalton's law of partial pressure, partial pressure of a gas in a mixture of gases is directly proportional to its mole fraction. $P_i \propto x_i$

Let equal masses of CH_4 and O_2 are 32 g each. Thus it will be 2 moles of CH_4 and 1 mole of O_2 .

in the order, $\text{Cl}_2 > \text{SO}_2 > \text{CO}_2 > \text{NH}_3$.

Therefore, their rate of diffusion will be in the inverse order i.e., $\text{NH}_3 > \text{CO}_2 > \text{SO}_2 > \text{Cl}_2$

Hence mole fraction of $\text{O}_2 = \frac{1}{3}$, and

$$\text{mole fraction of } \text{CH}_4 = \frac{2}{3}$$

Since mole fraction of O_2 is $1/3$, therefore, its pressure will also be $1/3$ of total pressure.

(ix) Ans: (d)

At high P , molecules of gas come closer together. Therefore, they form intermolecular forces. Hence, gas show deviations from ideal behaviour.

(x) Ans: (a)

Gases show deviations from ideal behaviour at low temperature and high pressure.

In (a) temperature is lowest and pressure is highest than in (b), (c) and (d).

(xi) Ans: (b)

'a' measures attraction among the gas molecules. Larger 'a', greater attractions.

'b' measures volume occupied by gas molecules. Larger 'b', greater volume occupied.

For a gas to behave as an ideal gas volume of gas molecules should be negligible and there should be no forces of attractions among gas molecules. Hence, 'a' and 'b' should be small.

Q.2. Fill in the blanks

- (i) The product PV has the S.I. unit of _____.
- (ii) Eight grams each of O_2 and H_2 at 27°C will have total KE in the ratio of _____.
- (iii) Smell of cooking gas during leakage from gas cylinder is detected because of the property of _____.
- (iv) Equal _____ of ideal gases at the same temperature and pressure contain _____ number of molecules.
- (v) The temperature above, which a substance exists only as a gas, is called _____.

Answers:

- (i) Nm
- (ii) 1:16
- (iii) diffusion
- (iv) volume, equal
- (v) critical temperature

Q.3. Label the following sentences as true or false.

- (i) KE of molecules is zero at 0°C .
- (ii) A gas in a closed container will exert much higher pressure at the bottom due to gravity than at the top.
- (iii) Real gases show ideal gas behaviour at low pressure and high temperature.
- (iv) Liquefaction of gases involves decrease in intermolecular spaces.
- (v) An ideal gas on expansion will show Joule-Thomson effect.

Answers:

- (i) False
- (ii) False
- (iii) True
- (iv) True
- (v) False

Q.4. (a) What is Boyle's law of gases? Give its experimental verification.

Solved on Page 99

- (b) What are isotherms? What happens to the positions of isotherms when they are plotted at high temperature for a particular gas?

Solved on Page 100

- (c) Why do we get a straight line when pressures exerted on a gas are plotted against inverse of volumes? This straight line changes its position in the graph by varying the temperature. Justify it.

Solved on Page 101

- (d) How will you explain that the value of the constant K in $(PV=K)$ in Boyle's law depend upon

- (i) Temperature of the gas

According to Boyle's law $PV=K$ at constant temperature. When temperature is increased, volume of a given mass of gas is increased. Therefore, the product PV hence K will also increase.

- (ii) Quantity of the gas

According to Boyle's law $PV=K$ at constant temperature. When quantity of gas is increased, volume of gas is increased at constant temperature. Therefore, the product PV and hence K will also increase.

- Q5 (a) What is the Charles's law? Which scale of temperature is used to verify that $V/T = k$ (pressure and number of moles are constant)?

Solved on Page 100

- (b) A sample of carbon monoxide gas occupies 150 mL at 25°C . It is then cooled at constant pressure until it occupies 100 mL. What is the new temperature?

Solution:

$$V_1 = 150 \text{ mL} \quad V_2 = 100 \text{ mL}$$

$$T_1 = 25^\circ\text{C} + 273 = 298 \text{ K} \quad T_2 = ?$$

According to Charles's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad T_2 = V_2 \times \frac{T_1}{V_1} = 100 \times \frac{298}{150} = 198.7 \text{ K}$$

- (c) Do you think that the volume of any quantity of a gas become zero at -273°C . Is it not against the law of conservation of mass? How do you deduce the idea of absolute zero from this information?

No volume of a gas cannot be zero at -273°C , since mass of gas will be destroyed. It is against the law of conservation of mass which states

Mass can neither be created nor destroyed

Since temperature -273°C is unattainable and it is the lowest temperature. Therefore, it is taken as Absolute zero of Kelvin scale. Thus, absolute zero is defined as

The hypothetical temperature at which the volumes of all gases become zero is called absolute zero.

- Q.6. (a) What is Kelvin scale of temperature? Plot a graph for one mole of an a real gas to prove that a gas becomes liquid, earlier than -273.16°C .

Solved on Page 104

- (b) Throw some light on the factor $1/273$ in Charles's law

According to quantitative statement of Charles's law,

At constant pressure, the volume of a given mass of an ideal gas increases or decreases by $1/273$ of its original volume at 0°C for every 1°C rise or fall in temperature respectively.

- Thus, If we have 273 dm^3 of a gas at 0°C , then its $\frac{1}{273}$ part will be $\frac{1}{273} \times 273 \text{ dm}^3 = 1 \text{ dm}^3$

So, for every 1°C rise or fall in temperature, volume will increase or decrease by 1 dm^3 .

- Also, If we have 546 dm^3 of a gas at 0°C , then its $\frac{1}{273}$ part will be $\frac{1}{273} \times 546 \text{ dm}^3 = 2 \text{ dm}^3$

So, for every 1°C rise or fall in temperature, volume will increase or decrease by 2 dm^3 .

In both cases if temperature is decreased to -273°C , the volume of the gas will become zero. Thus, -273°C is taken as zero of Kelvin scale and is called Absolute zero. Hence, this factor $1/273$ has helped in the development of Kelvin Scale.

- Q.7. (a) What is the general gas equation? Derive it in various forms?

Solved on Page 106

- (b) Can we determine the molecular mass of an unknown gas if we know the pressure, temperature and volume along with the mass of the gas.

Solved on Page 107

- (c) How do you justify from general gas equation that increase in temperature or decrease of pressure decreases the density of the gas?

Solved on Page 108

- (d) Why do we feel comfortable in expressing the densities of gases in the units of g/dm^3 rather than g/cm^3 .

In gases, molecules are widely separated from each other, and small amount of gas is present in a large volume. Therefore, values of densities will be much smaller, if expressed in cm^3 .

Thus, a bigger unit dm^3 is used because appreciable amount of gas will be present per dm^3 of gas and values of densities will be greater. e.g. density of CH_4 at STP is 0.71 g/dm^3 but in g/cm^3 , it is 7.1×10^{-4} which is inconvenient.

Q.8. Derive the units for general gas constant R in general gas equation.

(a) When the pressure is in atmospheres and volume in dm^3

$$\text{Since } R = \frac{PV}{nT}$$

$$\text{Thus } R = \frac{\text{atm} \times \text{dm}^3}{\text{mol} \times \text{K}} = \text{atm dm}^3 \text{ mol}^{-1} \text{K}^{-1}$$

(b) When the pressure is in N/m^2 and volume in m^3

$$\text{Since } R = \frac{PV}{nT}$$

$$\text{Thus } R = \frac{\text{Nm}^{-2} \times \text{m}^3}{\text{mol} \times \text{K}} = \text{Nm mol}^{-1} \text{K}^{-1} = \text{J mol}^{-1} \text{K}^{-1}$$

(c) When energy is expressed in ergs

In SI system, The units of R in terms of energy are $R = \text{J mol}^{-1} \text{K}^{-1}$

In CGS system the energy is expressed in units of erg and

$$1 \text{ J} = 10^7 \text{ erg}$$

Therefore, units of R will be $R = \text{erg mol}^{-1} \text{K}^{-1}$

However, numerical value of R will be different than SI units.

Q.9. (a) What is Avogadro's law of gases?

Solved on Page 111

(b) Do you think that 1 mole of H_2 and 1 mole of NH_3 at 0°C and 1 atm pressure will have Avogadro's number of particles. If not, why?

1 mole of every substance has Avogadro's number of particles.

Thus, 1 mole of each of H_2 and NH_3 will also have Avogadro's number of particles i.e. 6.02×10^{23} particles.

(c) Justify that 1 cm^3 of H_2 and 1 cm^3 of CH_4 at STP will have same number of molecules, when one molecule of CH_4 is 8 times heavier than that of hydrogen. (Calculate CH_4 2002)

In gases distance between two molecules is approximately 300 times its molecular size. Thus volume occupied by gas molecule does not depend upon the molecular size or mass of molecules. The volume of gas only depends upon the number of molecules. Hence according to Avogadro's law equal volume of H_2 and CH_4 at STP will have same number of molecules, although CH_4 molecule is 8 times heavier than H_2 molecule.

Q.10. (a) Dalton's law of partial pressures is only obeyed by those gases, which don't have attractive forces among themselves. Explain it.

Dalton's law is an ideal gas law which assumes that there are no attractive forces among the gas molecules. Hence, every gas molecule moves independently.

On mixing different gases, if they have negligible attractions for each other, then every gas molecule moves independently, exerts its full pressure and thus obeys Dalton's law.

However, if gases develop strong forces for each other on mixing, then their molecules will not be independent in their motion. Therefore, pressure exerted by each gas will be different than expected. Hence, gases with strong attractive forces do not obey Dalton's law.

(b) Derive an equation to find out the partial pressure of a gas knowing the individual moles of component gases and the total pressure of the mixture.

Solved on Page 114

(c) Explain that the process of respiration obeys the Dalton's law partial pressure.

Solved on Page 116

(d) How do you differentiate between diffusion and effusion? Explain Graham's law of diffusion.

Solved on Page 118

Q.11. (a) Gases show non-ideal behaviour at low temperature and high pressure. Explain this with the help of a graph.

Solved on Page 130

(b) Do you think that some of the postulates of kinetic molecular theory of gases are faulty? Point out these postulates.

Solved on Page 131

(c) Hydrogen and helium are ideal at room temperature, but SO_2 and Cl_2 are non-ideal. How do you explain it? (Rawalpindi Board, 2007, 2010; Lahore Board, 2012)

H_2 and He have very low liquefaction temperatures. ($\text{H}_2 = -252.87^\circ\text{C}$ $\text{He} = -268.934^\circ\text{C}$)

While SO_2 and Cl_2 have sufficiently high liquefaction temperatures, which are close to room temperature. ($\text{SO}_2 = -10.2^\circ\text{C}$ $\text{Cl}_2 = -34.6^\circ\text{C}$)

Due to very low liquefaction temperature of H_2 and He, their molecules have negligible attractions for each other at room temperature. Hence, these gases behave ideally at room temperature.

While liquefaction temperature of SO_2 and Cl_2 are close to room temperature, therefore, at room temperature, their molecules have appreciable forces of attraction. Hence SO_2 and Cl_2 are non-ideal at room temperature.

Q.14 (a) Derive van der Waal's equation for real gases.

Solved on Page 132

(b) What is the physical significance of Vander Wall's constants a and b give their units. (Multan Board, 2013)

'a'

In van der Waals equation ' a ' is called as the co-efficient of attraction or attraction per unit volume. It indicates the strength of intermolecular forces in gases. Stronger the intermolecular forces, higher the value of ' a ' and vice versa. Its units are $\text{Nm}^4\text{mol}^{-2}$

'b'

In Van der Waal's eq, ' b ' is called as the effective volume (excluded volume) occupied by gas molecules in highly compressed state but not in liquid state. It depends upon the size of gas molecules. Greater the size of gas molecules greater will be the value of ' b ' and vice versa. Its units are $\text{dm}^3\text{mol}^{-1}$

Q.15: Explain the following facts

(i) The plot of PV versus P is a straight line at constant temperature and with a fixed number of moles of an ideal gas. (D.G. Khan Board, 2012; Azad Kashmir Board, 2012)

According to Boyle's Law

At constant T , the product of pressure and volume of a fixed mass of gas is always constant. i.e. $PV = K$

Hence, if a graph is plotted between P on X -axis and PV on Y -axis, then a straight line parallel to X -axis is obtained, showing that PV (or K) is a constant quantity.

(ii) The straight line in (a) is parallel to x -axis and goes away from the pressure axis at higher pressure.

For real gases, when pressure is increased, the value of PV does not remain constant. At higher pressure, molecules of gases come closer together, develop forces of attraction, and thus show deviation from ideal behaviour. Hence, at high P , the straight line changes to curves and goes away from the pressure axis. The nature of curve depends upon the nature of real gas.

(iii) The van der Waals constant ' b ' of a gas is four times the molar volume of gas.

The gas molecules are incompressible spheres. Thus when gas molecules are packed together like spheres, there remains some empty spaces in between the spheres which are also incompressible.

Hence, the incompressible volume ' b ' is actually more than the actual molar volume of gas molecules. Actually, it is four times than their molar volume.

$$\text{i.e. } b = 4 V_m$$

(iv) Pressure of NH_3 gas at given condition (say 20 atm pressure and room temperature) is less as calculated by Van der Waal's equation than that calculated by general gas equation. (Lahore Board, 2012)

NH_3 is a polar gas. Its molecules have forces of attraction. Thus, it shows non-ideal behaviour. Therefore, molecules of NH_3 are attracted inward while striking the wall of container. Thus, these exert less pressure than expected ideal pressure.

Hence, real pressure of NH_3 as calculated by Van der Waal's equation is less than the ideal pressure.

(v) Water vapours do not behave ideally at 273 K.

(D.G. Khan Board, 2007; Lahore Board, 2012; Gujranwala Board, 2013, 2014)

273 K is the freezing point of water. At this temperature, vapours of water have appreciable forces of attraction between them. Hence water vapours behave non-ideally at 273 K

(vi) SO_2 is non-ideal at 273 K but behaves ideally at 327°C .

(Multan Board, 2008; Lahore Board, 2011, 2013)

At low temperature of 273 K, molecules SO_2 gas (-10.2°C) have considerable attractions for each other and thus SO_2 gas behaves non-ideally.

When temperature is increased to 327 K, the forces of attractions among SO_2 molecules decreases and hence SO_2 gas shows ideal behaviour at 327°C .

IMPORTANT FORMULAS

GAS LAWS

Boyle's Law

$$PV = k \quad P_1V_1 = P_2V_2$$

Charles's Law

$$\frac{V}{T} = k \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

General Gas Equation

$$PV = nRT \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$M = \frac{nRT}{PV} \quad d = \frac{PM}{RT}$$

van der Waal's Equation

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

Avogadro's Law

To find the number of particles (atoms, ions, molecules or formula units)

$$\text{number of particles} = \frac{\text{Given Mass}}{\text{Molar Mass}} \times N_A$$

or number of particles = moles $\times N_A$

Molar Mass may be Molecular mass or atomic mass or formula mass or ionic mass

Dalton's Law of Partial Pressure

$$P_t = P_1 + P_2 + \dots$$

$$P_t V = n_t RT \quad \text{or} \quad P_g V = n_g RT$$

$$P_g = \frac{n_g}{n_t} \times P_t \quad \text{or} \quad P_g = x_g \times P_t$$

P_t = Total Pressure of mixture of gases P_g = Partial Pressure of gas in mixture
 n_g = number of moles of a gas in mixture x_g = Mole Fraction of gas in mixture
 n_t = Total number of moles of a mixture of gases

Graham's Law of Diffusion or Effusion Of Gases

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \quad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Kinetic Equation for an Ideal Gas

$$PV = \frac{1}{3} mN_c \bar{c}^2$$

NUMERICAL PROBLEMS (Exercise)

Q.16. Helium gas in a 100 cm³ container at a pressure of 500 torr is transferred to a container with a volume of 250 cm³. What will be the new pressure (a) if no change in temperature occurs (b) if temperature changes from 20°C to 15°C.

Solution:

(a) When the T is not changed

$$V_1 = 100 \text{ cm}^3 \quad V_2 = 250 \text{ cm}^3$$

$$P_1 = 500 \text{ torr} \quad P_2 = ?$$

According to Boyle's law

$$P_1V_1 = P_2V_2$$

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

$$P_2 = \frac{500 \times 100}{250} = \boxed{200 \text{ torr}}$$

(b) When the T is changed

$$V_1 = 100 \text{ cm}^3 \quad V_2 = 250 \text{ cm}^3$$

$$P_1 = 500 \text{ torr} \quad P_2 = ?$$

$$T_1 = 20^\circ\text{C} + 273 = 293 \text{ K} \quad T_2 = 15^\circ\text{C} + 273 = 288 \text{ K}$$

According to general gas equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

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

$$P_2 = \frac{500 \times 100}{293} \times \frac{288}{250} = \boxed{196.6 \text{ torr}}$$

Q.17. What are the densities in kg/m^3 for following gases at STP ($P=101325 \text{ Nm}^{-2}$, $T=273 \text{ K}$, Molecular mass are in kg mol^{-1}).

- (i) Methane
(ii) Oxygen
(iii) Hydrogen

Solution:

(i) **Methane**

$$T = 273 \text{ K}$$

$$P = 101325 \text{ Nm}^{-2}$$

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

$$= \frac{16}{1000} \text{ kg mol}^{-1} = 0.016 \text{ kg mol}^{-1}$$

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

$$d = ?$$

The density of gas is given by

$$d = \frac{PM}{RT} = \frac{101325 \times 0.016}{8.314 \times 273}$$

$$= \boxed{0.714 \text{ kg m}^{-3}}$$

(ii) **Oxygen**

$$T = 273 \text{ K}$$

$$P = 101325 \text{ Nm}^{-2}$$

$$\text{Molar mass} = M = 32 \text{ g mol}^{-1}$$

$$= \frac{32}{1000} \text{ kg mol}^{-1} = 0.032 \text{ kg mol}^{-1}$$

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

$$d = ?$$

The density of gas is given by

$$d = \frac{PM}{RT} = \frac{101325 \times 0.032}{8.314 \times 273}$$

$$= \boxed{1.428 \text{ kg m}^{-3}}$$

(iii) **Hydrogen**

$$T = 273 \text{ K}$$

$$P = 101325 \text{ Nm}^{-2}$$

$$\text{Molar mass} = M = 2 \text{ g mol}^{-1}$$

$$= \frac{2}{1000} \text{ kg mol}^{-1} = 0.002 \text{ kg mol}^{-1}$$

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

$$d = ?$$

The density of gas is given by

$$d = \frac{PM}{RT} = \frac{101325 \times 0.002}{8.314 \times 273}$$

$$= \boxed{0.0893 \text{ kg m}^{-3}}$$

(b) **Compare the values of densities in proportion to their molar masses.**

Density of a gas is directly proportional to its molar masses.

Since molar masses of O_2 , CH_4 and H_2 are in the order

$$\text{O}_2 > \text{CH}_4 > \text{H}_2$$

Hence, under given conditions, their densities are also in the same order.

(c) **How do you justify that increase of volume upto 100 dm^3 at 27°C of 2 moles of NH_3 will allow the gas behave ideally?**

By increasing the volume of NH_3 , the molecules of NH_3 become widely separated. As a result, forces of attraction between NH_3 molecules become less and hence it behaves ideally.

Q.18. A sample of Krypton with a volume of 6.25 dm^3 and a pressure of 765 torr and a temperature of 20°C is expanded to a volume of 9.55 dm^3 and a pressure of 375 torr. What will be its final temperature ($^\circ\text{C}$)

Solution:

$$V_1 = 6.25 \text{ dm}^3$$

$$V_2 = 9.55 \text{ dm}^3$$

$$P_1 = 765 \text{ torr}$$

$$P_2 = 375 \text{ torr}$$

$$T_1 = 20^\circ\text{C} + 273 = 293\text{K}$$

$$T_2 = ?$$

According to General Gas Equation

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

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

$$T_2 = \frac{293}{765 \times 6.25} \times 375 \times 9.55$$

$$T_2 = 219.46 \text{ K}$$

$$\text{Hence } T \text{ in } ^\circ\text{C} = \text{K} - 273 = 219.46 - 273.16 = -53.7^\circ\text{C}$$

Q.19. Working at vacuum line a chemist isolated a gas in weighing bulb with a volume of 255 cm^3 at a T of 25°C and under a pressure in the bulb of 10 torr . The gas weighed 12.1 mg . What is the molecular mass of this gas?

Solution:

$$V = 255 \text{ cm}^3 = \frac{255}{1000} \text{ dm}^3 = 0.255 \text{ dm}^3$$

$$T = 25^\circ\text{C} + 273 = 298\text{K}$$

$$P = 10 \text{ torr} = \frac{10}{760} \text{ atm} = 0.01316 \text{ atm}$$

$$\text{Mass of the gas} = m = 12.1 \text{ mg} = \frac{12.1}{1000} \text{ g} = 0.0121 \text{ g}$$

$$R = 0.0821 \text{ atm} \cdot \text{dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

$$\text{Mol. Mass of gas} = M = ?$$

According to general gas equation

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

$$M = \frac{0.0121 \times 0.0821 \times 298}{0.01316 \times 0.255}$$

$$M = 88.23 \text{ g/mol}$$

Q.20. What pressure is exerted by a mixture of 2 g of H_2 and 8 g of N_2 at 273 K in a 10 dm^3 vessel?

Solution:

$$\text{Mass of } \text{H}_2 = 2 \text{ g}$$

$$\text{Moles of } \text{H}_2 = n_{\text{H}_2} = \frac{2}{2} = 1 \text{ moles}$$

$$\text{Mass of } \text{N}_2 = 8 \text{ g}$$

$$\text{Moles of } \text{N}_2 = n_{\text{N}_2} = \frac{8}{28} = 0.286 \text{ moles}$$

$$\begin{aligned} \text{Total number of moles} &= n_t = n_{\text{H}_2} + n_{\text{N}_2} \\ &= 1 + 0.286 = 1.286 \text{ moles} \end{aligned}$$

$$\text{Volume of mixture} = V = 10 \text{ dm}^3$$

$$T = 273\text{K}$$

$$R = 0.0821 \text{ atm} \cdot \text{dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

$$\text{Total pressure} = P_t = ?$$

According to General Gas Equation

$$P_t V = n_t R T$$

$$P_t = \frac{n_t R T}{V}$$

$$\begin{aligned} P_t &= \frac{1.286 \times 0.0821 \times 273}{10} \\ &= 2.88 \text{ atm} \end{aligned}$$

Q.21. (a) The relative densities of two gases A and B are $1 : 1.5$. Find out the volume of B, which will diffuse in the same time in which 150 dm^3 of A will diffuse?

Solution:

Since relative densities of gases A and B are $1 : 1.5$

Hence relative density of gas A = $d_A = 1$

And relative density of gas B = $d_B = 1.5$

Volume of gas A diffused = 150 dm^3

Volume of gas B diffused = ?

According to Graham's Law of diffusion of gases

$$\frac{r_B}{r_A} = \sqrt{\frac{d_A}{d_B}}$$

Since rate of diffusions are directly proportional to the volume of gases diffused, hence

$$\frac{\text{volume of gas B diffused}}{\text{volume of gas A diffused}} = \sqrt{\frac{d_A}{d_B}}$$

$$\text{or } \frac{\text{volume of gas B diffused}}{150} = \sqrt{\frac{1}{1.5}}$$

$$\text{volume of gas B diffused} = \sqrt{\frac{1}{1.5}} \times 150 = \boxed{122.47 \text{ dm}^3}$$

(b) Hydrogen diffuses through a porous plate at a rate of 500 cm^3 per minute at 0°C . What is the rate of diffusion of oxygen through the same porous plate at 0°C ?

Solution:

Rate of diffusion of hydrogen = $r_{\text{H}_2} = 500 \text{ cm}^3/\text{min}$

Rate of diffusion of oxygen = $r_{\text{O}_2} = ?$

Molecular Mass of H_2 = $M_{\text{H}_2} = 2 \text{ g mol}^{-1}$

Molecular Mass of O_2 = $M_{\text{O}_2} = 32 \text{ g mol}^{-1}$

According to Graham's Law of diffusion of gases

$$\frac{r_{\text{O}_2}}{r_{\text{H}_2}} = \sqrt{\frac{M_{\text{H}_2}}{M_{\text{O}_2}}}$$

$$\text{or } \frac{r_{\text{O}_2}}{500} = \sqrt{\frac{2}{32}}$$

$$r_{\text{O}_2} = \sqrt{\frac{2}{32}} \times 500 = \boxed{125 \text{ cm}^3/\text{min}}$$

(c) The rate of effusion of an unknown gas A through a pinhole is found to be 0.279 times the rate of effusion of H_2 through the same pinhole. Calculate the molecular mass of the unknown gas at STP.

Solution:

Let Rate of effusion of unknown gas A = $r_A = 0.279$

Then Rate of effusion of unknown gas = $r_{\text{H}_2} = 1$

Molecular mass of hydrogen = $M_{\text{H}_2} = 2 \text{ g/mol}$

Molecular mass of unknown gas A = $M_A = ?$

According to Graham's law of effusion

$$\frac{r_{\text{H}_2}}{r_A} = \sqrt{\frac{M_A}{M_{\text{H}_2}}}$$

$$\text{or } \frac{M_A}{M_{\text{H}_2}} = \left(\frac{r_{\text{H}_2}}{r_A}\right)^2$$

$$\text{or } \frac{M_A}{M_{\text{H}_2}} = \left(\frac{r_{\text{H}_2}}{r_A}\right)^2$$

$$M_A = \left(\frac{1}{0.279}\right)^2 \times 2$$

$$M_A = 12.85 \times 2$$

$$= \boxed{25.7 \text{ g mol}^{-1}}$$

Q.22. Calculate the number of molecules and the number of atoms in given amounts of each gas

(a) 20 cm^3 of CH_4 at 0°C and a pressure of 700 mm of Hg

(Gujranwala Board, 2009)

Solution:

$$V = 20 \text{ cm}^3 = \frac{20}{1000} \text{ dm}^3 = 0.02 \text{ dm}^3$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

$$P = 700 \text{ mm of Hg} = \frac{700}{760} \text{ atm} = 0.921 \text{ atm}$$

$$R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

$$n = ?$$

According to general gas equation

$$PV = nRT$$

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

$$n = \frac{0.921 \times 0.02}{0.0821 \times 273} = 0.000822 \text{ mol}$$

$$\begin{aligned} 1 \text{ mole of CH}_4 \text{ contains molecules} &= 6.02 \times 10^{23} \\ 0.000822 \text{ moles of CH}_4 \text{ contains} &= 6.02 \times 10^{23} \times 0.000822 \\ &= 4.95 \times 10^{20} \text{ molecules} \end{aligned}$$

$$\begin{aligned} 1 \text{ molecule of CH}_4 \text{ contains atoms} &= 5 \\ 4.95 \times 10^{20} \text{ molecules contains} &= 5 \times 4.95 \times 10^{20} \\ &= \boxed{2.475 \times 10^{21} \text{ atoms}} \end{aligned}$$

(b) 1 ml of NH_3 at 100°C and a pressure of 1.5 atm

$$V = 1 \text{ cm}^3 = \frac{1}{1000} \text{ dm}^3 = 0.001 \text{ dm}^3$$

$$T = 100^\circ\text{C} + 273 = 373 \text{ K}$$

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

$$R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

$$n = ?$$

According to general gas equation

$$PV = nRT$$

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

$$n = \frac{1.5 \times 0.001}{0.0821 \times 373} = 4.9 \times 10^{-5} \text{ mol}$$

$$\begin{aligned} 1 \text{ mole of NH}_3 \text{ contains molecules} &= 6.02 \times 10^{23} \\ 0.000049 \text{ moles of NH}_3 \text{ contains} &= 6.02 \times 10^{23} \times 0.000049 \\ &= 2.95 \times 10^{19} \text{ molecules} \end{aligned}$$

$$\begin{aligned} 1 \text{ molecule of NH}_3 \text{ contains atoms} &= 4 \\ 2.95 \times 10^{19} \text{ molecules contains} &= 4 \times 2.95 \times 10^{19} \\ &= \boxed{1.18 \times 10^{20} \text{ atoms}} \end{aligned}$$

Q.23. Calculate the masses of 10^{20} molecules of each of H_2 , O_2 and CO_2 at STP. What will happen to the masses of these gases, when the T of these gases is increased by 100°C and the pressure is decreased by 100 mm of Hg?

Solution:

For H_2

$$\begin{aligned} \text{Number of molecules} &= 10^{20} \\ \text{Mass of Hydrogen} &= ? \end{aligned}$$

$$\text{Since } 6.02 \times 10^{23} \text{ molecules of H}_2 = 2 \text{ g}$$

$$\begin{aligned} \text{Therefore } 10^{20} \text{ molecules of H}_2 &= \frac{2}{6.02 \times 10^{23}} \times 10^{20} \\ &= \boxed{3.322 \times 10^{-4} \text{ g}} \end{aligned}$$

For O_2

$$\text{Number of molecules} = 10^{20}$$

$$\text{Mass of Oxygen} = ?$$

$$\text{Since } 6.02 \times 10^{23} \text{ molecules of O}_2 = 32 \text{ g}$$

$$\begin{aligned} \text{Therefore } 10^{20} \text{ molecules of O}_2 &= \frac{32}{6.02 \times 10^{23}} \times 10^{20} \\ &= \boxed{5.3152 \times 10^{-3} \text{ g}} \end{aligned}$$

For CO_2

$$\text{Number of molecules} = 10^{20}$$

$$\text{Mass of Carbon dioxide} = ?$$

$$\text{Since } 6.02 \times 10^{23} \text{ molecules of CO}_2 = 44 \text{ g}$$

$$\begin{aligned} \text{Therefore } 10^{20} \text{ molecules of CO}_2 &= \frac{44}{6.02 \times 10^{23}} \times 10^{20} \\ &= \boxed{7.3084 \times 10^{-3} \text{ g}} \end{aligned}$$

The change in T and P has no effect on masses of gas.

Q.24. Two moles of NH_3 are enclosed in a 5 dm^3 flask at 27°C .

(a) Calculate the pressure exerted by the gas assuming that

i. Gas behaves like an ideal gas

ii. Gas behaves like a real gas

$$a = 4.17 \text{ atm dm}^6 \text{ mol}^{-2} \quad b = 0.0371 \text{ dm}^3 \text{ mol}^{-1}$$

Solution:

(i) When the gas is ideal

$$V = 5 \text{ dm}^3$$

$$n = 2 \text{ mol}$$

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

$$R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

$$P = ?$$

According to general gas equation

$$PV = nRT$$

Or $P = \frac{nRT}{V}$
 $P = \frac{2 \times 0.0821 \times 300}{5}$
 $P = 9.852 \text{ atm}$

(ii) When the gas is real

$V = 5 \text{ dm}^3$
 $n = 2 \text{ mol}$
 $T = 27^\circ\text{C} + 273 = 300 \text{ K}$
 $R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$
 $a = 4.17 \text{ atm dm}^6 \text{ mol}^{-2}$ $b = 0.0371 \text{ dm}^3 \text{ mol}^{-1}$
 $P = ?$

According to Van der Waal's eq

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

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

or $P = \frac{nRT}{(V - nb)} - \frac{an^2}{V^2}$

Putting values

$$P = \frac{2 \times 0.0821 \times 300}{(5 - 2 \times 0.0371)} - \frac{4.17 \times 2^2}{5^2}$$

$$P = \frac{49.26}{4.926} - \frac{16.68}{25}$$

$$P = 10 - 0.67$$

$$P = 9.33 \text{ atm}$$

(b) Also calculate the amount of pressure lessened due to forces of attractions at these conditions of volume and temperature.

$$\text{Amount of pressure lessened} = 9.85 - 9.33 = 0.52 \text{ atm}$$

(c) Do you expect the same decreases of pressure when 2 moles of NH_3 is maintained in a volume of 40 dm^3 and T of 27°C

If the volume of the gas is increased to 40 dm^3 , then the pressure of gas will be less as compared to that calculated above in part (a).

It is because since volume is large, therefore molecules will be more widely spaced than before and they will have lesser forces of attractions. Hence gas will be more ideal and pressure difference from ideal behaviour will also be small.

HELLO! Mr. Question here!

Multiple Choice Questions from PAST PAPERS

- If "a" and "b" are zero for certain gas then gas is: (Lahore board, 2014)
 (a) Ideal (b) Real (c) Non-ideal (d) May be any diatomic gas
- Escape out of gas molecules one by one through tiny hole is: (Lahore board, 2014)
 (a) Diffusion (b) Effusion (c) Osmosis (d) All of these
- The spreading of fragrance or scent in air is due to: (Rawalpindi board, 2011)
 (a) Effusion (b) Diffusion (c) Osmosis (d) Density
- Partial pressure of O_2 in lungs (in torr) is: (Gujranwala board, 2008)
 (a) 150 (b) 760 (c) 116 (d) 159
- The constant factor in Charles's law is: (Gujranwala board, 2008)
 (a) volume (b) temperature (c) pressure (d) all of these
- Partial pressure of oxygen in human lungs in torr is: (Gujranwala board, 2011)
 (a) 161 (b) 116 (c) 159 (d) 760
- The S.I. unit of pressure is expressed in (Rawalpindi Board, 2013) (Multan Board, 2013)
 (a) Nm^{-1} (b) Nm^{-2} (c) Nm^{-3} (d) mm Hg
- Which gas will diffuse more rapidly: (Lahore Board, 2009)
 (a) CO_2 (b) NH_3 (c) HCl (d) SO_2
- Plasma is conductor of electricity: (Faisalabad Board, 2009)
 (a) Bad (b) Poor (c) Good (d) None
- SI unit of pressure is: (Rawalpindi Board, 2009)
 (a) torr (b) mmHg (c) Nm^{-2} (d) pound inch⁻²
- Density of an ideal gas can be calculated by the formula. (Faisalabad Board, 2011)
 (a) $d = nRT$ (b) $d = \frac{PM}{RT}$ (c) $d = \frac{m}{M}RT$ (d) $d = \frac{PV}{M}$
- Absolute zero is equal to (Lahore Board, 2007)
 (a) 273°C (b) -273°C (c) 0°C (d) 273K
- Which gas will diffuse more rapidly among the following? (D.G. Khan Board, 2010)
 (a) N_2 (b) H_2 (c) CO (d) NH_3
- To calculate the pressure of volume of a real gas under the non-ideal conditions, alternative kinetic equation has been developed. This is known as: (Faisalabad Board, 2010)
 (a) General gas equation (b) Arrhenius equation
 (c) Clausius Clapeyron equation (d) Vander Wall's equation

15. An ideal gas has volume 1 dm^3 at 303 K . Keeping pressure constant, at which Kelvin temperature its volume becomes 2 dm^3 (Sargodha Board, 2011)
 (a) 240 (b) 303 (c) 330 (d) 606
16. Which gas will diffuse more rapidly (Sargodha Board, 2014)
 (a) CO_2 (b) NH_3 (c) HCl (d) SO_2

Answers to Multiple Choice Questions from Past Papers.

Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans
1	(a)	2	(b)	3	(b)	4	(c)	5	(c)
6	(b)	7	(b)	8	(b)	9	(c)	10	(c)
11	(b)	12	(b)	13	(b)	14	(d)	15	(d)
16	(b)								

Detailed Explanation of Past Papers MCQs & answers to all Past Papers SHORT QUESTIONS in COLLEGE CHEMISTRY OBJECTIVE BOOK-I

SHORT & LONG QUESTIONS FROM PAST PAPERS

PRESSURE, BOYLE'S LAW, CHARLES'S LAW

Short Questions

- Define pressure. Give its units. (Lahore Board, 2009; Multan Board, 2009; D.G. Khan Board, 2011; Sargodha Board, 2012) OR Compare the different units of pressure. (Sargodha Board, 2007)
- Define Boyle's law. Give its expression. (Faisalabad Board, 2009, 2012; Gujranwala Board, 2013)
- The product of pressure and volume at constant temperature and number of moles is constant. Why? (Sargodha Board, 2009)
- Why the graph plotted between pressure and volume moves away from pressure axis at higher temperature. (Sargodha Board, 2013)
- Why do we get a straight line when pressure are plotted against inverse volume of a gas. (Bahawalpur Board, 2009; Gujranwala Board, 2011; Rawalpindi Board, 2013)
- Greater the temperature of the gas, closer the straight line of P versus $1/V$ to the pressure axis. Justify it. (Lahore Board, 2007)
- State Charles's Law and write its mathematical form (Multan Board, 2012)
- Volume of a gas is doubled when temperature is raised from 0°C to 273°C . Why? (Bahawalpur Board, 2008)

Long Questions

Define Boyle's law. Give its experimental verification. (Bahawalpur Board, 2011)

ABSOLUTE ZERO

Short Questions

- Give the quantitative definition of Charles's law? (Faisalabad Board, 2008; Bahawalpur Board, 2012)
- What do you mean by absolute zero? (Faisalabad Board, 2008; Bahawalpur Board, 2008; Sargodha Board, 2009; D.G. Khan Board, 2012) OR What is absolute zero? What happens to real gases while approaching it? (Sargodha Board, 2012; Faisalabad Board, 2013)
- The volume of a given mass of a gas becomes theoretically zero at -273°C . Justify it. (Rawalpindi Board, 2007, 2009; Lahore Board, 2014)

SCALES OF THERMOMETRY

Short Questions

- What are different scales of thermometry? OR What is thermometry? Name the scales and device used for thermometry. (Multan Board, 2011) OR How the various scales of thermometry can be interconverted? (Lahore Board, 2007)
- Give two important scales of thermometry. How are these related? (Lahore Board, 2008; Faisalabad Board, 2009)
- -273°C is regarded as the lowest possible temperature. Justify it. (Gujranwala Board, 2010)
- Volume of real gas cannot be zero at any temperature. Why? (Sargodha Board, 2007)
- Convert 80°C to Fahrenheit scale. (Faisalabad Board, 2010)
- Convert 40°F to Kelvin temperature (Bahawalpur Board, 2010)
- Convert -40°F to (a) Centigrade scale (b) Kelvin scale (D.G. Khan Board, 2011)
- Convert -40°C to $^\circ\text{F}$ (Faisalabad Board, 2012)
- Convert -40°F temperature to kelvin temperature. (Rawalpindi Board, 2012)

GENERAL GAS EQUATION, IDEAL GAS CONSTANT

Short Questions

- What is general gas equation? Derive it in various forms. (Multan Board, 2010)
- Derive the unit of ' R ' in general gas equation when the pressure is in atmosphere and volume in dm^3 (Multan Board, 2012)
- Calculate the value of R in SI units OR Derive the units for gas constant R in general gas equation when pressure is in Nm^{-2} and volume in m^3 (Lahore Board, 2008, 2013; Gujranwala Board, 2011, 2013, 2014; Multan Board, 2008, 2012; D.G. Khan Board, 2009; Rawalpindi Board, 2009, 2010; Sargodha Board, 2012, 2013)
- Derive general gas equation for one mole of a gas? (Gujranwala Board, 2011)

AVOGADRO'S LAW, DALTON'S LAW OF PARTIAL PRESSURE AND ITS APPLICATIONS.

Short Questions

- What is Avogadro's law of gases? (Rawalpindi Board, 2010; Sargodha Board, 2014; Lahore Board, 2014) OR Explain Avogadro's law with example (D.G. Khan Board, 2012)
- Define Dalton's law of partial pressure. Give an example. (Lahore Board, 2009; Gujranwala Board, 2011; Bahawalpur Board, 2011) OR Define Dalton's law of partial pressure. Give mathematical expression (Faisalabad Board, 2011; Azad Kashmir Board, 2012)
- Prove that $P_i = P \cdot X_i$ (Lahore Board, 2014)

- (4) The process of respiration obeys the Dalton's law of partial pressure. Justify. (D.G. Khan Board, 2012; Sargodha Board, 2013)
- (5) Why pilots (or we) feel uncomfortable breathing at high altitude? (Faisalabad Board, 2007; Sargodha Board, 2010, 2011)
- (6) Why regular air cannot be used in divers tank? (Lahore Board, 2010) **OR** Explain procedure of sea divers breath. (Rawalpindi Board, 2011)
- (7) How Dalton's law of partial pressure is useful in determining pressure of a gas collected at water? (D.G. Khan Board, 2010; Multan Board, 2011)
- (8) Explain two applications of Dalton's law of partial pressures. (D.G. Khan Board, 2007; Gujranwala Board, 2012) **Ans:** Combine S1 and S4
- (9) Calculate fraction of total pressure exerted by oxygen when equal masses of CH_4 and O_2 are mixed in an empty container at 25°C . (Lahore Board, 2014)

Long Questions

- (1) Describe Dalton's law of partial pressure. Write its two/three/four applications. (Lahore Board, 2010; Faisalabad Board, 2011; Lahore Board, 2013; Gujranwala Board, 2013)
- (2) What are the applications of Dalton's law of partial pressure? (Sargodha Board, 2014) **OR** Write down the applications of Dalton's Law of Partial Pressure. (Rawalpindi Board, 2010)

DIFFUSION, EFFUSION, GRAHAM'S LAW OF DIFFUSION**Short Questions**

- (1) Define diffusion and effusion. (Faisalabad Board, 2013; Sargodha Board, 2013)
- (2) Differentiate between diffusion and effusion of gases. (Faisalabad Board, 2007; Bahawalpur Board, 2012)
- (3) State Graham's law of diffusion and write its mathematical form. (D.G. Khan Board, 2012; Rawalpindi Board, 2012)
- (4) Rate of diffusion of NH_3 gas is more than HCl gas. Why? (Bahawalpur Board, 2011) **OR** Lighter gases can diffuse more rapidly than heavier gases. Why? (Gujranwala Board, 2008, 2009; D.G. Khan Board, 2009)

Long Questions

- (1) Define and verify Graham's law of diffusion of gases. (Faisalabad Board, 2010; D.G. Khan Board, 2011)
- (2) State and explain Graham's Law of diffusion of gases. (Lahore Board, 2011)

KINETIC MOLECULAR THEORY OF GASES**Short Questions**

- (1) List four postulates of kinetic molecular theory of gases. (Faisalabad Board, 2011; Lahore Board, 2013)
- (2) Give mathematic expression of Kinetic equation and root mean square velocity. (Multan Board, 2012)
- (3) Give mathematical expression for mean square velocity and root mean square velocity. (Rawalpindi Board, 2009)
- (4) Why gases do not settle down in a vessel? (Gujranwala Board, 2008)
- (5) What are elastic collisions? Give an example. (Multan Board, 2007)

Long Questions

- (1) What is Kinetic molecular theory of gases. Give its Postulates. (Sargodha Board, 2013)

EXPLANATION OF GAS LAWS ON THE BASIS OF KINETIC MOLECULAR THEORY OF GASES**Short Questions**

- (1) Explain Boyle's law according to kinetic molecular theory of gases. (Lahore Board, 2013, 2014)
- (2) Apply kinetic molecular theory to explain the Avogadro's Law. (Sargodha Board, 2011)
- (3) Derive Graham's law of diffusion from kinetic molecular theory of gases. (Gujranwala Board, 2008)

Long Questions

- (1) Define Graham's law. Explain Graham's law of diffusion according to kinetic equation. (D.G. Khan Board, 2010)
- (2) Explain Avogadro's law and Graham's law of diffusion on the basis of Kinetic Molecular Theory of gases. (Multan Board, 2011)
- (3) What is kinetic molecular theory of gases? Derive Boyle's law from kinetic equation. (Faisalabad Board, 2013)
- (4) How are Boyle's law and Charle's law derived from kinetic molecular theory of gases? (Multan Board, 2013)

CRITICAL TEMPERATURE, PRESSURE, LIQUEFICATION OF GASES, LINDE'S METHOD**Short Questions**

- (1) Define critical temperature and pressure. (Sargodha Board, 2010) **OR** Differentiate between critical temperature and critical pressure. (Gujranwala Board, 2012)
- (2) What is the critical temperature of a gas? Give its importance for liquefaction of gases (D.G. Khan Board, 2008; Faisalabad Board, 2009; Bahawalpur Board, 2012)
- (3) What is critical temperature? It depends upon what factors? (D.G. Khan Board, 2010; Multan Board, 2013)
- (4) State Joule-Thomson effect. What is its applications? (Lahore Board, 2010; Faisalabad Board, 2010; Bahawalpur Board, 2010; Gujranwala Board, 2012)
- (5) Explain Linde's method of liquefaction of gases. (Multan Board, 2010)
- (6) Joule-Thomson effect is operative in the Linde's Method of Liquefaction of air. How? (Lahore Board, 2007)
- (7) State Joule-Thomson effect. Why it is not applicable to H_2 gas? (Lahore Board, 2011) **OR** Ammonia gas can be liquefied quite easily than H_2 gas. Justify it. (D.G. Khan Board, 2008)

Long Questions

- (1) Describe Linde's method of liquefaction of gases. (Lahore Board, 2013)

IDEAL AND NON-IDEAL GASES, CAUSES OF DEVIATION**Short Questions**

- (1) What is compressibility factor? What is its value for an ideal gas? (Multan Board, 2007)
- (2) What are faulty points in kinetic molecular theory of gases? (Faisalabad Board, 2008; Lahore Board, 2009; Bahawalpur Board, 2011; Lahore Board, 2012) **OR** Describe two causes of deviation from ideality. (Gujranwala Board, 2009; Rawalpindi Board, 2011, 2013) **OR** Why real gases deviate from ideal behaviour? (Bahawalpur Board, 2010; Faisalabad Board, 2010; D.G. Khan Board, 2012; Sargodha Board, 2011, 2012, 2014)
- (3) Why high pressure and low temperature make a gas non-ideal? **OR** Gases deviate significantly at high pressure and low temperature. (Sargodha Board, 2007, 2010; Gujranwala Board, 2010)
- (4) 2010: Faisalabad Board, 2011, 2012; Multan Board, 2012; Multan Board, 2013)

- (5) Gases deviate from ideal behaviour more significantly at high pressure. Why? (Multan Board, 2007; Bahawalpur Board, 2009)

Long Questions

- (1) Why real gases deviate from the ideal behaviour? Discuss its causes. (Sargodha Board, 2011)

VAN DER WAAL'S EQUATIONShort Questions

- (1) Why the pressure correction is done by van der Waals? (Rawalpindi Board, 2007; Multan Board, 2009; D.G. Khan Board, 2010)
 (2) Give (S.I.) units of 'a' and 'b' in van der Waals equation. (D.G. Khan Board, 2009, 2011; Rawalpindi Board, 2012; Multan Board, 2012; Faisalabad Board, 2013; Rawalpindi Board, 2013)
 (3) Gases deviate more from ideal behaviour at 0°C than at 100°C. Why? (Lahore Board, 2008)

Long Questions

- (1) Derive Vander Wall's equation for real gases. (Multan Board, 2010; Bahawalpur Board, 2010)
 (2) Derive Vander Wall's equation for real gases and give the physical significance of van der Waal's constants 'a' and 'b'. (Sargodha Board, 2013)

PLASMA STATEShort Questions

- (1) What is plasma state (or fourth state of matter)? How it is formed? (D.G. Khan Board, 2008; Bahawalpur Board, 2009; Multan Board, 2011)
 (2) Where is the plasma formed (or found)? (Bahawalpur Board, 2008; Multan Board, 2009; Gujranwala Board, 2009)
 (3) What are characteristics of plasma? (Multan Board, 2008; Azad Kashmir Board, 2012)
 (4) What is plasma? Write four uses (or applications) of plasma. (D.G. Khan Board, 2007; Multan Board, 2010, 2012; Bahawalpur Board, 2012; D.G. Khan Board, 2012; Rawalpindi Board, 2012; Gujranwala Board, 2012, 2014; Lahore Board, 2011, 2013; Sargodha Board, 2013)
 (5) State what is natural and artificial plasma (Faisalabad Board, 2012)

Long Questions

- (1) What is plasma? Give three applications of plasma. (Sargodha Board, 2010)

TEST YOUR SKILLS

Marks: 85

OBJECTIVE

Time: 20 Minutes

Marks: 17

Note: Over writing, cutting, erasing, using lead pencil will result in loss of marks.

Q1. Each question has four possible answers. Choose the correct answer and encircle it.

- (i) If absolute temperature of a gas is doubled and the pressure is also doubled, the volume of the gas will
 (a) remain unchanged (b) increase four times (c) reduce to $\frac{1}{4}$ (d) be doubled
- (ii) Charles's law is applicable to
 (a) ideal gases (b) real gases (c) all gases (d) none
- (iii) Partial pressure of O_2 in lungs (in torr) is?
 (a) 150 (b) 760 (c) 116 (d) 159
- (iv) If 2 moles of an ideal gas at 546 K occupy a volume of 44.8 liters, the pressure must be
 (a) 2 atm (b) 3 atm (c) 4 atm (d) 1 atm
- (v) The rate of diffusion of H_2 , compared with He is
 (a) $\frac{1}{2}$ times (b) 1.4 times (c) 2 times (d) 4 times
- (vi) The constant factor in Charles's law is:
 (a) volume (b) temperature (c) pressure (d) all of these
- (vii) The molar volume of helium is 44.8 dm³ at
 (a) 100°C and 1 atm (b) 25°C and 0.25 atm (c) 0°C and 0.5 atm (d) 40°C and 0.5 atm
- (viii) The total pressure exerted by a number of non-reacting gases is equal to the sum of the partial pressures of the gases under the same condition is known as
 (a) Boyle's law (b) Avogadro's law (c) Charles's law (d) Dalton's law
- (ix) An ideal gas has volume 1 dm³ at 303K keeping pressure constant at which Kelvin temperature its volume will become 2 dm³
 (a) 240 (b) 303 (c) 330 (d) 606
- (x) Which of the following gases will have the highest rate of diffusion?
 (a) O_2 (b) CO_2 (c) NH_3 (d) N_2
- (xi) The van der Waal's equation explains the behaviour of
 (a) ideal gas (b) real gases (c) liquids (d) non-real gases
- (xii) Absolute zero is equal to
 (a) -273.15 K (b) -273.15 °F (c) -273.15 °C (d) 0°C
- (xiii) The value of R in SI units is
 (a) 0.0821 dm³ . atm . J⁻¹ mole⁻¹ (b) 62.4 dm³ . torr K⁻¹ mole⁻¹
 (c) 8.31 dm³ . atm . K⁻¹ mole⁻¹ (d) 8.31 JK⁻¹ mole⁻¹
- (xiv) One dm³ of O_2 at STP has mass
 (a) 32g (b) 16g (c) 4.438g (d) 1.428g
- (xv) The volume of 1 mole of N_2 is maximum at
 (a) 0°C and 2 atm (b) STP (c) 100°C and 1 atm (d) 150°C and atm
- (xvi) The critical temperature for NH_3 is
 (a) less than Argon (b) equal to Argon (c) greater than Argon (d) Not known
- (xvii) The highest temperature above which a gas cannot be liquefied is called as
 (a) Critical temperature (b) Absolute (c) Liquefaction (d) Boiling point

SUBJECTIVE

Marks: 68

Time: 2-10 Hours

Note: Out of Questions 2,3 and 4, Write any TWENTY TWO(22) short answers. While writing answers write question numbers carefully. (22x 2)=44

Section - I

Q2. Answer any Eight parts from the followings.

- (i) What is Joule-Thomson effect?
 (ii) What is absolute zero?
 (iii) What is the physical significance of van der Waals constants 'a' and 'b'?
 (iv) What is plasma? Where it is found?

- (v) Gases behave ideally at high temperature and Low pressure. Why?
 (vi) How does heat transfer from a hot body to cold body?
 (vii) Prove Graham's Law of diffusion of gases.
 (viii) Differentiate between an ideal & a non ideal gas.
 (ix) Dalton's law of partial pressure is not obeyed by polar gases. Why?
 (x) 1 dm^3 of H_2 at STP weighs 0.0899 g & 1 dm^3 of O_2 at STP weight 1.438 g , but they have same volume why?
 (xi) Calculate the value of R in SI units.
 (xii) Why gases diffuse into each other?

Q3. Answer any Six parts from the followings.

- (i) Why the graph between $1/V$ on x-axis and P on Y-axis is a straight line?
 (ii) What is the effect of increasing temperature on isotherm of a given mass of a gas?
 (iii) The graph PV on y-axis and P on x-axis is a straight line parallel to x-axis. Why?
 (iv) What is the general equation?
 (v) Why pilots feel uncomfortable breathing at high altitude?
 (vi) How can Charles's law be obtained from the general gas equation $PV = nRT$?
 (vii) What is the effect of temperature and pressure on the density of a gas?
 (viii) Dalton's law of partial pressure is not obeyed by a mixture of NH_3 and SO_2 gases. Why?
 (ix) Why gases collected over water are impure gases?
 (x) How is artificial plasma produced?
 (xi) Why fluorescent tube produce light?
 (xii) Lighter gases diffuse more rapidly than heavier gases explain?

Q4. Give the short answers (only 06 questions).

- (i) Why real gases deviate from ideal behaviour?
 (ii) Why did van der Waals correct the pressure in the general gas equation?
 (iii) The product of pressure and volume at constant temperature and number of moles is constant why?
 (iv) Why spiral pipe is used in Lind's apparatus for liquefaction of gases?
 (v) Why high pressure and low temperature make gas non ideal?
 (vi) In solids, liquids and gases, temperature is the measure of which kind of energies?
 (vii) Define Critical temperature and critical pressure?
 (viii) The volume of the given mass of a gas becomes theoretically zero at 273°C . justify it?
 (ix) Polar gases have high critical temperature. Why?

Section - II (Attempt any three questions) (8x 3)=24

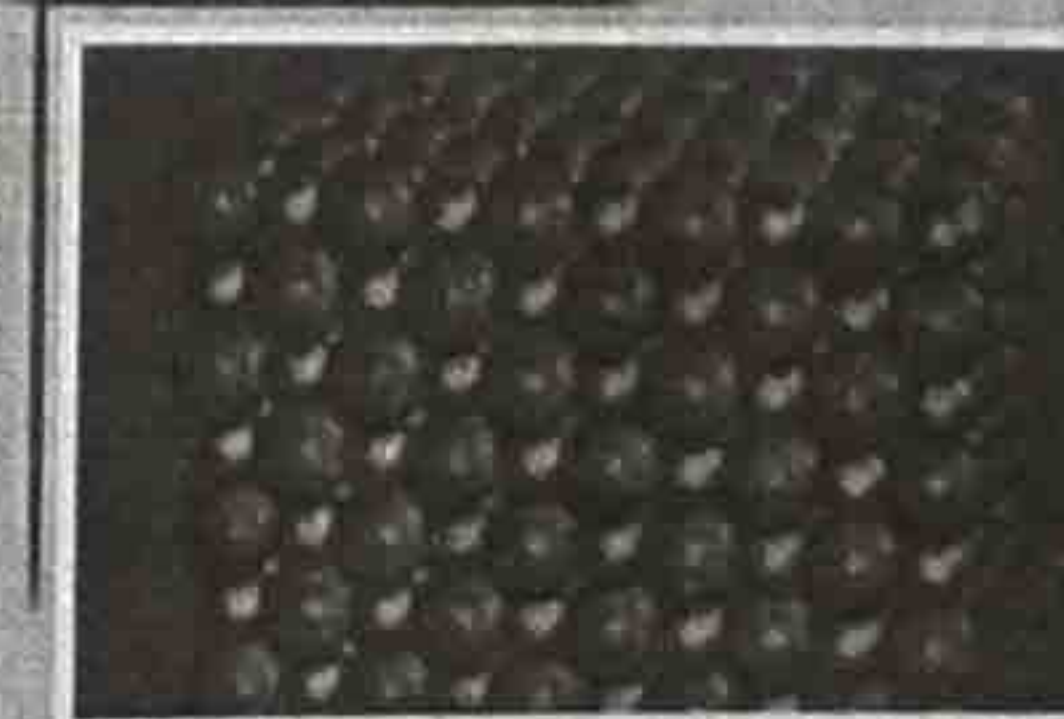
- Q5.** (a) How general gas equation can be used to determine the molecular mass of a gas. (02)
 (b) Calculate the mass of 11.2 dm^3 of H_2 at 0°C and 2 atm pressure considering hydrogen to be ideal under these condition (03)
 (c) Prove Avogadro's law using kinetic equation of kinetic molecular theory. (03)
Q6. (a) Derive van der Waals equation for real gases (04)
 (b) Three moles of SO_2 are enclosed in a 5 dm^3 flask at 27°C .
 (i) Calculate the pressure exerted by the gas assuming that
 (a) Gas behaves like an ideal gas
 (b) Gas behaves like a real gas For SO_2 $a = 6.170 \text{ atm dm}^6 \text{ mol}^{-2}$ $b = 0.0564 \text{ dm}^3 \text{ mol}^{-1}$
 (ii) Also calculate the amount of pressure lessened due to forces of attractions at these conditions of volume and temperature. (03)
Q7. (a) What are ideal and non-ideal gases? Why do real gases deviate from ideal behaviour? Explain with graph. (02)
 (b) What pressure is exerted by a mixture of 2 g of He , 16 g of O_2 and 10 g of CO_2 at 10°C in a 5 dm^3 vessel? (03)
 (c) Write a short note on Absolute zero. (04)
Q8. (a) What is the role of critical temperature in the liquefaction of gases? Describe Lind's method for the liquefaction of gases. (04)
 (b) Write down the postulates of kinetic molecular theory of gases. (02)
Q9. (a) Give experimental verification of Graham's law. (03)
 (b) Derive the expression between kinetic energy and temperature for a gas using kinetic equation for gases.
 (c) The rate of effusion of an unknown gas A through a porous plate is found to be $1/4$ times the rate of effusion of H_2 through the same porous plate. Calculate the molecular mass of the unknown gas at STP. (03)

Chapter 4

LIQUIDS AND SOLIDS



van der Waals



NaCl Crystal

