CHAPTER

4 STATES OF MATTER-I [GASES]

Q1: Define gaseous state. What are the characteristic properties of gases?

Ans. **Definition:** "The highest energy state of matter having no definite shape and volume is called a gaseous state."

Characteristics Properties of Gases:

1. Most of the gases are colourless, odourless and tasteless.
2. Gases expand by heating or decreasing external pressure.
3. All gases exert pressure.
4. Gases have low densities as compared to solids and liquids.
5. Non-reacting gases diffuse into each others.
6. All gases can effuse through a hole in the wall of the container.
7. Gases have weak intermolecular attractive forces.
8. Gases can be liquefied by increasing external pressure and decreasing temperature.
9. Gases can be compressed by increasing external pressure.

Q2: State basic postulates of kinetic molecular theory of gases. (ETEA: 2002, BISE D.I Khan 2017)

Ans. The kinetic molecular theory of gases was first postulated by Daniel Bernoulli [Swedish physicist 8/2/1700 – 17/3/1782], which was further developed by James Maxwell and Boltzmann to explain the various properties of gases [Bernoulli presented KMT in 1738].

Main Postulates:

1. All gases are composed of small invisible particles called molecules.
2. Gas molecules are far away (بعید) from each others.
3. Gas molecules have higher K.E than those of liquids and solids.
4. Gas molecules are widely separated from each others. One molecule is at a distance of 300 times of its own diameter from the other molecule. Due to long distance among molecules, volume of

the gas consists of mostly empty (خالی) spaces.

5. The volume of a single gas molecule is very very small as compared to the total volume of the gas. OR volume of one molecule is negligible as compared to total volume of the gas.
6. The attractive or repulsive forces among the gas molecules are negligible. Therefore, every gas molecule behaves independently.
7. Gas molecules are in constant random motion. They collide with each others and also with the walls of the container. The pressure of gas molecules is due to the collision of molecules.
8. All the collisions of a gas molecules are elastic i.e. no gain or loss of energy takes place.
9. The average kinetic energy of gas molecules is directly proportional to absolute temperature.
10. The molecules of different gases have the same kinetic energy at the same temperature and pressure.
11. Force of gravity has almost no effect on the gas molecules.

Q3: Define pressure. What are its units? Explain the informations used for the measurement of pressure. (BISE Malakand 2019)

Ans. Pressure:

Force per unit area is called pressure.

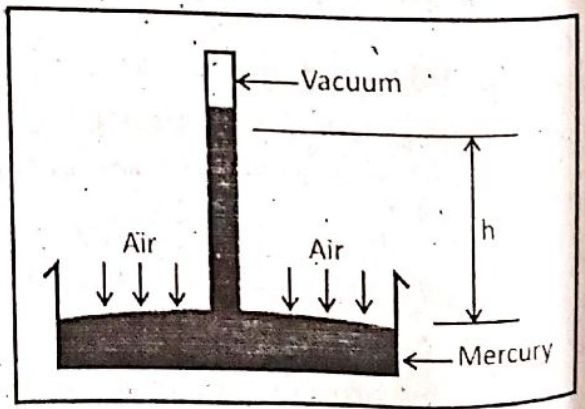
$$\text{Pressure } (P) = \frac{\text{Force } (F)}{\text{Area } (A)}$$

$$P = \frac{F}{A}$$

The apparatus used for the measurements gas pressure are:

1. Manometer
2. Barometer

A common type of barometer is the Torricellian barometer [Italian scientist Evangelista Torricelli who invented the barometer]. Torricellian barometer is shown below:



"Take 100cm [1m] long glass tube having cross sectional area of $1cm^2$. Fill this tube with mercury [hydragyrum - Hg] and place it invertedly in

a dish of mercury. The mercury in the column 1st come downward due to gravitational force and then will stay at a height of 76cm. The atmosphere exert pressure on the surface of the liquid mercury in the tub, forcing it to keeps its level in the tube at 76cm.

This experiment is carried out at sea level and at 0°C. The space above the column of mercury in the tube is vacuum, so that the only downward pressure in the tube is that exerted by the weight of the liquid column of mercury. Thus the atmospheric pressure may be directly measured by the height of the mercury column.

1. Atmospheric Pressure:

Definition: "The pressure exerted by 76cm long mercury column upon 1cm^2 area at sea level and at 0°C [273K] is called 1-atmospheric pressure."

Explanation: Atmospheric pressure is the pressure exerted by earth's atmosphere. The atmosphere is the thick blanket of air [gases] surrounding the earth, which exert pressure on the earth surface. Gravitational force attracts air molecules. The force experienced by any area exposed to earth's atmosphere is equal to the weight of the column of air above it.

Atmospheric pressure at sea level is about equal to the weight of 1.03 kg mass per square centimeter of surface OR 101.325KN/m^2. The pressure of the atmosphere is the weight of the gases that compose the atmosphere. Atmospheric pressure is the sum of the individual pressures of the various gases, [N = 78%, O = 21% and other gases = 1%] in atmosphere. The actual value of atmospheric pressure depends on location, temperature and weather conditions.

Units of Pressure: (BISE Bannu 2017)

The SI unit of force is the Newton and the unit of area is m^2 so

$$P = \frac{F}{A} \rightarrow \frac{N}{m^2} \Rightarrow (N \cdot m^{-2})$$

Unit	Symbol	
Pascal (Newton per square meter	Pa	SI pressure unit, $1Pa = \frac{1N}{m^2}$
Millimeter of mercury)	mmHg	Pressure that support a 1mm Hg column in a barometer.
Torr	torr	1 torr = 1mmHg
Atmosphere	atm	Average atmospheric pressure at sea level at 273k. 1 atm = 760mmHg 1 atm = 760 torrs 1 atm = 101325 Pascals 1 atm = 1.01325×10^5 Pa 1 atm = 101.325 K·Pa
Pound per square inch	psi	1 atm = 14.716/in2 1 atm = 143.7 psi
Bar	Bar	1 atm = 1.01325 bar
1 atm = 76cmHg = 760mmhg = 760 torrs = 101325Pa = 14.7 psi		

EXAMPLE 4.1:

The pressure of a gas is 49 torr. Convert this pressure into both atmospheres and pascal.

Solution:

(i) $\frac{1 - atm\ pressure = 760\ torrs}{Pressure\ in\ torr = 49\ torrs}$

As 760 torrs = 1 atm.

$$1\ torr = \frac{1\ atm}{760\ torrs}$$

$$49\ \cancel{torrs} = \frac{1\ atm}{760\ \cancel{torrs}} \times 49\ torrs = (0.06447\ atm)$$

$$= (6.447 \times 10^{-2}\ atm)$$

(ii) Pressure in atm = 6.447×10^{-2} atm

We know that 1 atm = 101325 Pascals

$$6.447 \times 10^{-2}\ atm = 101325 \times 6.447 \times 10^{-2}\ Pascal$$

$$= [6.5327 \times 10^3\ Pascals]$$

PRACTICE PROBLEM 4.1:

Reading on barometer is 2.21 atm. Calculate the corresponding pressure in (a) torr (b) Pascal

Solution:

Given pressure = 2.21atm

(a) We know that,

1 atm = 760 torrs

2.21 atm = 2.21 x 760 = (1679.6 torrs)

(b) We know that,

1 atm = 101325 Pascals

2.21 atm = 2.21 x 101325 Pascals

(223928 Pascals)

Q4: State and explain Boyle's law.

Ans: Boyle's Law:

In 1662, Robert Boyle gave the relationship between the pressure and volume of a fixed mass of a gas, at constant temperature.

Statement:

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

Mathematical Form:

$$V \propto \frac{1}{P} \quad [T \text{ and mass} = \text{constant}]$$

$$V = K \cdot \frac{1}{P} \quad [K = \text{proportionality constant}]$$

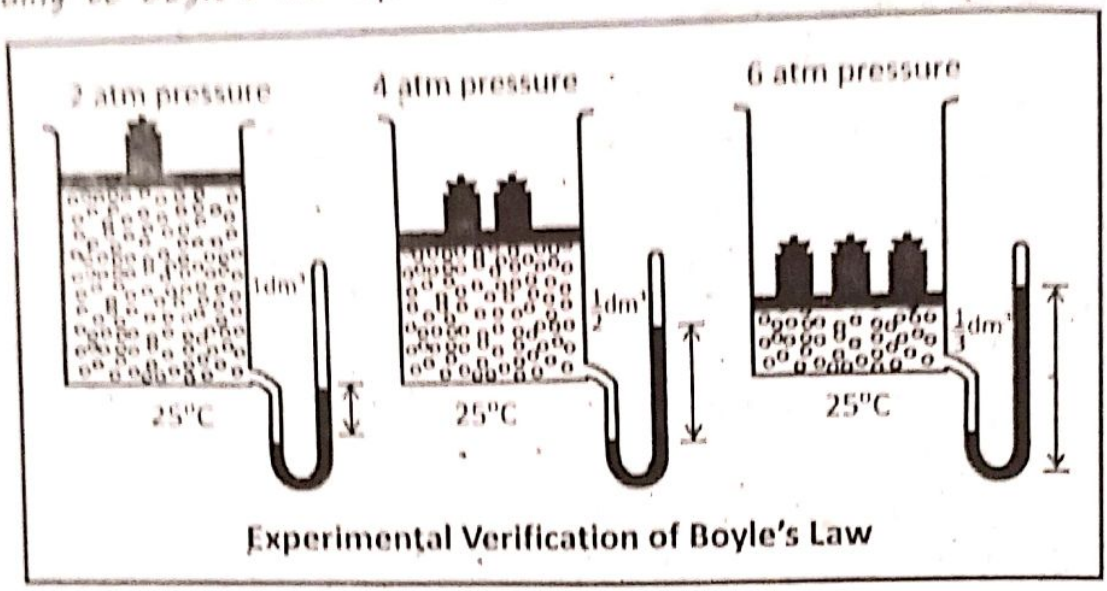
$$PV = K$$

It means that the product of pressure and volume is a constant quantity.

Experimental Verification:

Take a gas in a cylinder having moveable piston. The cylinder is attached with a manometer to read the pressure of the gas directly. A weight equal to two atmospheres is placed on the piston. Let its initial volume (V_1) is $2dm^3$ and initial pressure (P_1) is 2atm. When pressure (P_2) on the piston becomes double by placing two weights on the piston, volume (V_2) of the gas is reduced to $1dm^3$ and so on.

According to Boyle's law equation,



Experimental Verification of Boyle's Law

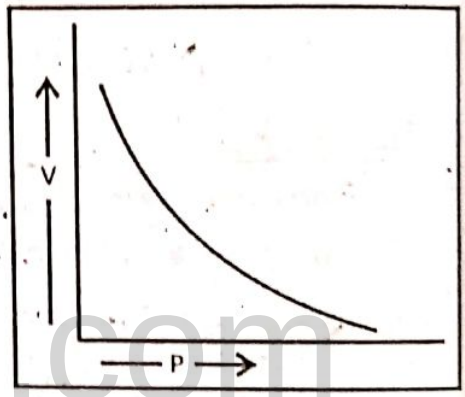
$P_1V_1 = 2atm \times 2dm^3 = 4atm \cdot dm^3$

$P_2V_2 = 2atm \times 1dm^3 = 4atm \cdot dm^3$

It means that $P_1V_1 = P_2V_2 \ldots\ldots\ldots P_nV_n$.

Graphical Representation:

When pressure of the given mass of a gas is plotted against its volume, a curve called hyperbola is obtained. As the temperature remains constant, so this graph is also called isothermic graph. This graph indicates that "V" is inversely proportional to the applied pressure.



EXAMPLE 4.2.

What pressure is required to compress $5dm^3$ of a gas at 1atm pressure to $1dm^3$ at constant temperature?

Solution:

Initial volume $(V_1) = 5dm^3$

Initial pressure $(P_1) = 1atm$

Final volume $= (V_2) = 1dm^3$

Final pressure $(P_2) = ?$

Using Boyle's equation

$P_1V_1 = P_2V_2$

$$P_2 = \frac{P_1V_1}{V_2} = \frac{1atm \times 5\,dm^3}{1\,dm^3} = (5atm)$$

Q5: State and explain Charles law.

Ans: Charles Law:

In 1787 a French physicist Jacques Charles and Gay-Lussac independently observed relationship between volume and temperature at constant external pressure.

Statement:

"The volume of the fixed mass of a gas is directly proportional to the absolute temperature at constant pressure."

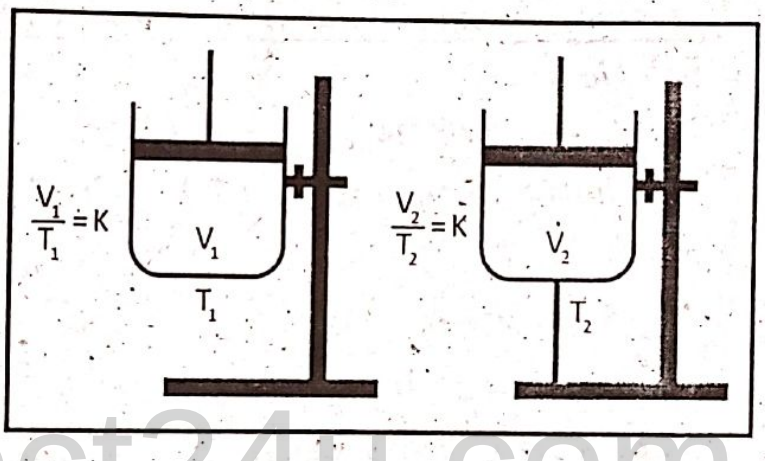
Mathematical Form:

Mathematically it can be written as;

$V \propto T$ [Pressure & mass = constant]

$V = KT$ [K = proportionality constant]

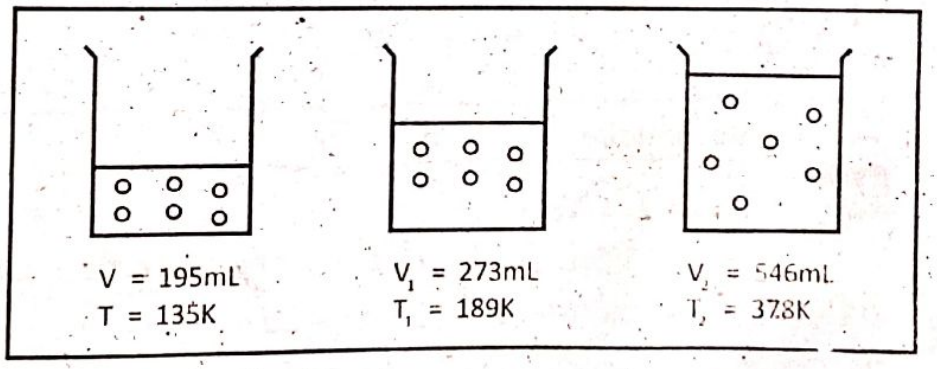
$$\boxed{\frac{V}{T} = K}$$



It means that the ratio between volume to temperature remains constant.

Experimental Verification:

Consider a gas is enclosed in a cylinder, fitted with a moveable and frictionless piston, and a constant pressure is applied. When the gas is heated, K.E of its molecules increases, as a result the piston moves upward and thus volume of the gas increases. But the ratio of volume to temperature (V/T) remains constant.



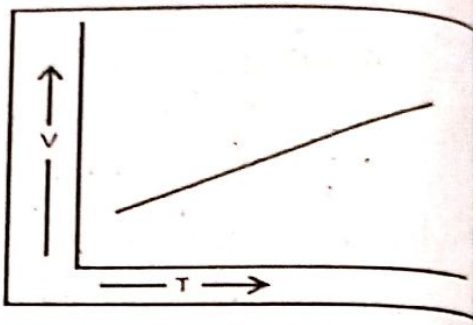
$$\left[\frac{V}{T} = \frac{195mL}{135K} = 1.44mL.K^{-1}\right]\left[\frac{V_1}{T_1} = \frac{273mL}{189K} = 1.44mL.K^{-1}\right]\left[\frac{V_2}{T_2} = \frac{546mL}{378K} = 1.44mL.K^{-1}\right]$$

It shows that $\frac{V_1}{T_1} = \frac{V_2}{T_2} \ldots\ldots \frac{V_n}{T_n}$ (Charles equation)

From the above equation, the Charles law can be defined as "the ratio between the volume and absolute temperature of the given mass of a gas is constant at constant pressure."

Graphical Representation:

When volume is plotted against temperature, a straight line is obtained. It indicates that "V" is directly proportional to the absolute temperature "T".



Example: $800cm^3$ of a gas at 400 torr pressure and 60°C was heated until the volume of the gas was $2000cm^3$ at the same pressure. What is the final temperature of the gas?

Solution:

Initial volume $(V_1) = 800cm^3$

Initial temperature $(T_1) = 60C°$ $K = C° + 273$

$K = 60 + 273 = (333K)$

Final volume $(V_2) = 2000cm^3$

Final temperature $(T_2) = ?$

Using Charles law's equation;

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

$$T_2 = \frac{2000cm^3 \times 333K}{800cm^3} = [832.5K]$$

Q6: Write a note on absolute zero.

Ans: Absolute Zero:

"The hypothetical (ضمنی) / theoretical temperature at which the volume of a gas becomes / considered to be zero is called absolute zero." Absolute zero means zero temperature on absolute [Kelvin] scale. Its value is 0 K or -273°C. This temperature is just theoretical and cannot be achieved practically, because before reaching this temperature, gas will either change to liquid state or solid state.

Quantitative Statement of Charles Law:

In 1787 Jacques Charles and Joseph Gay-Lussac independently discovered quantitative relationship between volume and temperature. Their experiments show that all gases expand to the same extent when heated to the same temperature.

Statement:

"At constant pressure, the volume of the fixed mass of a gas increases by 1/273 of the original volume at 0°C for each degree Celsius (C°) rise in temperature."

Similarly, at constant pressure the volume of the fixed mass of a gas decreases by 1/273 of the original volume at 0°C for each degree Celsius fall in temperature. Equation used to calculate the volume of a gas at various temperatures is as follow:

$$V_T = V_o\left(1 + \frac{T}{273}\right)$$

Where V_T = volume of the gas at temperature T.

V_o = volume of the gas at 0°C

T = temperature on Celsius scale.

Example: The volume of a gas at 0°C is 546cm³. We can calculate the new volume of a gas at various temperature.

At temperature T = 1°C:

$$V_{1C^\circ} = 546\left(1 + \frac{1}{273}\right) = 546\left(\frac{274}{273}\right) = 546(1.003) = 548cm^3$$

Follow LCM rules here

At temperature T = 10°C:

$$V_{10C^\circ} = 546\left(1 + \frac{10}{273}\right) = 546\left(\frac{283}{273}\right) = 566cm^3$$

Similarly for decrease in temperature.

At temperature T = -273°C:

$$V_{-273C^\circ} = 546\left(1 + \frac{(-273)}{273}\right) = 546\left(\frac{0}{273}\right) = 0cm^3$$

Looking to the above calculations, the volume of a gas at 0°C is 546cm³. At 1°C it becomes 548cm³. It means 1°C rise in temperature causes the volume of the gas to increase by 2cm³ and so on. It must

be noted that volume does not increases in direct proportion to the Celsius temperature. For example when temperature is increased ten times i.e. from 10°C - 100°C, the volume does not increase ten times but increases only from 566 - 746cm³. Look at the table.

Celsius temperature C°	Temperature (K)	Volume (cm^3)	V/T = K
273	546	1092 ⟶	$2cm^3 \cdot K^{-1}$
100	373	746 ⟶	$2cm^3 \cdot K^{-1}$
10	283	566 ⟶	$2cm^3 \cdot K^{-1}$
1	274	548 ⟶	$2cm^3 \cdot K^{-1}$
0	273	546 ⟶	$2cm^3 \cdot K^{-1}$
-1	272	544 ⟶	$2cm^3 \cdot K^{-1}$
-10	263	526 ⟶	$2cm^3 \cdot K^{-1}$
-100	173	346 ⟶	$2cm^3 \cdot K^{-1}$
-273	0	0	

It means that Charles law does not obey the Celsius scale temperature. According to the data in table;

$V_1 = 566cm^3$ at $T_1 = 10°C$
$V_2 = 746cm^3$ at $T_2 = 100°C$

Applying Charles equation;

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{566}{10} \neq \frac{746}{100}$$

As a result a new temperature scale has been developed. It starts from -273°C which is called absolute zero.

Relationship between Kelvin Scale and Celsius Scale:

$$K = C° + 273$$

Now according to the data in table;

$V_1 = 566cm^3$ $T_1 = 10°C$ equal to 283K

$V_2 = 746cm^3$ $T_2 = 100°C$ equal to 373K.

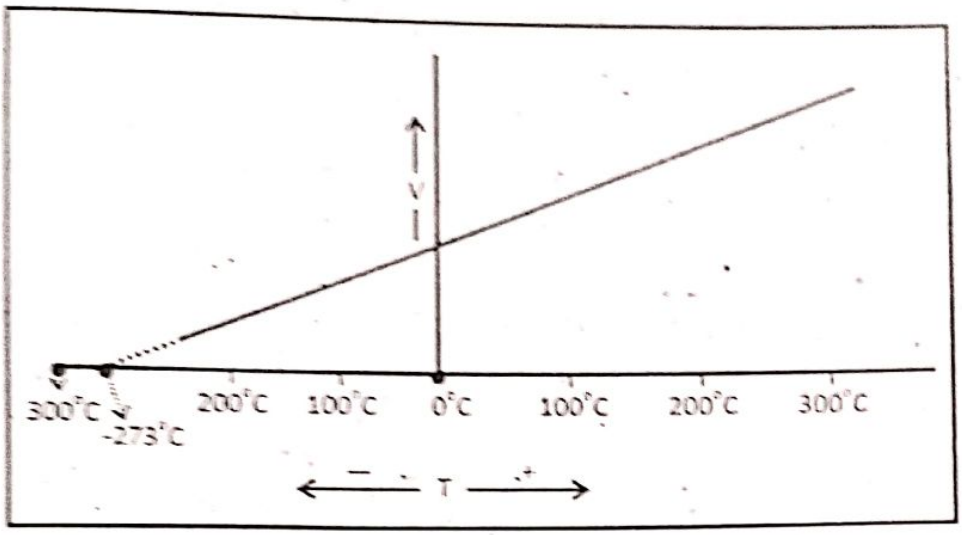
Applying Charles equation;

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{566}{283} = \frac{746}{373} \Rightarrow 2 = 2$$

When temperature is taken in Kelvin scale then Charles law is obeyed.

Graphical Representation:

A graph is plotted between temperature "T" on x-axis and volume "V", for a given mass of a gas at constant pressure, we get a straight line. When we extra plot, the graph upto -273°C, this line cuts the temperature axis at -273°C, which is the lowest possible temperature. The line breaks before reacting the temperature of a gas to -273°C. It means that at this temperature all the gases have zero volume or not in gaseous state. The gases are condensed and converted into liquid or solid. Therefore gas laws cannot be applied on them. All gases liquefies before this temperature.



Absolute temperature must be used in all gas laws problems involving temperature. The B.P, F.P and absolute zero on Fahrenheit, Celsius and Kelvin scales are given below:

	Fahrenheit	Celsius	Kelvin
B.P of water	212F	100C	373K
F.P of water	32F	0C	273K
Absolute zero	-459F	-273C	0K

Q7: State and explain Avogadro's law [1813].
(ETEA: 2008, 2010, 2014-15)

Ans. Statement:

"At the same temperature and pressure equal volumes of different gases contain the same number of molecules or atoms." OR

"The volume of the given [fixed] mass of a gas is directly proportional to the number of moles at constant temperature and pressure."

Mathematical Form:

$V \propto N$ [T and P = constant]

$V = Kn$

Where K = proportionality constant

We know that, 1 mole of any gas at STP occupy a volume of $22.4dm^3$.

OR $22.4dm^3 = 1$ mole

$22.4dm^3 = 6.023 \times 10^{23}$ molecules

$1dm3 = \frac{6.023}{22.4} \times 10^{23}$ molecules

$\boxed{1dm^3 = 0.268 \times 10^{23} \Rightarrow (2.68 \times 10^{22} \text{ molecules})}$

If we have one dm^3 [$1dm^3$] of each, H_2, O_2, N_2 and CH_4 in separate containers at STP, then the number of molecules in each case will be 2.68×10^{22}.

We know that $1dm^3$ of H_2 at STP = 0.0899gms and $1dm^3$ of oxygen gas at STP = 1.44 grams. But their number of molecules is the same. Although oxygen molecules is 16 times heavier than hydrogen but this does not affect volume. It is because gas molecules are far apart and one molecule is approximately at a distance of 300 times of its own diameter from its neighbour at room temperature.

Q8: Derive an ideal gas equation / expression. Write down the applications of ideal gas equation. (ETEA: 2013,15,18,19)

Ans: Ideal Gas Equation / General Gas Equation:

Definition: "The combine equation of four variables i.e. volume, pressure, temperature and number of moles of a gas, is known as ideal gas equation."

Derivation:

General gas equation can be derived by combining Boyle's law, Charles law and Avogadro's law.

According to Boyle's law, "the volume of the fixed mass of a gas is inversely proportional to the applied external pressure at constant temperature."

$$V \propto \frac{1}{P}, \quad \text{[constant temperature]} \_\_\_\_\_ (i)$$

According to Charles's law, "the volume of the given mass of a gas is directly proportional to absolute temperature at constant external pressure."

$V \propto T$ [P = constant] _____ (ii)

According to Avogadro's law, "the volume of the given mass of a gas is directly proportional to the number of moles at constant pressure and temperature."

$V \propto n$ [T, P = constant] _____ (iii)

Combining all the three equations i.e. (i), (ii) and (iii)

$$V \propto \frac{1}{P} \cdot T \cdot n$$

$$V = \text{constant} \cdot \frac{1}{P} \cdot T \cdot n$$

$$V = R \cdot \frac{1}{P} \cdot T \cdot n \quad [R = \text{General gas constant or universal gas constant}]$$

$$PV = RTn$$

$$PV = nRT \_\_\_\_\_ (iv) \quad \text{re-arranging}$$

Equation (iv) is called general gas equation.

Since standard moles for any gas, n = 1 so equation (iv) becomes;

$$PV = RT$$

$$\frac{PV}{T} = R \_\_\_\_\_ (v)$$

The ideal gas equation is useful for problems that do not involve changes in P, V, T and n for a gas sample. Sometimes, however, we need to deal with changes in pressure, volume and temperature or even the number of moles. When conditions changes we must employ a modified form of the ideal gas equation.

For initial states equation (v) can be written as;

$$\frac{P_1V_1}{T_1} = R \_\_\_\_\_ (vi)$$

For final states equation (v) can be written as;

$$\frac{P_2V_2}{T_2} = R \_\_\_\_\_ (vii)$$

Combining equation (vi) and equation (vii),

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

This is the generalized form of ideal gas equation.

EXAMPLE 4·2:

750cm³ of a gas at 300mmHg and 50°C is heated until the volume of gas is 2000cm³ at a pressure of 700mmHg. What is the final temperature of the gas?

Solution:

$V_1 = 750cm^3$ $P_1 = 300mmHg$ $T_1 = 50°C + 273 = 323K$

$V_2 = 2000cm^3$ $P_2 = 700mmHg$ $T_2 = ?$

According the general gas equation;

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

By cross multiplication;

$$P_1V_1T_2 = P_2V_2T_1$$

$$T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

$$T_2 = \frac{700mmHg \cdot 2000cm^3 \cdot 323K}{300mmHg \cdot 750cm^3}$$

$$T_2 = 2010K$$

Applications of General Gas Equation [PV = nRT]:

General gas equation can be used to determine:

1. To calculate units and values of "R" at STP.
2. To calculate units and values of "R" in system international.
3. To calculate mass of a gas (m).
4. To calculate density of a gas (d).
5. To calculate concentration of gas (C).
6. To calculate molecular mass of gas (M).

1. Calculation of Units and Values of "R" at STP:

Standard temperature = T = 0°C → 273K

Standard volume = V = 22·4dm³

Standard pressure = P = 1atm

Standard number of moles = n = 1mole

According to the general gas equation;

$$PV = nRT$$

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

$$R = \frac{1atm \times 22{\cdot}4dm^3}{1mole \times 273K}$$

$$\boxed{R = 0{\cdot}0821\ atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}}$$

2. Calculation of Value of "R" in International System:

In international system of units, it is known that,

Standard temperature = T = 273K ["K" SI unit of temp.]

Standard pressure = P = 1atm → $101325Nm^{-2}$ → SI unit

Amount i.e. number of moles = n = 1 mole

Volume = V = $22{\cdot}4dm^3 \rightarrow 0{\cdot}0224m^3$ [divide 22·4 by 1000]

The general gas equation is, PV = nRT

(OR) $R = \frac{PV}{nT}$

$$R = \frac{101325Nm^{-2} \cdot 0{\cdot}0224m^3}{1mole \cdot 273K}$$

$$R = 8{\cdot}31Nm^{-2} \cdot m^3 \cdot mol^{-1} \cdot K^{-1}$$

$$R = 8{\cdot}31Nm \cdot mol^{-1} \cdot K^{-1}$$

$$R = 8{\cdot}31J \cdot mol^{-1} \cdot K^{-1} \qquad [1Nm = J]$$

$$R = \frac{8{\cdot}31}{4{\cdot}18} Calori \cdot mol^{-1} \cdot K^{-1} \qquad [1Cal = 4{\cdot}18\ joules]$$

$$R = 1{\cdot}987\ Cal \cdot mol^{-1} \cdot K^{-1}$$

$$\boxed{V = 22{\cdot}4dm^3}$$

$$\boxed{V = \frac{22{\cdot}4}{100} = [0{\cdot}0224m^3]}$$

1atm = 760 torrs

1 torr = $133{\cdot}32Nm^{-2}$

760 torrs = 760 × $133{\cdot}32Nm^{-2}$ = [$101325Nm^{-2}$]

So [1atm = $101325Nm^{-2}$]

3. Calculation of Mass of a Gas (m):

PV = nRT

$PV = \frac{m}{M}RT$ As $n = \frac{m \rightarrow mass}{M \rightarrow molar\ mass}$

PVM = mRT

$$\frac{PVM}{RT} = m \quad OR \quad \boxed{m = \frac{PVM}{RT}}$$

4. Calculation of Density of a Gas (d):

$$PV = nRT$$

$$PV = \frac{m}{M}RT \qquad As \left[n = \frac{m}{M}\right]$$

$$PM = \frac{m}{V}RT$$

$$PM = dRT \qquad As \left[d = \frac{m}{V}\right]$$

$$\frac{PM}{RT} = d \quad OR \quad \boxed{d = \frac{PM}{RT}}$$

5. Calculation of Concentration of Gas (C):

$$PV = nRT$$

$$P = \frac{n}{V}RT \qquad As\ C = \frac{n}{V}$$

$$P = CRT \text{ and } \boxed{C = \frac{P}{RT}}$$

6. Calculation of Molecular Mass of Gas (M):

$$PV = nRT$$

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

$$PVM = mRT$$

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

PRACTICE PROBLEM 4.1:

Reading on barometer is 2.21atm. Calculate the corresponding pressure in (a) torr (b) Pascal.

Solution:

(a) We know that 1atm = 760 torr

So 2.21atm will be = 2.21 x 760 torrs = (1679.6 torrs)

(b) We know that,

1 atm = 101325 Pascals

So 2·21 atm = 2·21 x 101325 Pascals

2·21 atm = (223928·25 Pascals)

PRACTICE PROBLEM 4·2:

(i) Find the value of "R" in $mmHg \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$

(ii) Find the value of "R" in $torr \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$

(iii) Find the value of "R" in $torr \cdot cm^3 \cdot mol^{-1} \cdot K^{-1}$

Solution:

Using the general gas equation, $PV = nRT$

OR $R = \frac{PV}{nT}$

$R = \frac{1atm \cdot 22{\cdot}4dm^3}{1mole \cdot 273K}$

$R = 0{\cdot}0821 atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$

(i) $R = 0{\cdot}0821 \times 760 mmHg \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$ [As 1atm = 760mmHg]

$R = 62{\cdot}4\ mmHg \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$

(ii) $R = 62{\cdot}4\ torrs \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$ [As 1mmHg = 1 torr]

(iii) $R = 62{\cdot}4 \times 1000\ torrs \cdot cm^3 \cdot mol^{-1} K^{-1}$ [As $1dm^3 = 1000cm^3$]

$R = 62400\ torrs \cdot cm^3 \cdot mol^{-1} \cdot K^{-1}$

PRACTICE PROBLEM 4·3:

What pressure will be exerted by 0·400 mol of a gas in a $5dm^3$ containers at 17°C?

Solution: Pressure = P = ? n = 0·400 mol V = $500dm^3$

T = 17°C + 273 = 290K R = $0{\cdot}0821\ atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$

We know that,

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

$P = \frac{0{\cdot}400\ \cancel{mol} \cdot 0{\cdot}0821\ atm \cdot \cancel{dm^3} \cdot \cancel{mol^{-1}} \cdot \cancel{K^{-1}} \times 290K}{500\ \cancel{dm^3}}$

(P = 0·0190 atm)

EXAMPLE 4·3:

Calculate the mass of $1dm^3$ of NH_3 gas at 30°C, and 1000mmHg pressure, considering that ammonia is behaving ideally.

Solution:

$V = 1dm^3$ T = 30°C, 30 + 273 = 303K

P = 1000mmHg $\frac{1000}{760} = 1{\cdot}316$ atm

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

Molar mass of $NH_3 = 14 + 3 = 17 g/mol$

By using the relation;

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

$$m = \frac{1.316\ atm \times 1\ dm^3 \times 17\ g \cdot mol^{-1}}{0.0821\ atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1} \times 303 K}$$

$(m = 0.911\ grams)$

EXAMPLE 4.4:

What is the density in $g \cdot dm^{-3}$ of SO_2 at 20°C and 750mmHg pressure?

Solution:

$$P = 750mmHg = \frac{750}{760} = 0.98\ atm$$

$T = 20°C = 20 + 273 = 293K$

Molar mass (M) of $SO_2 = M = 32 + 32 = 64 g/mol$

Using the density formula;

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

$$d = \frac{0.98\ atm \times 64 g \cdot mol^{-1}}{0.0821\ atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1} \times 293K}$$

$(d = 2.60 g \cdot dm^{-3})$

PRACTICE PROBLEM 4.4:

Calculate the density in g·dm-3 of NH3 at 25oC and 1.2atm.

Solution:

$d = ?$ $\quad T = 25°C \rightarrow 25 + 273 = 298K$

$P = 1.2\ atm$ $\quad M_{NH_3} = 14 + 3 = 17 g \cdot mo^{-1}$

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

Using the density formula;

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

$$d = \frac{1.2\ atm \times 17 g \cdot mol^{-1}}{0.0821\ atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1} \times 298K}$$

$(d = 0.833 g \cdot dm^{-3})$

Q9: What are ideal and non-deal gases? Why do real gases [non-ideal] deviate from ideal behaviours? Explain with a graph.

Ans. Ideal Gases: (BISE Swat 2017)

"Gases which obey all gas laws [Boyle's law, Charles law, Avogadro's law] at all temperatures and pressures are called ideal gases." For example, hydrogen, helium and nitrogen are ideal at low pressure and high temperature.

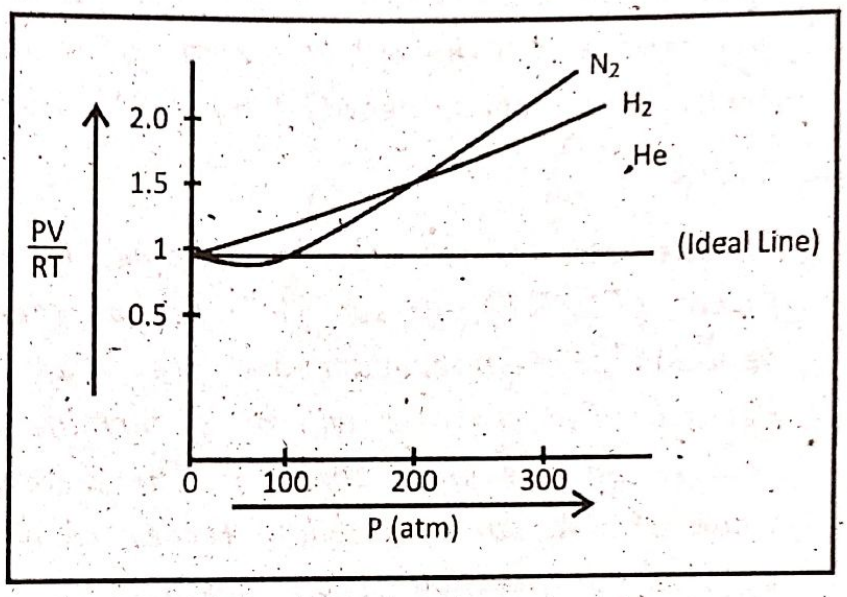
Non-Ideal / Real Gases:

"Gases which do not obey gas laws at all temperature and pressure and called non-ideal / real gases." OR

"Gases which obey gas laws under certain condition of temperature and pressure are called non-ideal gases."

Non-Ideal Behaviour:

"The behaviour of a gas which do not obey gas laws at all temperatures and pressures is called non-ideal behaviour."

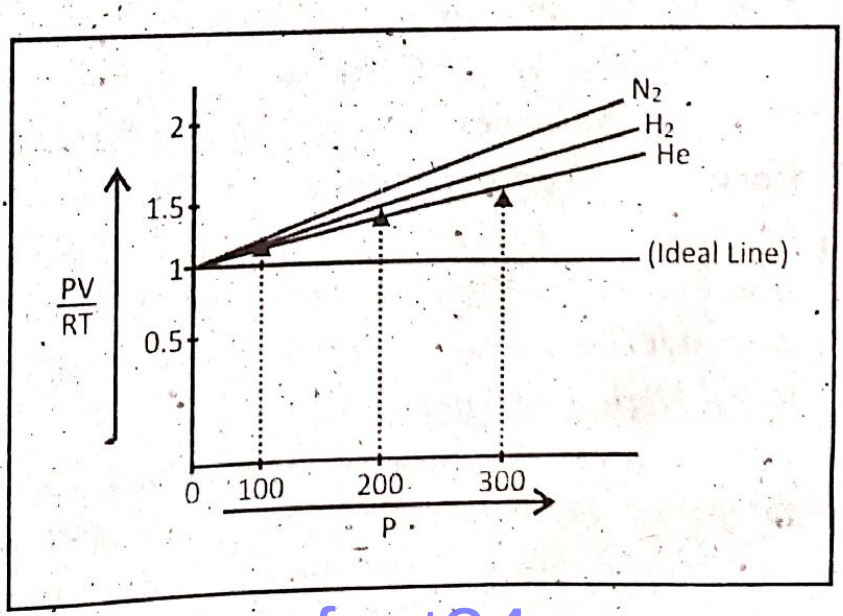


By plotting a graph between PV/RT and pressure the graph is a straight line for an ideal gas but the graph shows deviation for real gases. The graph plotted at 0°C is shown above:

The above graph shows that nitrogen gas shows negative deviation at low pressure but at high pressure all three gases show positive deviation. Now look at the graph drawn at 100°C:

The next graph "at 100°C" shows that at high pressure all the three gases shows positive deviation. The above discussion gives the following results:

1. Gases are ideal at low pressure.
2. Gases are non-ideal at high pressure.



3. Gases are ideal at high temperature.
4. Gases are non-ideal at low temperature.

For an ideal gas PV/RT equals 1 under all conditions. This ratio is called compressibility factor (Z).

$$Z = \frac{\text{Molar volume of gas at same T and P}}{\text{Molar volume of ideal gas at same T and P}} = \left(\frac{PV}{RT}\right)$$

$$\frac{PV}{RT} = \frac{1atm \cdot 22.4dm^3}{0.0821\ atm \cdot dm^3\ mol^{-1} \cdot K^{-1} \times 273K} = 1$$

Compressibility Factor "Z":

"The extent to which a real gas deviate from the ideal behaviour may depicted (shown) in terms of a new function called the compressibility factor, denoted by "Z". Mathematically it is written as;

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

The deviation from ideality may be shown by plotting compressibility factor (PV/RT), versus "P". For an ideal gas, Z = 1 and it is independent of temperature and pressure. The deviation from ideal behaviour of a real gas will be determined by the value of "Z" being greater or less than one (1). The difference between unity and the value of the compressibility factor of a gas is a measure of the degree of non-ideality of the gas.

Causes of Deviation:

The real gases show non-ideal behaviours at low temperature and high pressure. It is due to two incomplete assumptions of kinetic theory of gases.

1. The volume of a single gas molecule is negligible as compared to total volume of the gas.
2. Gas molecules have no force of attraction.

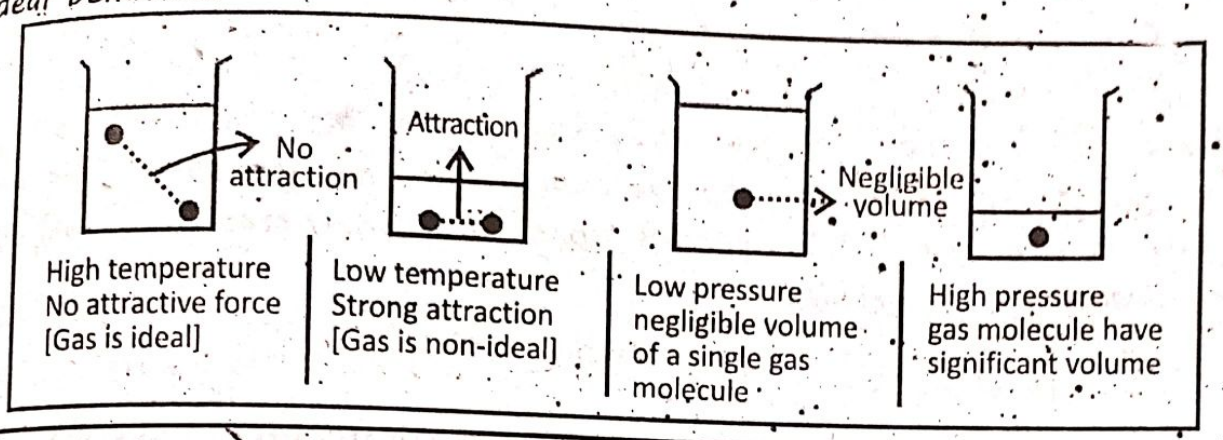
Very Low Temperature:

At very low temperature the K.E and velocity of gas molecules decreases, thus force of attraction increases and the point of KMT does not hold true. Therefore gas shows non-ideal behaviours.

Very High Pressure:

Under normal conditions, the actual volume of a gas is negligible as compared to volume of container. But at very high pressure it is not true because actual volume becomes significant. At high pressure

the gas molecules come closer, empty spaces are reduced. It means that the gas molecules have some volume at high pressure. So the point of KMT is wrong at high pressure. Therefore gas shows non-ideal behaviour.

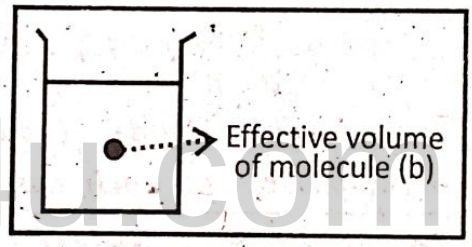


Q10: Derive non-ideal gas equation. The Vander Waal's equation. ETEA: 2014, 2018, 2019

Ans: Vander Waal's Equation: [1873 Dutch scientist]

"A modified [alternative] kinetic equation used to calculate pressure or volume of a real gas under non-ideal conditions is called Vander Waal's equation." Vander Waal suggested that pressure and volume factors in the idea gas equation / general gas equation needs correction in order to make it applicable to real gases.

1. Volume Correction:

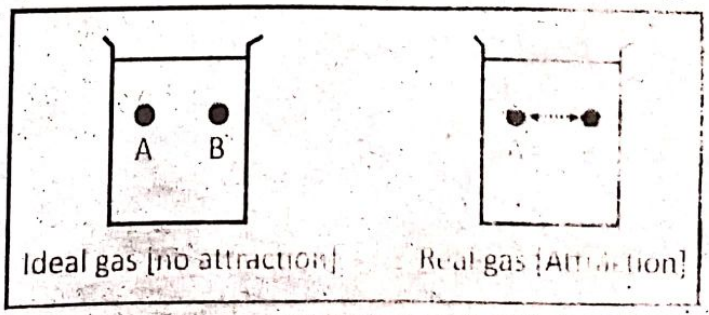


The molecule of an ideal gas has zero volume. Thus the volume of an ideal gas (V) consists of free space and is equal to the volume of the container. However molecules of real gases have a definite volume. This volume cannot be neglected at high pressure. If the effective volume of the gas molecules per mole is "b" then the free space available to gas molecule is: $V - b$

For "n" moles we have; $V_{free} = V - nb$ _____(i)

2. Pressure Correction:

At high pressure and low temperature attraction among gas molecules increases. A molecule striking with the walls of the container is at-



tracted by other molecules. So decreases its velocity. Hence the observed pressure is less than the ideal pressure (P_i) by an amount "P'".

$$P_{observed} = P_i - P'$$

$$P_i = P + P' \quad \text{_____(ii)}$$

Where "P'" is the decrease in pressure due to attraction among real gas molecules at low temperature and high pressure. The decrease in pressure "P'" depends upon the concentration of hitting and dragging of molecules.

$$P' \propto \frac{n}{V} \cdot \frac{n}{V} \qquad \frac{n}{v} = \text{concentration}$$

$$P' = a\frac{n^2}{V^2} \quad \text{_____(iii)} \qquad (a = \text{coefficient of attraction})$$

Putting the value of "P'" in equation (ii)

$$P_i = P + a\frac{n^2}{V^2} \quad \text{_____(iv)}$$

Putting the values of corrected pressure and corrected volume in ideal gas equation [$PV = nRT$], we have,

$$PV = nRT$$

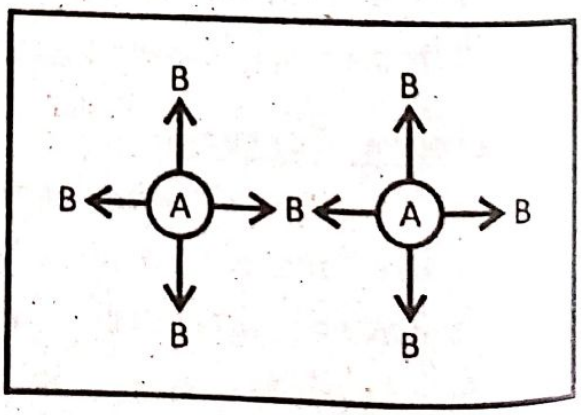
$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT \qquad (\text{when } n = 1 \text{ then})$$

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT \rightarrow \text{Put forward by Vander Waal in 1879.}$$

This is called Vander Waal's equation. The constant "a" and "b" are characteristic of each gas.

Proof of $P' = an^2/V^2$:

Let we have a gas in a container. There are two types of molecules in the mixture. (A) Type molecules are striking the walls and (B) type molecules are pulling them inward. Let C_A and C_B are molar concentrations of (A) type and (B) type molecules.



$$P' \propto C_A \cdot C_B P$$

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

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

If $n = 1$ then $\left(P' = \frac{a}{V^2}\right)$

Unit of "a":

We have, $P' = \frac{an^2}{V^2}$

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

Unit of "b":

$V = nb = 0$

$$b = \frac{V}{n} \longrightarrow \frac{volume}{mole} \longrightarrow [dm^3 \cdot mol^{-1}]$$

Q11: State and explain Daltons law of partial pressure.
ETEA: 2005,2006,2008,2015-17, BISE D.I Khan 2019

Ans: Statement:

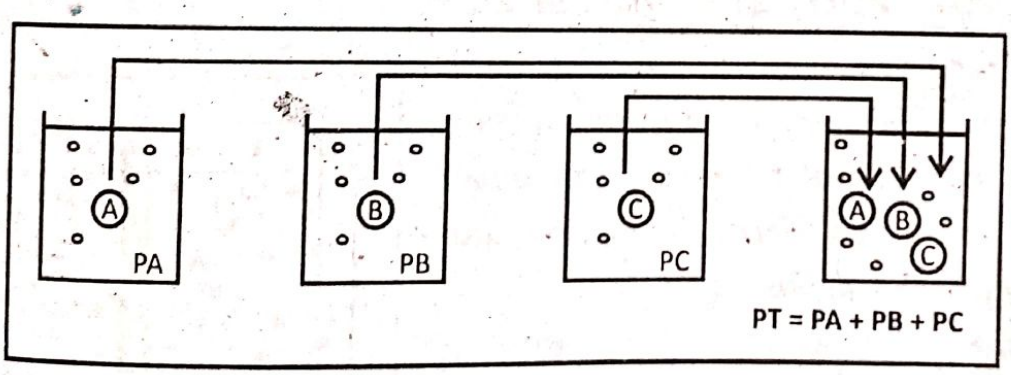
"At constant temperature and volume, the total pressure exerted by a non-reacting gaseous mixture is equal to the sum of the partial pressures of all the gases present in the mixture."

Dalton presented / summarized his observations in 1803.

Partial Pressure:

"The pressure exerted by one kind of gas molecules in a mixture is called partial pressure of that gas."

Let we have three non-reacting gases i.e. A, B, C present / taken in separate containers at the same temperature. Their partial pressures are P_A, P_B, P_C respectively. When these three gases are mixed together in a single container of the same volume at the same temperature, then according to Dalton's law of partial pressure, total pressure (P_T) of the mixture is the sum of the partial pressure of all the gases present in the mixture i.e.



$$P_T = P_A + P_B + P_C$$

Similarly $P_{atm} = P_{N_2} + P_{O_2} + P_{CO_2} + P_{inert\ gases}$

The partial pressure of each gas depends upon the amount / number

of moles of that gas in the mixture. According to general gas equation; $PV = nRT$

$$P = n\frac{RT}{V}$$

$P = n$ [As R, T, V are constant for all gases]

$P \propto n$

It means that higher the number of moles of a gas in the gaseous mixture, greater will be its pressure and vice versa.

Example: If we have a mixture of three non-reacting gases A, B and C, exerting partial pressure of 300torr, 600torr and 450torr respectively, the total pressure will be 1350torr.

$P_T = P_A + P_B + P_C$

P_T = 300torr + 600torr + 450torr = (1350torr)

EXAMPLE 4.5:

A gaseous mixture contains 9.6% NH_3, 22.6% N_2 and 67.8% H_2 gases. If its total pressure is 50atm, find the partial pressure of each gas.

Solution:

Partial pressure of $NH_3 = \frac{9.6}{100} \times 50 = [4.8\ atm]$

Partial pressure of $N_2 = \frac{22.6}{100} \times 50 = [11.3\ atm]$

Partial pressure of $H_2 = \frac{67.8}{100} \times 50 = [33.9\ atm]$

$P_T = P_{NH_3} + P_{N_2} + P_{H_2}$

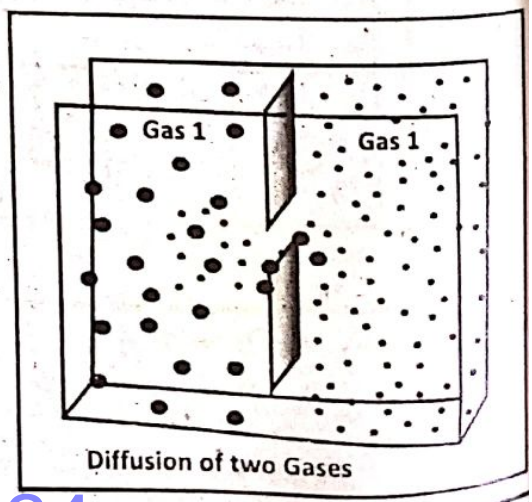
$P_T = 4.8 + 11.3 + 33.9 = [50\ atm]$

Q12: (a) Explain the terms diffusion and effusion.
(b) State and explain Graham's law.
ETEA: 2007, 2017, BISE Mardan 2019

Ans. Diffusion:

Definition: "The random mixing of two or more than two different gas molecules in order to form a homogenous mixture is called diffusion."

Explanation: For example, spreading of fragrance of rose. Although the molecules of the gas move with great / high speed but the diffusion process is

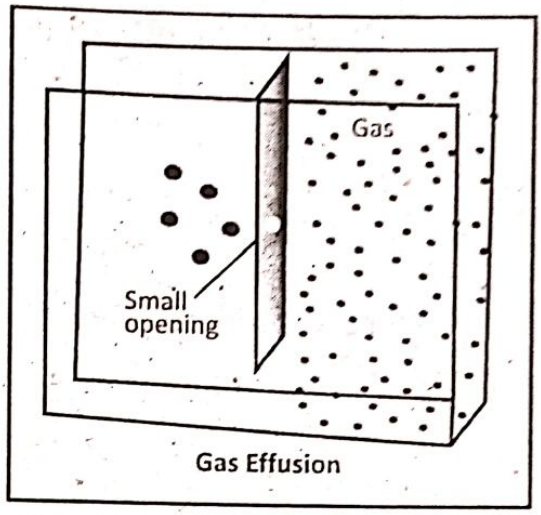


Diffusion of two Gases

slow. It takes relatively long time to occur. For example; when a bottle of perfume is opened at one corner of a room, it take some time before a person, at the other corner of the room can smell it. The reason is that a molecule collides with other molecules on the way reaching to the other corner of the room.

Effusion:

"The escape (نکلنا) of one kind of gas molecules [pure gas] from high pressure region to vacuum [low pressure region] one by one through a hole [hole size is equal to molecule size] in the wall of the container is called effusion."



A general example to differentiate between diffusion and effusion is the bursting and puncturing of a vehicle tyre. During bursting gas molecules spread suddenly and this phenomenon is called diffusion. While during puncturing gas molecules spread by escaping through a small hole, this phenomenon is called effusion.

Rate of Diffusion and Effusion of a Gas:

ETEA: 2005,6,8,17,19, BISE D.I Khan 2017, Malakand 2018, Swat, Abbottabad 2019

"The distance covered / travelled by a gas molecule per unit time is called rate of diffusion or effusion of a gas."

The rate of diffusion of a gas is directly proportional to its energy and inversely proportional to its mass. It means that heavy molecules will diffuse slowly as compared to light molecules.

Graham's Law of Diffusion and Effusion:

In 1832, the Scottish chemist Thomas Graham proposed his law of diffusion of gases.

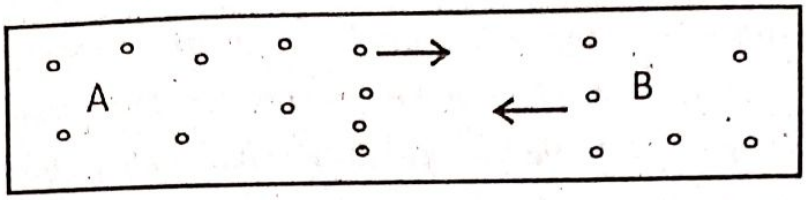
Statement:

"At the same temperature and pressure the rate of diffusion of two different gases is inversely proportional to the square root of their densities or molecular masses."

Mathematical Form:

$$\frac{r_A}{r_B} = \sqrt{\frac{d_B}{d_A}} \quad \text{(i)}$$

Derivation:



According to the statement;

Rate of diffusion of gas-A $r_A \propto \frac{1}{\sqrt{d_A}}$

$$r_A = K \frac{1}{\sqrt{d_A}}$$

$$r_A \times \sqrt{d_A} = K \quad \text{_____(i)}$$

Rate of diffusion of gas-B $r_B \propto \frac{1}{\sqrt{d_B}}$

$$r_B = K \frac{1}{\sqrt{d_B}}$$

$$r_B \times \sqrt{d_B} = K \quad \text{_____(ii)}$$

Divide equation (i) by equation (ii)

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

$$\frac{r_A \times \sqrt{d_A}}{r_B \times \sqrt{d_B}} = \frac{1}{1}$$

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

$$\frac{r_A}{r_B} = \sqrt{\frac{d_B}{d_A}} \quad \text{_____(iii)}$$

As density of gas molecule is directly proportional to its molecular mass so, the above reaction (iii) can also be written as,

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$$

Where; r_A = rate of diffusion of gas-A molecules
r_B = rate of diffusion of gas-B molecules
d_A = density of gas-A molecules
d_B = density of gas-B molecules

M_A = molecular mass of gas-A molecules
M_B = molecular mass of gas-B molecules

Note:
Graham's law mathematical form verification / derivation can also be carried out as;

We know that $d = \frac{m}{V}$ _____ (i)

Where "m" is the molecular mass of the gas.
According to the ideal gas equation;

$PV = nRT$

$PV = RT$ _____ (ii) when $n = 1$

Re-arranging equation (ii) we get;

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

Now put the value of "V" in equation (i)

$d = \frac{m}{RT/P}$ or $d = \frac{mP}{RT}$

For gas-A molecules: $d_A = \frac{m_A P}{RT}$

For gas-B molecules: $d_B = \frac{m_B P}{RT}$

Putting the value of "d_A" and "d_B" in equation (i) i.e.

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

$$\frac{r_A}{r_B} = \sqrt{\frac{m_B P}{RT} \div \frac{m_A P}{RT}}$$

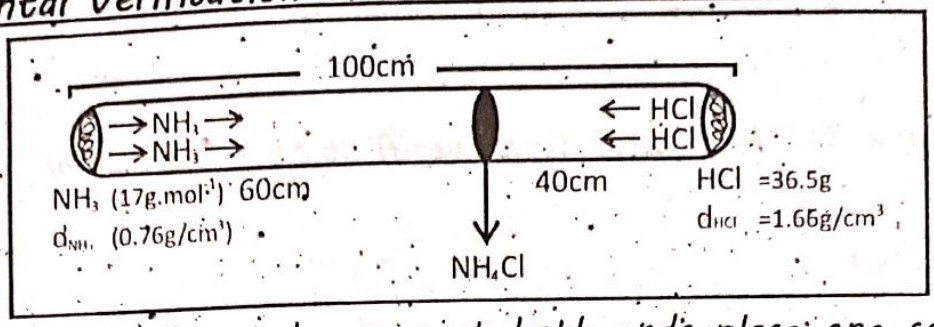
$$\frac{r_A}{r_B} = \sqrt{\frac{m_B P}{RT} \times \frac{RT}{m_A P}}$$

$$\frac{r_A}{r_B} = \sqrt{\frac{m_B}{m_A}}$$ where $m_B = M_B$, $m_A = M_A$

Therefore we can write the above equation as;

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$$ (دونوں طریقوں میں سے جو بھی آسان ہو وہی یاد کرو)

Experimental Verification of Graham Law of Diffusion:



Take 100cm long glass tube open at both ends place one cotton plug soaked (تر) in ammonia solution at one end of the tube and another cotton plug soaked in HCl solution at the other end of the tube simultaneously. The two gases escape from the cotton plugs. A white dense smoke of ammonium chloride [NH_4Cl] is formed at the point where the two gases met.

$$NH_3 + HCl \longrightarrow NH_4Cl$$

Ammonium chloride is formed at a distance of about 60cm from NH_3 end and 40cm from HCl end. The ratio of velocities of two gases depends on their distance travelled by the respective gases. Let

r_{NH_3} = rate of diffusion of ammonia

d_{NH_3} = density of NH_3

M_{NH_3} = molecular mass of NH_3

r_{HCl} = rate of diffusion of HCl

d_{HCl} = density of HCl

M_{HCl} = molecular mass of HCl

$$\frac{\text{Velocity of } NH_3}{\text{Velocity of HCl}} = \frac{\text{distance covered by } NH_3}{\text{distance covered by HCl}}$$

$$= \frac{60cm}{40cm} = (1{\cdot}5) \quad \text{[Experimental result]}$$

It shows that ammonia vapours diffuses 1·5 times faster than HCl vapours.

Applying Graham's law;

$$\frac{r_{NH_3}}{r_{HCl}} = \sqrt{\frac{d_{HCl}}{d_{NH_3}}} = \sqrt{\frac{1{\cdot}66}{0{\cdot}76}} = 1{\cdot}48 \cong \boxed{1{\cdot}5} \text{ theoretical result}$$

$$\frac{r_{NH_3}}{r_{HCl}} = \sqrt{\frac{M_{HCl}}{M_{NH_3}}} = \sqrt{\frac{36{\cdot}5}{17}} = 1{\cdot}48 \cong \boxed{1{\cdot}5} \text{ theoretical result}$$

As experimentally and theoretically, the results are the same, so it verifies Graham's law of diffusion.

EXAMPLE 4.6:

Determine the relative rates of diffusion of equal volumes of H_2 and CO_2 under the same conditions of temperature and pressure.

Solution:

Molecular mass of $H_2 = 2$

Molecular mass of $CO_2 = 44$

According to Graham's law of diffusion;

$$\frac{r_{H_2}}{r_{CO_2}} = \sqrt{\frac{M_{CO_2}}{M_{H_2}}}$$

$$= \sqrt{\frac{44}{2}} = \sqrt{\frac{22}{1}} = \frac{4.69}{1} = \frac{4.7}{1} \Rightarrow \boxed{r_{H_2} = 4.7 r_{CO_2}}$$

The ratios of rates of diffusion of hydrogen and carbon dioxide has been observed as 4.7 = 1. It means that hydrogen molecules diffuse 4.7 times faster than CO_2 under similar conditions.

PRACTICE PROBLEM 4.5:

Calculate the rates of diffusion of carbon dioxide and sulphur dioxide from the same container and at same temperature and pressure.

Solution:

Molecular mass of $CO_2 = 44$

Molecular mass of $SO_2 = 64$

Applying Graham's law;

$$\frac{r_{CO_2}}{r_{SO_2}} = \sqrt{\frac{M_{SO_2}}{M_{CO_2}}} = \sqrt{\frac{64}{44}} = \sqrt{\frac{1.4545}{1}} = 1.2 \quad \boxed{r_{CO_2} = 1.2\, r_{SO_2}}$$

It means that CO_2 diffuses 1.2 times faster than SO_2 under the same conditions.

Q13: Write down the applications of KMT. (BISE Bannu 2019)

Ans: **Root Mean Square Velocity:**

In the equation from the kinetic molecular theory of gases, the average velocity of the gas particles is a special kind of average.

Definition:

"The square root of mean square velocity is called root mean square

velocity." It is represented by u_{rms}.

Explanation:

We know that mean square velocity is the average of squares of all the possible velocities. It is represented by $\bar{u}^2$. Let there are "n" numbers of molecules having velocities V_1, V_2, V_3, V_4.....V_n. Mean square velocity of all the gas molecules will be written as;

$$\bar{u}^2 = \frac{V_1^2 + V_2^2 + V_3^2 + V_4^2 \cdots\cdots V_n^2}{n} \quad \text{_____(i)}$$

We also know that the square root of mean square velocity is known as root mean square velocity. So it can be written as;

$$u_{rms} = \sqrt{\bar{u}^2} \quad \text{_____(ii)}$$

Now putting the value of $\bar{u}^2$ from equation (i) into equation (ii), we get,

$$u_{rms} = \sqrt{\frac{V_1^2 + V_2^2 + V_3^2 + V_4^2 \cdots\cdots V_n^2}{n}}$$

"u" is thus the root mean square velocity or RMS velocity.

Value of RMS Velocity at given Temperature:

The value of RMS velocity at a given temperature can be calculated from the kinetic gas equation.

$$PV = \frac{1}{3} mNu^{-2} \quad \text{_____(1)}$$

Where P = pressure of the gas

V = volume of the gas

m = mass of each molecule

N = number of molecules

u^{-2} = mean square velocity

Solving / re-arranging equation (1), we get,

$$u^{-2} = \frac{3PV}{mN} \quad \text{_____(2)}$$

According to ideal gas equation;

$$PV = nRT$$

For n = 1, $PV = RT$ _____(3)

Putting the value of "PV" from equation (3) into equation (2), we get,

$$u^{-2} = \frac{3RT}{mN}$$

If "N" is Avogadro's number then;

$mN = M$ (molar mass of the gas)

$$u^{-2} = \frac{3RT}{M} \quad \text{_____(4)}$$

We know from equation (ii) that

$$u_{rms} = \sqrt{u^{-2}} \quad \text{_____(5)}$$

Putting the value of "u^{-2}" from equation (4) into equation (5), we get,

$$u_{rms} = \sqrt{\frac{3RT}{M}} \quad \text{_____(6)}$$

Putting the values of R, T and M, the value of rms velocity can be calculated.

2. Derivation of Graham's Law of Diffusion:

Graham law states that "the rate of diffusion of two gases is inversely proportional to the square root of their densities / molecular mass at the same temperature and pressure."

According to KMT of gases "at the same temperature and pressure, the molecules of different gases have the same kinetic energies."

If we have two gases, then K.E of gas-1 molecules will be equal to K.E of gas-2 molecules.

K.E of gas-1 = K.E of gas-2

We know that,

$$K.E_1 = \frac{1}{2} m_1 V_1^2 \quad \text{_____(1) and}$$

$$K.E_2 = \frac{1}{2} m_2 V_2^2 \quad \text{_____(2)}$$

Putting the values of $K.E_1$ and $K.E_2$ from equations (2) and (3) into equation (1), we get,

$$\frac{1}{2} m_1 V_1^2 = \frac{1}{2} m_2 V_2^2$$

Cancelling $\frac{1}{2}$ at both sides of the equation

$$m_1 V_1^2 = m_2 V_2^2$$

$$\frac{V_1^2}{V_2^2} = \frac{m_2}{m_1}$$

Taking square root at both sides of the equation

$$\sqrt{\frac{V_1^2}{V_2^2}} = \sqrt{\frac{m_2}{m_1}}$$

$$\frac{V_1}{V_2} = \sqrt{\frac{m_2}{m_1}}$$

As numerically masses (m_1, m_2) are equal to molecular masses (M_1, M_2) so,

$m_1 = M_1$ and $m_2 = M_2$

$$\frac{V_1}{V_2} = \sqrt{\frac{M_1}{M_2}} \quad \text{_____(1)}$$

We know that,

$V_1 \propto r_1$; $V_1 = Kr_1$

And $V_2 \propto r_2$; $V_2 = Kr_2$

Where r_1 = rate of diffusion of gas-1 and

r_2 = rate of diffusion of gas-2

Putting the values of "V_1" and "V_2" in equation (4)

$$\frac{Kr_1}{Kr_2} = \sqrt{\frac{M_2}{M_1}}$$

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \quad \text{_____(5)}$$

Molar masses of gases are directly proportional to densities of gases;

$M \propto d$

Therefore, $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$ _____(6)

Equations (5) and (6) are the mathematical forms of Graham's law of diffusion.

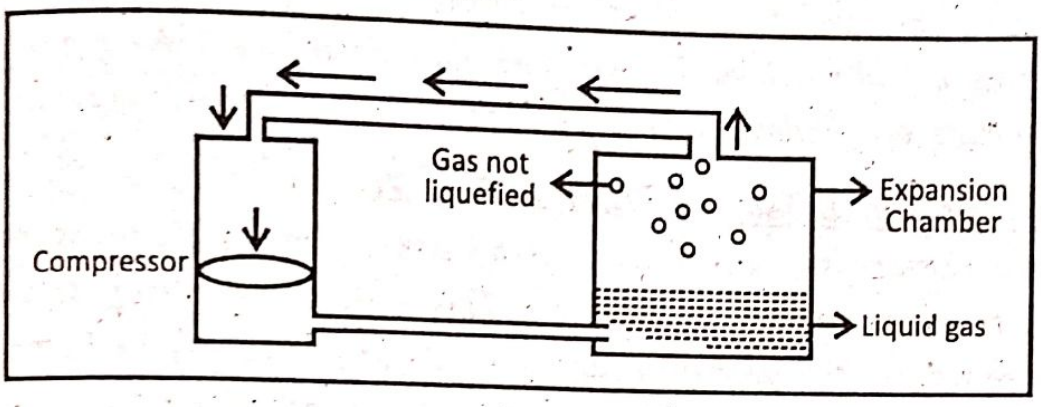
Q14: Write a note on liquefaction of gases. (ETEA: 2019, BISE Mdn 2018)

Ans. Liquifaction of Gases:

Definition: "The process of conversion of a gas into its liquid state by increasing external pressure and decreasing temperature is called liquefaction of gases."

Explanation: A gas can be converted into its liquid state by increasing external pressure and decreasing temperature simultaneously.

When external pressure is applied / increased the gas molecules come closer. By decreasing temperature the K.E decreases due to which attractive forces increases. As a result the molecular motion slows down and the molecules crowded together so closely that the attractive forces between them causes them to condense. The increased pressure is obtained by compressor machines.



Since expansion of gases causes cooling, therefore the low temperature is obtained by releasing the compressed gas to expend. For every gas there exist a temperature above which it cannot be liquefied by applying / increasing pressure alone. It is because K.E is so greater that it can overcome the attractive forces between them.

Critical Temperature (T_c):

"The temperature above which a gas cannot be liquefied by increasing/ applying external pressure alone is called critical temperature."

Critical Pressure (P_c):

"The minimum pressure that is required to liquefy a gas as its critical temperature is called critical pressure."

Critical Volume (V_c):

"The space occupied by one mole of a gas at critical temperature and critical pressure is called critical volume."

The critical temperature, critical pressure and critical volume are called critical constants of the gas.

At critical temperature and critical pressure the gas is in a state which is intermediate between gaseous and liquid state called critical state.

Critical Temperature and Critical Pressure of some Gases			
S·No	Gas	TC	PC
1	Oxygen gas	-118·8C°	47·7atm
2	Nitrogen gas	-147·1C°	33·5atm
3	Hydrogen gas	-239·9C°	12·8atm

The critical temperature of a gas depends upon its size, shape and intermolecular attractive forces. The non-polar gases have low critical temperature and polar gases have high critical temperatures.

Joule-Thomson Effect: (ETEA 2016, 2019)

"When a compressed gas is allowed to enter from high pressure region to low pressure region, it expands. This sudden expansion causes cooling of the gas. The phenomenon of producing lowering of temperature by sudden expansion is known as Joule-Thomson effect."

Joule-Thomson Experiment:

The molecules of a highly compressed gas are very close to each other. These have strong attractive forces. When a gas is expanded suddenly, molecules move away from each other. This requires energy in order to break forces of attraction. This energy is taken from the gas itself, hence gas is cooled down.

An insulated tube is fitted with a porous plug in the middle and two frictionless pistons on the sides. The gas is forced through the porous plug by a slow movement of piston. The gas in the right hand chamber is allowed to expand. The change in temperature is found by taking readings on the two thermometers. Most gases were found to undergo cooling on expansion through the porous plug. Hydrogen and helium were exceptions as these gases showed a warming up instead of cooling.

Lind's Method of Liquefaction of Gases / Air:

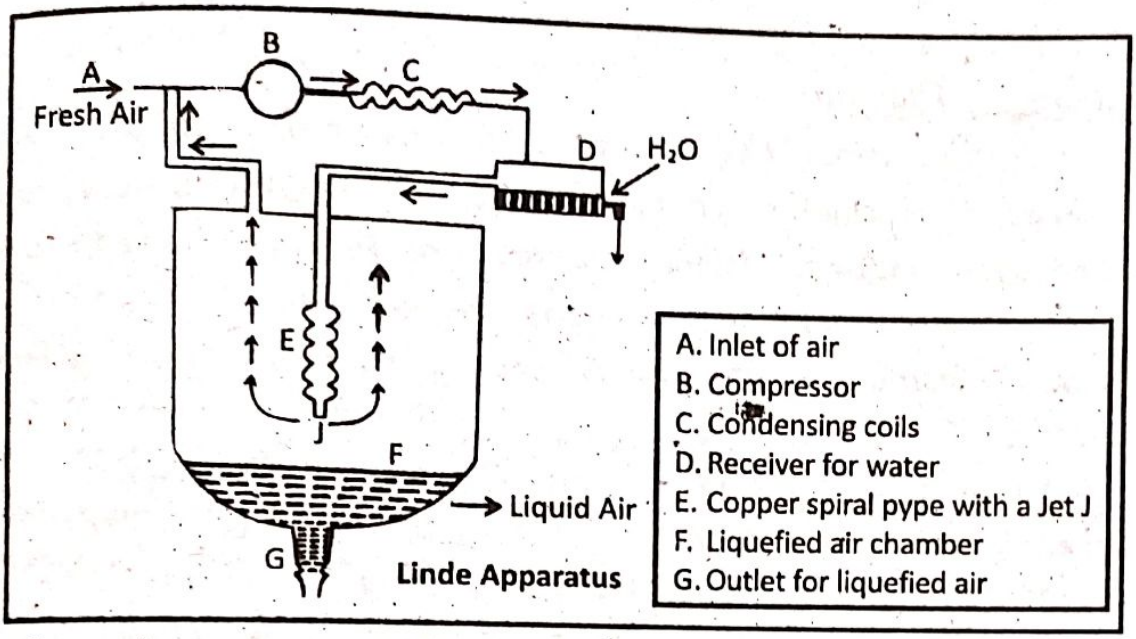
Lind's method is used for liquefaction of all gases except H_2 and helium. This method is usually applied for the liquefaction of air. The apparatus consists of

(1) Compressor
(2) Cooler
(3) Spiral pype

(4) Expansion chamber

This method is based on Joule-Thomson effect.

1. The air is compressed to 200atm by a compressor.
2. Heat is produced during compression of the air.
3. The heat is removed by cooler.
4. Then the compressed air passes through a spiral pype having a jet at the end.



5. When the air comes out of the jet into low pressure area [expansion chamber]. (1atm) it expands freely. The pressure decreases from 200-1atm.
6. Thus air is cooled down due to Joule-Thomson effect.
7. The cooled air moves up, cool the incoming air (gas) of the jet and then again enters into the compression pump, to be compressed again.

By repeating compression and expansion process air is liquefied except H_2 and He.

Note:

Water vapours are condensed and removed at the start.

Q15: Write down a comprehensive note on plasma state of matter.
ETEA: 2014, BISE Mardan 2018, Malakand 2019

Ans: Plasma means shapeless or formless matter. It is the fourth state of matter. It was identified by an English scientist Willium Crooks in 1879. Plasma is a Greek word which means diffused, unclear or semi-transparent matter.

Definition:
"The state of matter, consists of equal number of positive ions, electrons and neutral particles as well is called plasma state."

Origin:
The term plasma was originally applied by Irving Langmuire in 1928, for an ionized mixture. Since an ionized gas cannot exist at room temperature, that is why, it was observed for the 1st time in an electrical discharge at high temperature.

Formation of Plasma:
Plasma is formed after the ionization of atoms or molecules. When a solid is heated, it changes to liquid state. On further heating liquid is converted into vapours. When vapours are further heated, some of them loses electrons and positive ions are formed.

Definition: "Plasma is a mixture of neutral particles, positive ions and free electrons."

There are two main methods for the formation of plasma:

1. When more heat is given to a substance, then most of its atoms / molecules are ionized and plasma is formed.
2. When electric current is passed through a gas [neon etc], the both light and plasma is formed.

Note:
Plasma is neutral as a whole, because it is a mixture of equal number of positive ions, and free electrons [cations = negative electrons]. Due to free moment of electric charges plasma is good conductor of electricity.

Occurrence of Plasma:

1. Plasma is present in the region around sun and stars.
2. About 99% of the known universe is plasma.
3. Sun is 1.5 million kilometer ball of plasma.
4. Occurs in ionosphere [upper region of atm].
5. On the earth plasma is limited to lightening bulbs, flames and neon signs.

Applications:

1. A fluorescent tube is consisting of long glass tube, filled with neon gas. Neon gas in plasma form is used for lightening. The colour of plasma depends upon the gas used.

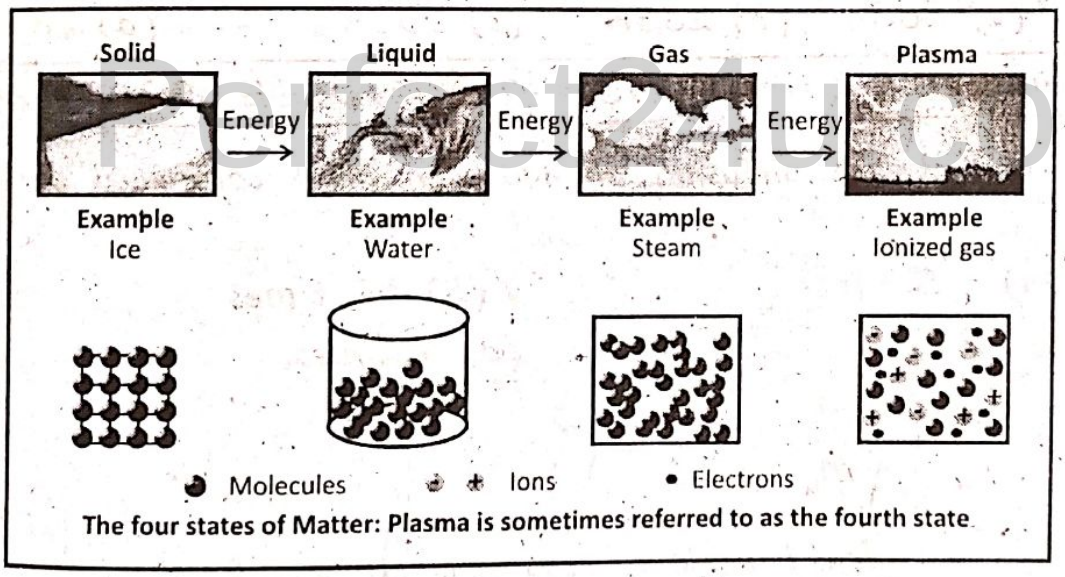
2. Plasma is used in semi-conductors, sterilization of some medical products [surgical instruments], fluorescent lamps, lasers, diamond coated films, high power microwaves sources and pulsed power switches.
3. Used in computers and televisions (TV).
4. Micro-plasma welding is a method which is used to join paper and thin sheets of metals. It is mainly used in the preparation of stainless steel, water storage tanks and kitchen equipments.
5. Cold plasma is used in sterilizers.
6. Used to make plastics attract or repel liquids, which is used in printers.
7. Plasma spray is a process which enable us to coat any material onto any other surface such as;

a. Metal onto Metal: Titanium onto steel to prevent corrosion.

b. Metal onto Non-Metal: Copper onto porcelain used in capacitors.

c. Non-Mental onto Metal: Alumina onto stainless steel, to reduce wear and tear (توڑ پھوڑ) on stainless steel.

d. Non-Metal onto Non-Metal: Teflon onto ceramics to prevent corrosion by acids.



The four states of Matter: Plasma is sometimes referred to as the fourth state

- کسی کو دکھ دینے والا کبھی خوش نہیں رہتا۔ (حضرت ابو بکر صدیق رضی اللہ عنہ)
- کسی کے بے بسی پہ مت ہنسو، کل یہ وقت تم پر بھی آسکتا ہے۔ (حضرت عمر فاروق رضی اللہ عنہ)
- کسی کی آنکھ تمہاری وجہ سے نم نہ ہو کیونکہ تمہیں اُس کے ہر آنسو کا قرض چُکانا ہو گا۔ (حضرت عثمان غنی رضی اللہ عنہ)
- مظلوم اور نمازی کی آہ سے ڈرو کیونکہ آہ کسی کی بھی ہو عرش کو چیر کر اللہ کے پاس جاتی ہے۔ (حضرت علی رضی اللہ عنہ)

Exercise

1: Choose the correct option from the given options:

1. If absolute temperature of a gas is doubled and the pressure reduced to one-half, the volume of the gas will:
   (a) Remain unchanged (b) Be doubled
   (c) Reduced to half ✓(d) Increase four times
2. One dm^3 of hydrogen at STP weighs approximately:
   (a) 0·0789g (b) 0·0799g
   (c) 0·0987g ✓(d) 0·0899g
3. In a factory producing liquid air, one of the pipes carrying dr air at -80°C is blocked with a white solid. This white solid is:
   (a) Argon (b) Ice
   (c) Nitrogen ✓(d) Carbon dioxide
4. The spreading of perfume or scent in air is due to:
   ✓(a) Diffusion (b) Effusion
   (c) Attraction with air (d) Low density
5. A gas has certain volume at 10°C. How much temperature shoul be raised to double its volume?
   ✓(a) 566K (b) 283K (c) 293K (d) 283°C

As 10°C = K = 10 + 273 = 283K → Double of 283 = 566K

6. The rate of diffusion of hydrogen (H_2) compared with helium (He) is:
   (a) 0·5 times ✓(b) 1·4 times
   (c) 2 times (d) 4 times

$$\Rightarrow \frac{r_{H_2}}{r_{He}} = \sqrt{\frac{M_{He}}{M_{H_2}}} = \sqrt{\frac{4}{2}} = \sqrt{\frac{2}{1}} = 1.4$$

7. The non-ideal behaviours results chiefly from:
   (a) Intermolecular attraction and infinite volume
   (b) Elastic collisions and finite volume
   ✓(c) Intermolecular attractions and finite volume
   (d) Intermolecular attraction only

8. The molar volume of helium is $44.8 dm^3$ at:
(a) 100°C and 1atm (b) 25°C and 0·25atm
✓(c) 0°C and 0·5atm (d) 40°C and 0·5atm

9. Which statement about the behaviour of the particles in a gas is not correct?
(a) They are able to move at great speed
(b) Force of attraction between particles are negligible
(c) There is large space among the particles
✓(d) They are arranged in regular pattern

10. At the same temperature and pressure which of the following gases has the greatest density.
(a) CO_2 (b) SO_2 ✓(c) Cl_2 (d) H_2O

11. Weight of $1 dm^3$ of O_2 at STP is:
✓(a) 1·4384gm (b) 1·4394gm
(c) 1·6384g (d) 1·3384g

12. The value of ideal gas constant in $dm^3 \cdot torr \cdot K^{-1} \cdot mol^{-1}$.
(a) 0·0821 (b) 1·98722
✓(c) 62·364 (d) 8·3143

13. 760 torr is equal to:
(a) 760 Pascals (b) 76 Pascals
✓(c) 101325 Pa (d) 1·01325 Pa

14. At 50°C, a gas has 1atm pressure, and $20 dm^3$ volume, its volume at STP will be:
✓(a) $16.94 dm^3$ (b) $10.92 dm^3$
(c) $3.66 dm^3$ (d) $42.2 dm^3$

$T_1 = 50C^\circ = 333K$, $P_1 = 1\ atm$, $V_1 = 20 dm^3$, $V_2 = ?$
$T_2 = 0C^\circ = 273K$, $P_2 = 1\ atm$

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

$$V_2 = \frac{P_1V_1T_2}{T_1P_2} = \frac{1\ atm \times 20 dm^3 \times 273K}{333K \times 1\ atm} = (16.94 dm^3)$$

15. Which of the following gases will have the fastest effusion rate?
✓(a) CH_4 (b) NH_3 (c) CO_2 (d) O_2

Short Questions

Q1: Justify that $1cm^3$ of hydrogen and $1cm^3$ of methane at STP will have same number of molecules, although one molecule of CH_4 is 8-times heavier than that of hydrogen.

Ans: According to Avogadro's law, equal volume of all gases contain equal number of molecules at the same temperature and pressure. The mass and size of molecules do not affect the volume. Thus similarly one molecule is approximately at a distance of 300 times of its own diameter from its near neighbour at room temperature. Thus we may say that $1cm^3$ of H_2 and $1cm^3$ of CH_4 at STP will have the same number of molecules.

Q2: Do you think that some postulates of KMT of gases are faulty? Write down these postulates.

Ans: Yes, I think that two postulates of KMT of gases are faulty. These points are actually incomplete and needs detail. These are;

1. There are no attractive and repulsive forces among gas molecules.
2. Volume of a single gas molecule is negligible as compared to total volume of the container.

Q3: Why high pressure and low temperature make the gas non-ideal?

Ans: When pressure is increased, molecules come closer, the empty spaces among them reduces and thus the volume of a single gas molecule is no longer negligible and possess effective volume. So at high pressure volume of a single molecule cannot be neglected as compared to total volume of the container.

Similarly by decreasing temperature, the K.E of molecules decreases, their motion become slow, and the molecules com closer. When the molecules come closer, attractive forces increases which make the gas non-ideal.

It becomes clear that at low temperature and high pressure gases become non-ideal (real).

Q4: What is your opinion about rapid expansion of gases causes cooling?

Ans: When the compressed gas suddenly expands, the molecules move

away from each others. During expansion, the breaking of intermolecular forces needs energy, which is obtained from the gases itself, energy becomes lower and thus gases are cooled down / liquefied.

Q5: Why do you think that lighter gases can diffuse more rapidly than heavier ones?

Ans. According to Graham's law of diffusion of gases, the rate of diffusion of two gases is inversely proportional to the square root of their densities or molecular masses.

It means that rate of diffusion of a gas depends upon the densities / molecular masses of the gas molecules. Higher the molecular mass, lower will be the rate of diffusion and lower the molecular mass, faster will be the rate of diffusion. It means lighter gases will diffuse quickly as compared to heavier gases.

- NH_3 diffuses 1.5 times faster than HCl.
- H_2 diffuses 1.4 times faster than He (helium).

Numerical Questions

Q1: A mass of gas is under a pressure of 760 torr, and occupies volume of 525cm³. If the pressure is doubled, what volume would the gas now occupy? Assume the temperature is constant.

Solution:

Initial pressure (P_1) = 760torrs
Initial volume (V_1) = 525cm³
Final pressure (P_2) = 1520torrs
Final volume (V_2) = ?

According to Boyle's law,

$$P_1V_1 = P_2V_2$$

$$= \frac{760\text{torrs} \times 525\text{cm}^3}{1520\text{cm}^3} = [262.5\text{cm}^3]$$

Q2: H_2 gas diffuses through a porous plate at a rate of 500cm per minute at 0°C. What is the rate of diffusion of oxygen gas through the same porous plate at 0°C?

Solution:

Rate of diffusion of H_2 = r_{H_2} = 500cm per minute at 0°C

Rate of diffusion of O_2 = r_{O_2} = ? cm per minute at 0°C

Molecular mass of H_2 = M_{H_2} = 2amu

Molecular mass of O_2 = M_{O_2} = 32amu

According to Graham's law of diffusion,

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \text{ or } \frac{r_1}{r_2} = \sqrt{\frac{M_2}{m_1}} \Rightarrow \frac{r_{H_2}}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{H_2}}}$$

$$r_{O_2} = \sqrt{\frac{M_{H_2}}{M_{O_2}}} \times r_{H_2}$$

$$r_{O_2} = \sqrt{\frac{2}{32}} \times 500cm$$

$$\left[r_{O_2} = 125cm \cdot min^{-1}\right]$$

Q3: It was desired to obtain a volume of 1000cm³ of oxygen at 100°C and 640mmHg. How many moles of oxygen would be required?

Solution:

Volume = V = 1000cm³ = 1dm³

Temperature = T = 100°C = 100 + 273 = 373K

Pressure = P = 640mmHg = 0·842atm

Number of moles = n = ?

According to general gas equation;

$$PV = nRT$$

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

$$n = \frac{0{\cdot}842\ \cancel{atm} \times 1\ \cancel{dm^3}}{0{\cdot}0821\ \cancel{atm} \cdot \cancel{dm^3} \cdot mol^{-1} \cdot \cancel{K^{-1}} \times 373K}$$

$$\left[n = 0{\cdot}0275\ mole\right]$$

Q4: Calculate the density of CH_4 at 0°C and 1atm.

Solution:

Density = d = ?

Temperature = T = 0°C ⟶ 273K

Pressure = P = 1atm

Molar mass of CH_4 = 16g/mol

General gas constant = $0.0821 atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1}$

Using the density formula:

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

$$d = \frac{1atm \times 16g \cdot mol^{-1}}{0.0821\ atm \cdot dm^3 \cdot mol^{-1} \cdot K^{-1} \times 273K} = (0.714g \cdot dm^{-3})$$

Q5: A sample of krypton with a volume of $6.25dm^3$ and a pressure of 765torr and a temperature of 20°C is expanded to a volume of $9.55dm^3$ and a pressure of 375torr. What will be its final temperature?

Solution:

$V_1 = 6.25dm^3$

$P_1 = 765torrs \rightarrow 1.006atm$

$T_1 = 20°C \rightarrow 20 + 273K = 293K$

$V_2 = 9.55dm^3$

$P_2 = 375torrs \Rightarrow 0.493atm$

$T_2 = ?$

According to general gas equation;

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

$P_1V_1T_2 = P_2V_2T_1$ [Cross multiplication]

$$T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

$$T_2 = \frac{0.493\ atm \times 9.55dm^3 \times 393K}{1.006\ atm \times 6.25dm^3}$$

$[T_2 = 219.40K]$

Descriptive Questions

Q1: (a) Explain the concept of diffusion and effusion of gases.

(b) Apply the knowledge of kinetic molecular theory of gases and derive an expression of Graham's law of diffusion.

(c) The rate of effusion of an unknown gas 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.

Ans (a) Please see question number 12

(b) Please see question number 13 part (2)

(c) Rate of effusion of an unknown gas = $r_x = 0.279$

Rate of effusion of hydrogen = $r_{H_2} = 1$

Molecular mass of unknown gas = $M_x = ?$

Molecular mass of hydrogen = $M_{H_2} = 2amu$

According to Graham's law of diffusion;

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \frac{r_{H_2}}{r_x} = \sqrt{\frac{M_x}{M_{H_2}}}$$

Taking square at both sides

$$= \frac{r^2 H_2}{r^2 x} = \frac{M_x}{M_{H_2}}$$

$$M_x = \frac{r^2 H_2}{r^2 x} \times M_{H_2}$$

$$M_x = \frac{1^2}{(0.279)^2} \times 2 \Rightarrow \frac{1}{0.0778} \times 2 \Rightarrow 12.85 \times 2$$

$$M_x \Rightarrow [25.70\ amu]$$

So molecular mass of the unknown as is 25.7amu which is approximately equal to 26amu.

Q2: (a) State Dalton's law of partial pressure. Write down the applications of Dalton's law.

(b) Explain significance of absolute zero, giving its value in degree Celsius and Kelvin.

(c) What pressure is exerted by a mixture of 2grams of helium (He), 16grams of oxygen (O_2) and 10grams of carbon dioxide (CO_2) at 10°C in a 5dm³ vessel?

Ans (a) Dalton's Law:

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

Applications of Dalton's Law:

1. When a gas is collected over the surface of water, it becomes wet and water vapours mixes with the gas molecules and forms a homogenous mixture. The pressure of dry gas can be calculated by using Dalton's law.

$P_T = P_{Gas} + P_{w.v}$

$P_{Gas} = P_T - P_{w.v}$ [w.v → water vapours]

2. The respiration of animals depends upon partial pressure of gases. The partial pressure of oxygen gas in human's lung is 116g/cm³ and in air it is 159g/cm³. So oxygen moves from high pressure region (air) into low pressure region (lungs). Similar is the case for CO_2 gas, which comes out of the lungs because its partial pressure inside the lungs is higher than that in air.

3. At high altitudes, partial pressure of oxygen is low [150g/cm³], so pilots feels uncomfortable in breathing. Thus pilots must use pressurized cabins.

4. Deep see divers, use mixture of oxygen and any inert gas for breathing. The reason is that after every 100 feet depth, the pressure increases three times than normal. Moreover due to increase in pressure, nitrogen diffuses into blood. Due to the above reasons, divers cannot use normal air for breathing.

(b) Absolute zero is important because, when a gas is cooled to absolute zero its molecules stop moving / molecular motion ceased up. This is a theoretical concept. It is the lowest possible temperature where nothing could be colder and no heat energy remains in a substance.

Absolute zero [0 K] = -273°C and -459°F

(c) Solution:

Mass of He = 2grams

Mass of O_2 = 16grams

Mass of CO_2 = 10g

$n_{H_2} = \frac{2}{4} = \frac{1}{2} \Rightarrow$ [0.5 moles]

$n_{O_2} = \frac{16}{32} = \frac{1}{2} \Rightarrow$ [0·5 moles]

$n_{CO_2} = \frac{10}{44} =$ [0·227 moles]

Total number of moles (n_T) = 0·5 + 0·5+0·227=[1·227 moles]

Volume of the container = V = 5dm³

Temperature = T = 10°C + 273 = 283K

According to ideal gas equation, PV = nRT

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

$$P = \frac{1{\cdot}227 \text{ mole} \times 0{\cdot}0821 \text{ atm}{\cdot}dm^3{\cdot}mol^{-1}{\cdot}K^{-1} \times 283K}{5dm^3} = [5{\cdot}70 \text{ atm}]$$

Q3: (a) How would you relate temperature to the average K·E of the particles in a substance?
(b) Explain Lind's method for the liquefaction of gases.
(c) Give two points of evidence to show that gases do not behave ideally under all conditions of temperature and pressure.

Ans. (a) As stated in the kinetic molecular theory of gases that average kinetic energy of molecules is directly proportional to absolute temperature. So temperature and average K·E of molecules are directly related. When a substance absorbs heat, the molecules move faster so the average kinetic energy, and therefore the temperature increase.

Actually temperature is the measurement of the average kinetic energy of the molecules in a system.

(b) Please see question number 14.

(c) All gases are ideal under certain conditions of temperature and pressure but gases do not behave ideally under all conditions of temperature and pressure. It is because:

1. At very low temperature attractive forces among gas molecules exist and gases show non-ideality at low temperature.
2. At high pressure the volume of a single gas molecules cannot be neglected and has some volume [effective volume], so gases show non-ideality at high pressure.

Q4: (a) Write the Vander Waal equation for real gas. Clearly explain the meaning of the corrective terms for pressure and volume.
(b) The temperature of a real gas usually drops when it is allowed to enter into a low pressure vacuum [explain].
(c) Define and describe the properties of plasma.

Ans (a) Please see question number 10.

(b) The temperature of a real gas usually drops when it is allowed to enter into a low pressure region, it is because, the molecules of highly compressed gases are very close and have very strong forces of attraction. When this gas suddenly expands, then molecules move apart and energy is required to break intermolecular forces / attractions. This energy is taken from the gas itself, thus temperature of the gas drops and cool down.

(c) Please see question number 15 for this section.

Q5: (a) State basic postulates of KMT of gases.
(b) Explain how gas laws can be derived from kinetic molecular theory of gases.
(c) Explain absolute zero on the basis of Charles law.

Ans (a) Please see question number 2.

(b) Gas laws can be derived from KMT as;

(i) Derivation of Boyle's Law:

According to KMT, pressure of the gas is due to the number of collisions of molecules per unit area. When volume of the gas is reduced by one half of its original volume at constant temperature. The number of collisions on the walls of the container becomes double. It means that pressure of the gas became double. This is the statement of Boyle's law, that when volume decreases pressure increases, i.e.

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

(ii) Charles Law:

According to KMT, average K.E of gas molecules is directly proportional to absolute temperature. By increasing temperature, K.E of molecules increases, which will exert more pressure on the walls of the container as well as upon the piston. As a result the piston moves upward and thus the volume will be increased. It shows that

volume is directly proportional to absolute temperature, which is the statement of Charles law. This relation can be represented as;

$T \propto K.E \propto$ molecular motion $\propto$ no of collisions $\rightarrow$ internal pressure $\rightarrow$ volume

(iii) Dalton's Law of Partial Pressure:

According to KMT, there are no attractive and repulsive forces among gas molecules. Therefore molecules of each gas in the mixture behave independently. These will move freely and will exert their own pressure [individual / partial pressure]. Thus total pressure of the non-reacting gases will be equal to the sum of the partial pressures of all the gases present in the mixture, which is the statement of Dalton's law.

(iv) Graham's Law of Diffusion:

According to KMT, at the same temperature, molecules of every gas have the same average kinetic energy. Let we have two gases, then K.E of gas-1 molecules will be equal to K.E of gas-2 molecules.

$$K.E_1 = K.E_2$$

$$\frac{1}{2} m_1 V_1^2 = \frac{1}{2} m_2 V_2^2$$

$$\frac{V_1^2}{V_2^2} = \frac{m_2}{\cancel{2}} \times \frac{\cancel{2}}{m_1}$$

$$\frac{V_1^2}{V_2^2} = \frac{m_2}{m_1}$$

Taking square root at both sides of the equation,

$$\frac{V_1}{V_2} = \sqrt{\frac{m_2}{m_1}}$$

This equation shows that, the average velocity of gas molecules is inversely proportional to the square root of their masses, which is the statement of Graham's law. As numerically masses are equal to molecular masses so,

$$\frac{V_1}{V_2} = \sqrt{\frac{M_2}{M_1}} \text{ or } \sqrt{\frac{d_2}{d_1}}$$

(c) Please see question number 6.

If you rest you will rust. If you work, then you will shine.
When you shine, the world will give you credit.