Ideal Gases

Introduction :

Matter as we know broadly exists in three states.

There are always two opposite tendencies between particles of matter which determine the state of matter.

- Inter molecular attractive forces.
- The molecular motion / random motion.

Matter

Properties	Solid state	Liquid state	Gaseous state
Attractive force	● large	● Smaller	● Almost zero
Thermal motion	● Almost zero	● Greater	● Random motion
Volume	● Fixed volume	● Fixed volume	● varies with container
Geometry	● Definite	● Not definite	● Not definite

In this chapter the properties and behaviour of the gases will be analysed and discussed in detail. These properties are measured with the help of the gas laws as proposed Boyle, Charles, Gay lussac etc.

Section (A) : Ideal gas equation & gas laws

Th1 Boyle's law and measurement of pressure :

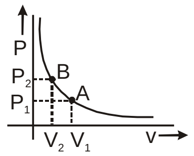
D1 **Statement :** For a fixed amount of gas at constant temperature, the volume occupied by the gas is inversely proportional to the pressure applied on the gas or pressure of the gas.

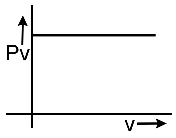
$$V \propto \frac{1}{P}$$

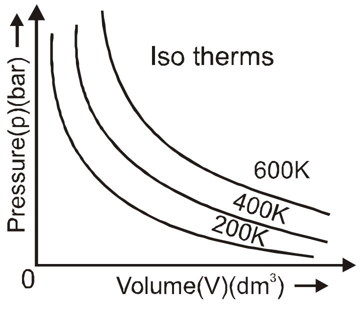
Hence PV = constant.

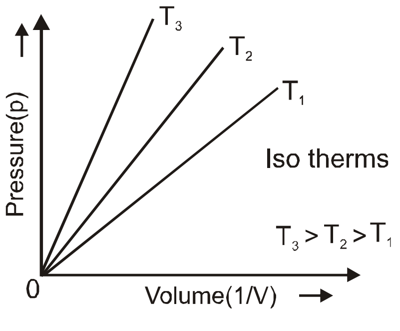
This constant will be dependent on the amount of the gas and temperature of the gas.

F1 $P_1V_1 = P_2V_2$









Application of Boyles Law: For the two points 'A' and 'B' $P_1V_1 = K$ & $P_2V_2 = K$

Hence it follows that $P_1V_1 = P_2V_2$.

Units

Volume	Pressure	Temperature
Volume of the gas is the Volume of the container S.I. unit → m^3 C.G.S. unit → cm^3 $1\ \ell = 10^{-3}\ m^3$ $1\ \ell = 10^3\ cm^9$ $1\ dm^3 = 1\ \ell = 10^{-3}\ m^3$ $1\ ml = 10^{-3}\ \ell = 1\ cm^3 = 1cc$	Pressure = N/m^2 = Pa → S.I. unit C.G.S unit = dyne-cm^2 Convert $1N/m^2$ into dyne/cm^2 $\frac{1\,N}{1\,m^2} = \frac{10^5\,dyne}{10^4\,cm^2}$ $1\ N/m^2 = 10\ dyne/cm^2$ $1\ atm = 1.013 \times 10^5\ N/m^2$ $1\ bar = 1 \times 10^5\ N/m^2$ 1 atm = 760 mm of Hg = 760 torr	Kelvin scale → Boiling point = 373 K ice point = 273 K Farenheit scale → B.P. = 212ºF ice point = 32ºF Celcius scale → B.P. = 100ºC ice point = 0ºC $\frac{C-0}{100-0} = \frac{K-273}{373-273} = \frac{F-32}{212-32} = \frac{R-R(0)}{R(100)-R(0)}$ where R = Temperature on unknown scale.

Atmospheric pressure :

The pressure exerted by atmosphere on earth's surface at sea level is called 1 atm.

1 atm = 1.013 bar

$1\ atm = 1.013 \times 10^5\ N/m^2 = 1.013\ bar = 760\ torr$

Solved Examples

Ex-1. A rubber balloon contains some solid marbles each of volume 10 ml. A gas is filled in the balloon at a pressure of 2 atm and the total volume of the balloon is 1 litre in this condition. If the external pressure is increased to 4atm the volume of Balloon becomes 625 ml. Find the number of marbles present in the balloon.

Sol. Let the no. of marbles be = n.

volume of marble = 10 n ml.

volume of balloon earlier = 1000 ml.

later = 625 ml.

Now for the gas inside the balloon temperature and amount of the gas is constant, hence Boyles law can be applied

$$P_1V_1 = P_2V_2$$

$$4\times (625 - 10n) = 2 \times (1000 - 10n)$$

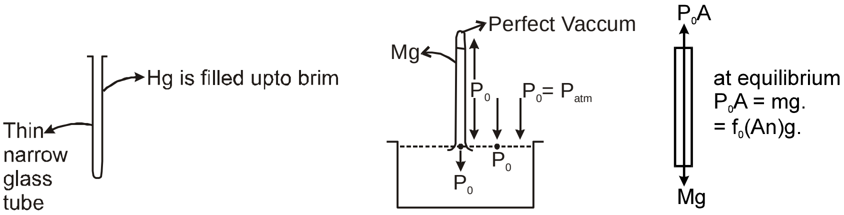
$$625 \times 4 = 2000 - 20n + 40n$$

$$625 \times 4 - 2000 = 20n$$

$$\frac{625 \times 4 - 2000}{20} = n. \qquad \frac{125}{5} = n \quad ; \qquad n = 25$$

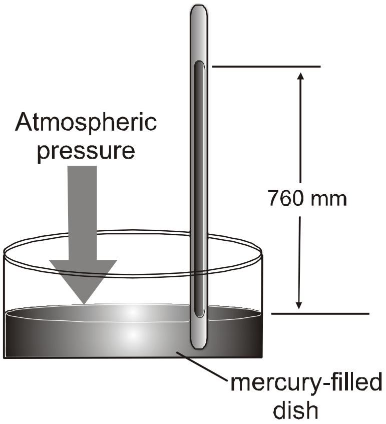
Th2 Measurement of Pressure

Barometer: A barometer is an instrument that is used for the measurement of pressure.The construction of the barometer is as follows



Cross sectional view of the capillary column

A mercury barometer is used to measure atmospheric pressure by determining the height of a mercury column supported in a sealed glass tube. The downward pressure of the mercury in the column is exactly balanced by the outside atmospheric pressure that presses down on the mercury in the dish and pushes it up the column.



A thin narrow calibrated capillary tube is filled to the brim, with a liquid such as mercury, and is inverted into a trough filled with the same fluid.Now depending on the external atmospheric pressure, the level of the mercury inside the tube will adjust itself, the reading of which can be monitored. When the mercury column inside the capillary comes to rest, then the net forces on the column should be balanced.

Applying force balance, we get,

$P_{atm} \times A = m \times g$ ('A' is the cross-sectional area of the capillary tube)

If 'ρ' is the density of the fluid, then $m = \rho \times v$

Hence, $P_{atm} \times A = (\rho \times g \times h) \times A$ $(v = A \times h)$

('h' is the height to which mercury has risen in the capillary)

F2 or, $P_{atm} = \rho gh$

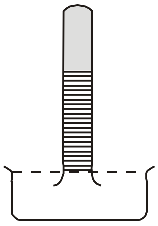
☞ Normal atmospheric pressure which we call 1 atmosphere (1 atm), is defined as the pressure exerted by the atmosphere at mean sea level. It comes out to be 760 mm of Hg = 76 cm of Hg. (at mean sea level the reading shown by the barometer is 76 cm of Hg)

$1 \text{ atm} = (13.6 \times 10^3) \times 9.8 \times 0.76 = 1.013 \times 10^5$ Pascal.

1 torr = 1 mm of Hg.

$1 \text{ bar} = 10^5 \text{ N/m}^2$ (Pa)

Faulty Barometer: An ideal barometer will show a correct reading only if the space above the mercury column is vacuum, but in case if some gas column is trapped in the space above the mercury column, then the barometer is classified as a faulty barometer. The reading of such a barometer will be less than the true pressure.



For such a faulty barometer

$P_0A = Mg + P_{gas} A$

$P_0 = \rho gh + P_{gas}$ or $\rho gh = P_0 - P_{gas}$

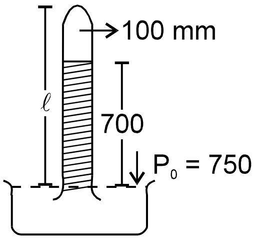
Solved Examples

Ex-2. The reading of a faulty barometer is 700 mm of Hg. When actual pressure is 750 mm of Hg. The length of the air column trapped in this case is 10 cm. Find the actual value of the atmospheric pressure when reading of this barometer is 750 mm of Hg. Assume that the length of the Barometer tube above mercury surface in the container remains constant.

Sol. $P_0 = P_{gas} + 700\ \rho g$

$\therefore$ $P_{gas} = 750\ \rho g - 700\ \rho g = 50\ \rho g$

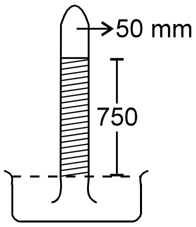
Now for the gas column in the capillary, amount and temperature are constant hence $P_1V_1 = P_2V_2$



$(50\ \rho g)\ (100\ A) = P'_{gas} \times (50\ A)$

$\therefore \quad P'_{gas} = 100\ \rho g$

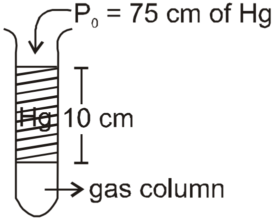
Now, applying force balance in the new conditions:



$P'_{atm} = P'_{gas} + 750\ \rho g = 100\ \rho g + 750\ \rho g = 850\ \rho g$

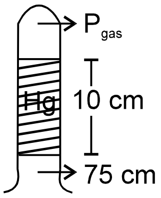
Hence, the atmospheric pressure is now, 850 mm of Hg.

Ex-3. In each of the following examples, find the pressure of the trapped gas.



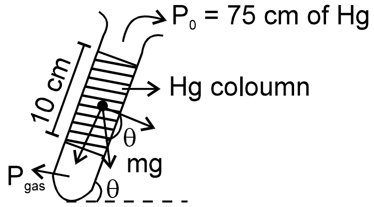
Sol. Total pressure of gas column = 75 + 10 = 85 cm of Hg.

Ex-4.



Sol. P_{gas} = 65 cm of Hg.

Ex-5.



$P_g = 75 + 10 \cos \theta.$

Sol. From the above problem, it can be generalised that, applying force balance every single time is not necessary. If we are moving up in a fluid, then substract the vertical length, and while moving down add the vertical length.

D2 Charles law :

For a fixed amount of gas at constant pressure volume occupied by the gas is directly proportional to temperature of the gas on absolute scale of temperature.

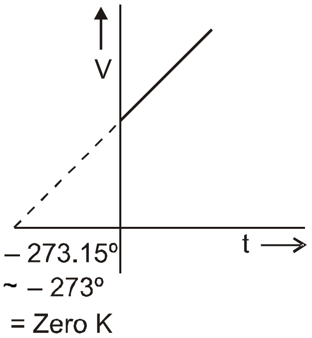
$V \propto T$ or $V = kT$

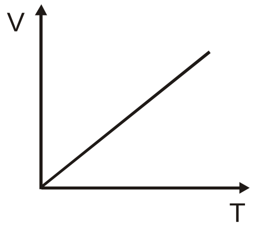
$\frac{V}{T}$ = constant where 'k' is a proportionality constant and is

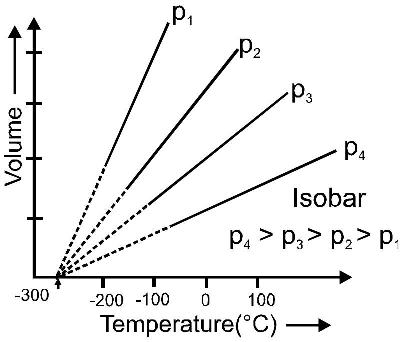
Dependent on amount of gas and pressure.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ Temperature on absolute scale, Kelvin scale or ideal gas scale.

$V = a + bt$ Temperature on centigrade scale.







Relation : $T = t + 273$

- Since volume is proportional to absolute temperature. The volume of a gas should be theoretically zero at absolute zero temperature.
- Infact no substance exists as gas at a temperature near absolute zero, though the straight line plots can be extra plotted to zero volume. Absolute zero can never be attained practically though it can be approached only.
- By considering –273.15°C as the lowest approachable limit, Kelvin developed temperature scale which is known as absolute scale.

Solved Examples

Ex-6. If the temperature of a particular amount of gas is increased from 27ºC to 57ºC, find final volume of the gas, if initial volume = 1 lt and assume pressure is constant.

Sol. $\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{1}{(273+27)} = \frac{V_2}{(273+57)}$ So $V_2 = 1.1$ lt.

Ex-7. An open container of volume 3 litre contains air at 1 atmospheric pressure. The container is heated from initial temperature 27ºC or 300 K to tºC or (t + 273) K the amount of the gas expelled from the container measured 1.45 litre at 17ºC and 1 atm.Find temperature t.

Sol. $\therefore \quad T_1 = 300$ K

It can be assumed that the gas in the container was first heated to (t + 273), at which a volume 'ΔV' escaped from the container.

Hence applying charles law : $\frac{3}{300} = \frac{3+\Delta V}{t+273}$

Now, this volume 'ΔV' which escapes when the container get cooled

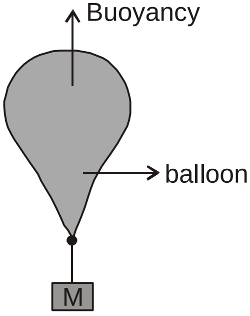
$\therefore \quad \frac{\Delta V}{t+273} = \frac{1.45}{290}$

Solve the two equations and get the value of ΔV and t.

Determine Δ V & calculate t that will be the answer.

Th3 Calculation of pay load :

Pay load is defined as the maximum weight that can be lifted by a gas filled balloon.



For maximum weight that can be lifted, applying force balance

$F_{buoyancy} = M_{balloon} \times g + M_{pay\ load} \times g$

$\Rightarrow \quad \rho_{air}$ v.g. $= \rho_{gas}$ v.g + Mg + mg.

mass of balloon = m
volume of balloon = v
density of air = ρ_{air}
density of gas inside the balloon = ρ_{gas}

net force on balloon = 0
(at equilibrium / when balloon is incoming with constant speed)

Solved Examples

Ex-8. A balloon of diameter 20 m weights 100 kg. Calculate its pay-load, if it is filled with He at 1.0 atm and 27ºC. Density of air is 1.2 kg m^{-3}. [R = 0.0082 dm^3 atm K^{-1} mol^{-1}]

Sol. Weight of balloon = 100 kg = 10×10^4 g

Volume of balloon = $\frac{4}{3}\pi r^3 = \frac{4}{3} \times \frac{22}{7} \times \left(\frac{20}{2} \times 100\right)^3 = 4190 \times 10^6$ cm^3 = 4190×10^3 litre

Weight of gas (He) in balloon = $\frac{PVM}{RT} = \frac{1 \times 4190 \times 10^3 \times 4}{0.082 \times 300} = 68.13 \times 10^4$ g $\left(\because \quad PV = \frac{w}{M}RT\right)$

$\therefore$ Total weight of gas and balloon = $68.13 \times 10^4 + 10 \times 10^4 = 78.13 \times 10^4$ g

Weight of air displaced = $\frac{1.2 \times 4190 \times 10^6}{10^3} = 502.8 \times 10^4$ g

$\therefore$ Pay load = wt. of air displaced – (wt. of balloon + wt. of gas)

$\therefore$ Pay load = $502.8 \times 10^4 - 78.13 \times 10^4$ = **424.67×10^4 g**

D3 Gay-lussac's law :

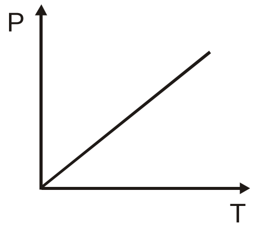
For a fixed amount of gas at constant volume, pressure of the gas is directly proportional to temperature of the gas on absolute scale of temperature.

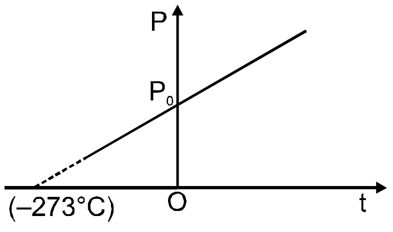
$P \alpha T$

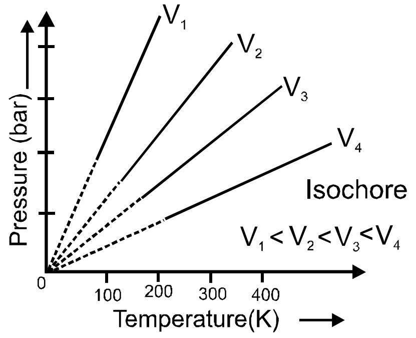
$\frac{P}{T}$ = constant → dependent on amount and volume of gas

F3 $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ → temperature on absolute scale

Originally, the law was developed on the centigrade scale, where it was found that pressure is a linear function of temperature $P = P_0 + bt$ where 'b' is a constant and P_0 is pressure at zero degree centigrade.







Ex. PV = K $\Rightarrow$ V = K$_1$/P

$\frac{V}{T} = K_2 \Rightarrow V = K_2 T$

$\frac{K_1}{P} = K_2 T$

$\underbrace{PT = \frac{K_1}{K_2}}$ = const. $\Rightarrow P \alpha = \frac{1}{T} \Rightarrow$?

where are we wrong ?

This is wrong because we are varying temperature &
K$_1$ = f(1) thus K$_1$ will change according to temperature

So $\frac{K_1}{K_2}$ will be a function of temp & not constant.

Solved Examples

Ex-9. The temperature of a certain mass of a gas is doubled. If the initially the gas is at 1 atm pressure. Find the % increase in pressure ?

Sol. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$; $\frac{1}{T} = \frac{P_2}{2T}$

% increase = $\frac{2-1}{1} \times 100 = 100\%$

Ex-10. The temperature of a certain mass of a gas was increased from 27°C to 37°C at constant volume. What will be the pressure of the gas.

Sol. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$; $\frac{P}{300} = \frac{P_2}{310}$; $P_2 = \frac{31}{30}P$

D4 Avogadro's Hypothesis :

For similar values of pressure & temperature equal number of molecules of different gases will occupy equal volume.

$N_1 \longrightarrow V$ (volume of N_1 molecules at P & T of one gas)

$N_1 \longrightarrow V$ (volume of N_1 molecules at P & T of second gas)

⇒ Molar volume & volume occupied by one mole of each and every gas under similar conditions will be equal.

One mole of any gas or a combination or gases occupies 22.413996 L of volume at STP.

The previous standard is still often used, and applies to all chemistry data more than decade old, in this definition ***Standard Temperature and Pressure STP*** *denotes the same temperature of 0°C (273.15K), but a slightly higher pressure of 1 atm (101.325 kPa) .*

Standard Ambient Temperature and Pressure (SATP)*, conditions are also used in some scientific works. SATP conditions means 298.15 K and 1 bar (i.e. exactly 10^5 Pa) At SATP (1 bar and 298.15 K), the molar volume of an ideal gas is 24.789 L mol^{-1}* ***(Ref. NCERT)***

Th4 Equation of State :

Combining all the gas relations in a single expression which discribes relationship between pressure, volume and temperature, of a given mass of gas we get an expression known as equation of state.

$\frac{PV}{T}$ = constant (dependent on moles of the gas n).

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

Ideal gas Equation : $\frac{PV}{nT}$ = constant [universal constant]

= R (ideal gas constant or universal gas constant)

R = 8.314 J/Kmole ≈ 25/3
= 1.987 cal/mole ≈ 2
= 0.08 Latm/mole ≈ 1/12

Solved Examples

Ex-11. Some spherical balloons each of volume 2 litre are to be filled with hydrogen gas at one atm & 27°C from a cylinder of volume 4 litres. The pressure of the H_2 gas inside the cylinder is 20 atm at 127°C. Find number of balloons which can be filled using this cylinder. Assume that temperature of the cylinder is 27°C.

Sol. No. of moles of gas taken initially = $\frac{20 \times 4}{R \times 400}$ = 2.43 L

No. of moles of gas left in cylinder = $\frac{1 \times 4}{R \times 300}$ = 0.162L

No. of moles of gas to be filled in balloons = 2.43 – 0.162 = 2.268

Let we have 'n' balloons that we can fill

No. of moles of gas that can be filled in 1 balloon = $\frac{1 \times 2}{0.082 \times 300}$ = 0.081

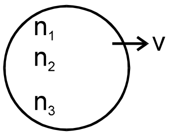
$\therefore$ 0.081 × n = 2.268

n = 28 balloons.

Section (B) : Daltons law of partial pressures

Th5 Daltons law of partial pressure :

Partial pressure :



In a mixture of non reacting gases partial pressure of any component of gas is defined as pressure exerted by this component if whole of volume of mixture had been occupied by this component only.

Partial pressure of first component gas

$P_1 = \frac{n_1RT}{v}$; $P_2 = \frac{n_2RT}{v}$; $P_3 = \frac{n_3RT}{v}$

Total pressure = $P_1 + P_2 + P_3$.

D5 Daltons law :

For a *non reacting* gaseous mixture total pressure of the mixture is the summation of partial pressure of the different component gases.

F4 $P_{Total} = P_1 + P_2 + P_3 = \frac{(n_1 + n_2 + n_3)RT}{v}$

$\frac{P_1}{P_T} = \frac{n_1}{n_T}$ = x_1 (mole fraction of first component gas)

$\frac{P_2}{P_T} = \frac{n_2}{n_T}$ = x_2 (mole fraction of second component gas)

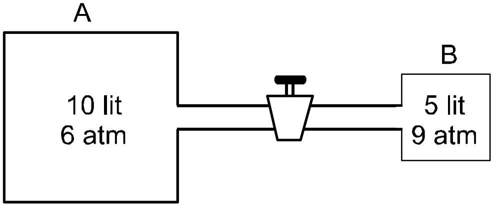
$\frac{P_3}{P_T} = \frac{n_3}{n_T}$ = x_3 (mole fraction of third component gas)

Section (C) : Mixing of Gases

Solved Examples

Ex-12. The stop cock connecting the two bulbs of volume 5 litre and 10 litre containing as ideal gas at 9 atm and 6 atm respectively, is opened. What is the final pressure if the temperature remains same.

Sol. After the opening of the stop cock the pressure of the each bulb will remain same.



At the beginning, the no. of moles of gas in A = $\frac{10 \times 6}{RT}$

At the beginning, the no. of moles of gas in B = $\frac{5 \times 9}{RT}$

$\therefore$ Total no. of mole at the beginning = $\frac{105}{RT}$

Total no. of mole of gas before opening the stop cock

Total no. of moles of gas after opening stop cock = $\frac{105}{RT}$

∴ Pressure after the opening of the stop cock

$$P = \frac{105}{RT} \times \frac{RT}{V_{total}} = \frac{105}{10+5} = 7 \text{ atm}$$

Th6 Analysis of gaseous mixture :

D6 **Vapour Density :** Vapour density of any gas is defined as the density of any gas with respect to density of the H_2 gas under identical conditions of temperature T and pressure P.

$$\text{Vapour density} = \frac{\text{density of gas at T \& P}}{\text{density of } H_2 \text{ under same P \& T}}$$

$$P = \frac{m}{V}.\frac{RT}{M} \quad \Rightarrow \quad P = \rho\frac{RT}{M} \qquad \rho = \frac{PM}{RT}$$

$$\text{Vapour density} = \frac{PM_{gas}RT}{RT \quad PM_{H_2}} = \frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2}$$

F5 $M_{gas} = 2 \times$ vapour density

Average molecular mass of gaseous mixture :

total mass of the mixture divided by total no. of moles in the mixture

$$M_{mix} = \frac{\text{Total mass of mixture}}{\text{Total no. of moles in mixture}}$$

If we have 'n_1', 'n_2' and 'n_3' are moles of three different gases having of molar mass 'M_1', 'M_2' and 'M_3' respectively.

$$M_{min} = \frac{n_1M_1 + n_2M_2 + n_3M_3}{n_1 + n_2 + n_3}$$

Solved Examples

Ex-13. Calculate the mean molar mass of a mixture of gases having 7 g of Nitrogen, 22 g of CO_2 and 5.6 litres of CO at STP.

Sol. Moles of N_2 = 7/28 = 1/4

Moles of CO_2 = 22/44 = 1/2

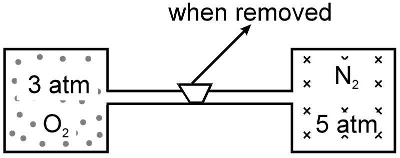
Moles of CO = 5.6 / 22.4 = 1/4

Mean molar mass = $M_{min} = \frac{n_1M_1 + n_2M_2 + n_3M_3}{n_1 + n_2 + n_3}$ = (7 + 7 + 22) / 1 = 36

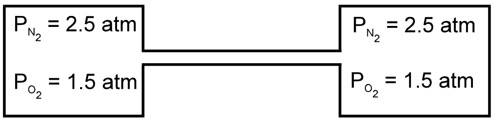
Section (D) : Graham's law of diffusion

Th7 Graham's Law of Diffusion/Effusion :

D7 **Diffusion :** Net spontaneous flow of gaseous molecules from region of high concentration (higher partial pressure) to the region of lower concentration or lower partial pressure



⇌ flow will be from both sides, N_2 will try to equalise its partial Pressure in both the vessels, and so will O_2.



Graham's Law :

"Under similar conditions of temperature and pressure (partial pressure) the rate of diffusion of different gases is inversely proportional to square root of the density of different gases."

- Rate of diffusion $r \propto \frac{1}{\sqrt{d}}$ d = density of gas

F6 $$\frac{r_1}{r_2} = \frac{\sqrt{d_2}}{\sqrt{d_1}} = \frac{\sqrt{M_2}}{\sqrt{M_1}} = \sqrt{\frac{V.D_2}{V.D_1}}$$ V.D is vapour density

r = volume flow rate = $\frac{dV_{out}}{dt}$

r = mole flow rate = $\frac{dn_{out}}{dt}$

r = distance travelled by gaseous molecule per unit time = $\frac{dx}{dt}$

- The general form of the grahams law of diffusion can be stated as follows, when one or all of the parameters are varied.

F7 rate $\propto \frac{P}{\sqrt{TM}} A$; P – Pressure, A – area of hole, T – Temp. , M – mol. wt.

- If partial pressure of gases are not equal.
  Then rate of diffusion is found to be proportional to partial pressure & inversely proportional to square root of molecular mass.

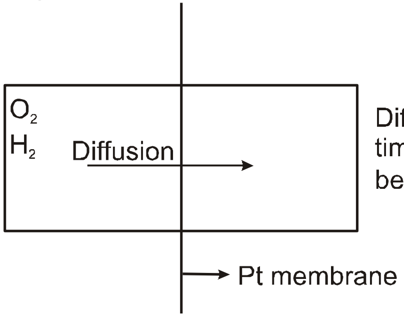
$r \propto P$

$r \propto \frac{1}{\sqrt{M}}$; $\frac{r_1}{r_2} = \frac{P_1}{P_2}\sqrt{\frac{M_2}{M_1}}$

Selective diffusion :

If one or more than one components of a mixture are allowed to diffuse and others are not allowed then it is selective diffusion of those components.

although total pressure may be different in both sides.



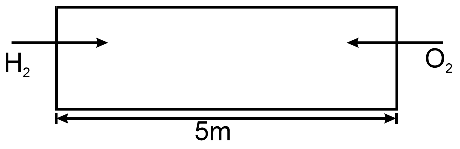
Diffusion will happen till the time partial pressure of H_2 becomes same on both sides.

☞ **Platinum allows only H_2 gas to pass through**

Effusion : (forced diffusion) a gas is made to diffuse through a hole by the application of external pressure.

Solved Examples

Ex-14. In a tube of length 5 m having 2 identical holes at the opposite ends. H_2 & O_2 are made to effuse into the tube from opposite ends under identical conditions. Find the point where gases will meet for the first time.



Sol. $\frac{r_1}{r_2} = \frac{ax}{dt} \times \frac{dt}{dx} = \sqrt{\frac{m_2}{m_1}} = \frac{dx_1}{dx_2} = \sqrt{\frac{32}{2}}$

$\frac{dx_1}{dx_2} = 4 \quad \Rightarrow \quad \frac{\text{distance travelled by } H_2}{\text{distance travelled by } O_2} = 4$

$\frac{x}{(5-x)} = 4 \quad \Rightarrow \quad x = (5 - x)\,4 \quad \Rightarrow \quad x = 20 - 4x$

$5x = 20 \quad \Rightarrow \quad x = 4$ from H_2 side

Ex-15. Assume that you have a sample of hydrogen gas containing H_2, HD and D_2 that you want to separate into pure components (H = ^1H and D = ^2H). What are the relative rates of diffusion of the three molecules according to Graham's law ?

Sol. Since D_2 is the heaviest of the three molecules, it will diffuse most slowly, and let we call its relative rate 1.00. We can then compare HD and H_2 with D_2.

Comparing HD with D_2, we have

$$\frac{\text{Rate of HD diffusion}}{\text{Rate of } D_2 \text{ diffusion}} = \sqrt{\frac{\text{Molecular mass of } D_2}{\text{Molecular mass of HD}}} = \sqrt{\frac{4.0 \text{ amu}}{3.0 \text{ amu}}} = 1.15$$

Comparing H_2 with D_2 we have

$$\frac{\text{Rate of } H_2 \text{ diffusion}}{\text{Rate of } D_2 \text{ diffusion}} = \sqrt{\frac{\text{Mass of } D_2}{\text{Mass of } H_2}} = \sqrt{\frac{4.0 \text{ amu}}{2.0 \text{ amu}}} = 1.41$$

Thus, the relative rates of diffusion are H_2(1.41) > HD (1.15) > D_2(1.00).

Section (E) : Kinetic theory of gases

Th8 Kinetic Theory of Gases :

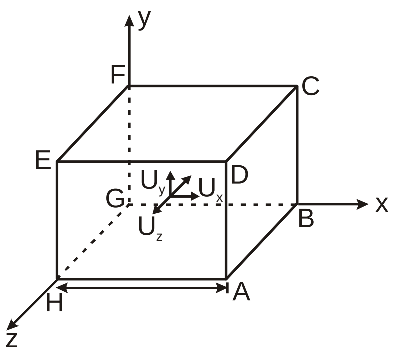
Postulates / assumptions of KTG :

- A gas consists of tiny spherical particles called molecules of the gas which are identical in shape & size (mass)
- The volume occupied by the molecules is negligible in comparision to the total volume of the gas.
  For an ideal gas, volume of the ideal gas molecule $\simeq$ 0.
- Gaseous molecules are always in random motion and collide with other gaseous molecules & with the walls of the container.
- Pressure of the gas is due to these molecular collisions among themselves and with walls of the container
- These collisions are elastic in nature
- Molecular attraction forces are negligible. Infact, for an ideal gas attractive or repulsive forces are equal to zero.
- Newton's laws of motion are applicable on the motion of the gaseous molecules.
- Effect of gravity is negligible on molecular motion.
- The average K.E. of gaseous molecules is proportional to the absolute temperature of the gas.

$\frac{1}{2} M(\overline{u^2}) \propto T$ (bar is for average)

Kinetic equation of gaseous state (expression for pressure of gas).

Der.1 m = mass of one molecule



$\vec{u} = u_x\hat{i} + u_y\hat{j} + u_z\hat{k}$

Consider collision with face ABCD

inital $\vec{P}_i = mu_x\,\hat{i}$; final $\vec{P}_f = -\,mu_x\,\hat{i}$

change in momentum due to collision $= 2\,U_x\,m$

time taken between two successive collision with face ABCD = $t = \dfrac{2\ell}{\mu_x}$

frequency of collisions (f) = $\dfrac{1}{t} = \dfrac{U_x}{2\ell}$

change in momentum in one sec. = force = $2\,m\,\dfrac{U_x \times U_x}{2\ell} = \dfrac{2mU_x^2}{\ell}$

force due to all the molecules = $\dfrac{m}{\ell}\,\{U_{x_1}^2 + U_{x_2}^2 + \ldots\ldots.. + U_{x_N}^2\}$

average value of $U_N^2 = \overline{U_N^2} = \dfrac{Ux_1^2 + U_{x_2}^2 + \ldots\ldots\ldots..U_{x_N}^2}{N}$

$$F_x = \frac{m}{\ell}\,\{N\ \ \overline{U_x^2}\}$$

all the three directions are equal as the motion is totally random in all directions, hence

$$\overline{U_x^2} = \overline{U_y^2} = \overline{U_z^2}$$

$$\overline{U^2} = \frac{U_1^2 + U_2^2 + U_3^2 + \ldots\ldots. + U_N^2}{N} = \overline{U_x^2} + \overline{U_y^2} + \overline{U_z^2} = 3\,\overline{U_x^2}$$

$$F = \frac{m}{\ell}.N\,\frac{1}{3}\,\overline{U^2}$$

Pressure = $\dfrac{F_x}{\ell^2} = \dfrac{1}{3}\,\dfrac{mN}{\ell^3}\,\overline{U^2}$ The volume of the container 'V' = $\square^3$

$\therefore$ $PV = \dfrac{1}{3}\,mN\,\overline{U^2}$ Kinetic equation of gases

where $\overline{U^2}$ is mean square speed

root mean square speed = $U_{rms} = \sqrt{\overline{U^2}} = \sqrt{\left(\dfrac{U_1^2 + U_2^2 + U_3^2 + \ldots\ldots\ \ U_N^2}{N}\right)}$

Th9 Verification of Gaseous Laws Using Kinetic Equation :

F8

- From postulates ; $PV = \dfrac{1}{3}\,mN\,\overline{U^2}$

  $\dfrac{1}{2}\,m\overline{U^2} \propto T = \lambda\,T$ Where 'λ' is a proportionality constant

  $PV = \dfrac{2}{3}\left(\dfrac{1}{2}m\overline{U^2}\right)N$; $PV = \dfrac{2}{3}\,\lambda\,NT$ (N = Total number of molecules)

- Boyle's Law : N : constant
  T : constant
  PV = constant

- Charles law : N : constant
  P : constant
  $V \propto T$

- Kinetic energy of gaseous molecule (translational K.E.)
  To calculate λ we have to use ideal gas equation (experimental equation)
  $PV = nRT$

  kinetic equation $PV = nRT = \dfrac{2}{3}\,\lambda\,NT = \dfrac{2}{3}\,\lambda\,(nN_A)\,T$ (n = number of moles of gas)

on comparing $\lambda = \dfrac{3}{2} \times \dfrac{R}{N_A}$

$\lambda = \dfrac{3}{2}\,K$ where $K = \dfrac{R}{N_A}$ = Boltzmann constant

F9 Average K.E. of molecules = $\frac{1}{2}m\overline{U^2} = \lambda T$

Average K.E. = $\frac{3}{2}KT$ (only dependent on temperature not on nature of the gas.)

F10 Average K.E. for one mole = $N_A\left(\frac{1}{2}m\overline{U^2}\right) = \frac{3}{2}KN_AT = \frac{3}{2}RT$

- **Root mean square speed :**

  $U_{rms} = \sqrt{\overline{U^2}} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{mN_A}}$ Where m-mass of one molecule

❍ Dependent on nature of gas i.e. mass of the gas

F11 $U_{rms} = \sqrt{\frac{3RT}{M}}$ M = molar mass

- **Average speed :**

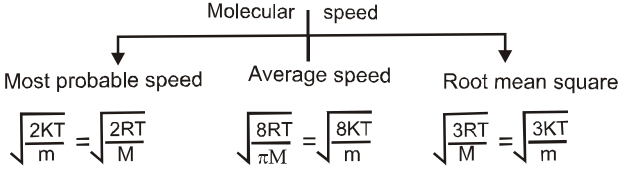
  $U_{av} = U_1 + U_2 + U_3 + \ldots\ldots\ldots\ldots U_N$

F12 $U_{av} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{\pi m}}$ K is Boltzmman constant

- **Most probable speed:** The speed possessed by maximum number of molecules at the given temperature

F13 $U_{mp} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2KT}{m}}$

Velocity can be decribed by maximum number of molecule.



$$U_{mp} : U_{av} : U_{rms} = \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3}$$

Solved Examples

Ex-16. In a container of capacity 1 litre there are 10^{23} molecules each of mass 10^{-22} gm. If root mean square speed is 10^5 cm/sec then calculate pressure of the gas.

Sol. $PV = \frac{1}{3}mNU_{rms}^2$

$P = ?$; $V = 10^{-3}\ m^3$

$m = 10^{-25}$ kg ; $N = 10^{23}$

$\sqrt{\overline{U^2}} = 10^5$ cm/sec $= 10^3$ m/sec

$\overline{U^2} = 10^6\ m^2/sec^2$

$P \times 10^{-3} = \frac{1}{3} \times 10^{-25} \times 10^{23} \times 10^6$

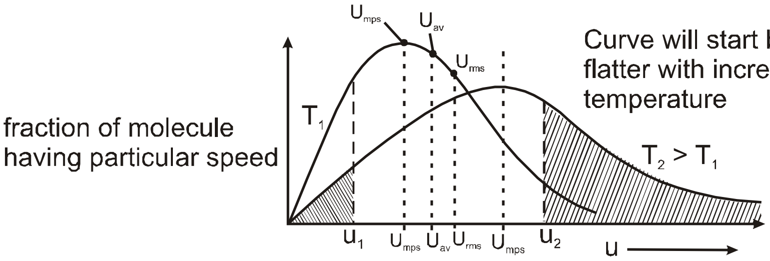
$P = \frac{1}{3} \times 10^{-2} \times 10^6 \times 10^3$; $P = \frac{1}{3} \times 10^7$ pascals

---

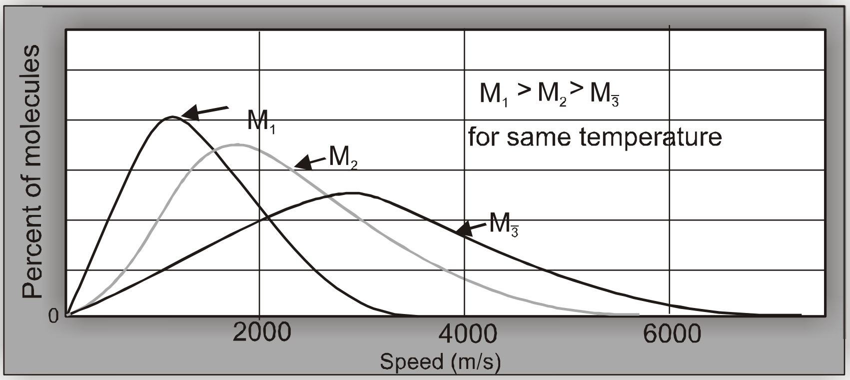
Th10 Maxwell's distributions of molecular speeds :

Postulates/Assumptions of speed distributions

- **It is based upon theory of probability.**
- **It gives the statistical averages of the speed of the whole collection of gas molecules.**
- Speed of gaseous molecules of may vary from 0 to ∞. The maxwell distribution of speed can be plotted against fraction of molecules as follows.



- The area under the curve will denote fraction of molecules having speeds between zero and infinity
- Total area under the curve will be constant and will be unity at all temperatures.
- Area under the curve between zero and u_1 will give fraction of molecules racing speed between 0 to u_1. This fraction is more at T_1 and is less at T_2.
- The peak corresponds to most probable speed.
- At higher temperature, fraction of molecules having speed less than a particular value decreases.
- For Gases with different molar masses will have following graph at a given temperature.



Section (F) : Eudiometry

Th11 Eudiometry :

The analysis of gaseous mixtures is called eudiometry. The gases are identified by absorbing them in specified and specific reagents.

Some Common Facts :

- Liquids and solutions can absorb gases.
- If a hydrocarbon is burnt, gases liberated will be CO_2 & H_2O. [H_2O is seperated out by cooling the mixture & CO_2 by absorption by aqueous KOH]
- If organic compound contains S or P, then these are converted into SO_2 & P_4O_{10} by burning the organic compound.
- If nitrogen is present, then it is converted into N_2.
  [The only exception : if organic compound contains – NO_2 group then NO_2 is liberated]
- If mixture contains N_2 gas & this is exploded with O_2 gas, do not assume any oxide formation unless specified.
- Ozone is absorbed in turpentine oil and oxygen in alkaline pyragallol.

Solved Examples

Ex-17. Carbon dioxide gas (CO_2) measuring 1 litre is passed over heated coke the total volume of the gases coming out becomes 1.6 litre. Find % conversion of CO_2 into carbon monoxide.

Sol. $CO_2 + C \longrightarrow 2CO$ $\quad CO_2 \quad CO$

1 0 at time t $1 - x$ $2x$

Initial volume = 1 litre

final volume = 1.6 litre Final volume = $(1 + x)$ litres

$1 + x = 1.6$

$x = 0.6$

$x = 0.6$

$\frac{0.6}{1} \times 10 = 60\%$ of CO_2 will be converted into CO

Ex-18. 100 ml of hydrocarbon is mixed with excess of oxygen and exploded. On cooling, the mixture was reported to have a contraction of 250 ml. The remaining gas when passed through a solution of aqueous KOH, the mixture shows a further contraction of 300 ml. Find molecular formula of the hydrocarbon.

Sol. $C_xH_y + \left(x + \frac{Y}{4}\right)O_2 \longrightarrow x\,CO_2 + \frac{y}{2}H_2O$

100 ml $\quad x\,100 \quad \frac{y}{2}100.$

mixture contains CO_2 , H_2O & excess O_2

on cooling, H_2O is separated, volume of H_2O = 250 ml

$\frac{y}{2} \times 100 = 250$; $y = 5$,

as KOH absorbs CO_2, hence 300 ml contraction is because of CO_2 that has been absorbed.

Volume of $CO_2 = 100\,x = 300$; $x = 3$

Empirical formula = C_3H_5 ; molecular formula = C_6H_{10}.

Note: If water is already condensed out then total contraction in reaction mixture = {volume of reactants} – {volume of products + volume of unused species excluding volume of H_2O}.

Ex-19. 100 ml of an hydrocarbon is burnt in excess of oxygen in conditions so that water formed gets condensed out the total contraction in volume of reaction mixture was found to be 250 ml when the reaction mixture is further exposed to aqueous KOH a further contraction of 300 ml is observed find molecular formula of hydrocarbon.

Sol. $C_xH_y + O_2 \longrightarrow CO_2 + H_2O$

100 ml excess 300 ml

By POAC on 'C' atoms

$x \times 100 = 300$

$x = 3$

POAC on 'H' atoms

$y \times 100 = 2 \times$ moles of H_2O

POAC on O atoms

$2 \times v = 2 \times 300 + 1 \times H_2O$ {V = volume of O_2 consumed}

$2 \times v = 600 + 50\,y$

$v = \frac{600 + 50y}{2}$ volume of O_2 consumed

The total volume contraction is 250 ml.

Hence, $100 + V - 300 = 250$

$-200 + V = 250$

$2 \times 450 - 600 = 50\,y$

$\frac{300}{50} = y = 6$

Hydro carbon will be C_3H_6

Alternative :

Using balanced chemical equation

$C_xH_y + \left(\frac{x+y}{4}\right)O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O$

t = 0 100 ml V 0 0

t   0 $V - 100\left(x + \frac{y}{4}\right)$ 100 x ml $\frac{100\,y}{2}$

volume remained

$V - 100\left(x + \frac{y}{4}\right) + 100x + 50y$

$-100 - V = 250$
$-25y + 50y = 150$
$25y = 150$
$y = 6$

Summary A gas is a collection of atoms or molecules moving independently through a volume that is largely empty space. Collisions of the randomly moving particles with the walls of their container exert a force per unit area that we perceive as pressure. The SI unit for pressure is the pascal, but the atmosphere and the millimeter of mercury are more commonly used. The physical condition of any gas is defined by four variables; pressure (P), temperature (T), volume (V) and molar amount (n). The specific relationship among these variables are called the gas laws :

Boyle's law :	The volume of a gas varies inversely with its pressure. That is, $V \propto 1/P$ or $PV = k$ at constant n. T.
Charles Law :	The volume of a gas varies directly with its Kelvin temperature. That is, $V \propto T$ or $V/T = k$ at constant n, P
Avogadro's Law :	The volume of a gas varies directly with its molar amount. That is, $V \propto n$ or $V/n = k$ at constant T,P.

The three individual gas laws can be combined into a single ideal gas law, PV = nRT. If any three of the four variables P, V, T and n are known, the fourth can be calculated. The constant R in the equation is called the gas constant and has the same value for all gases. At standard temperature and pressure (STP; 1 atm and 0°C), the standard molar volume of an ideal gas is 22.414 L.

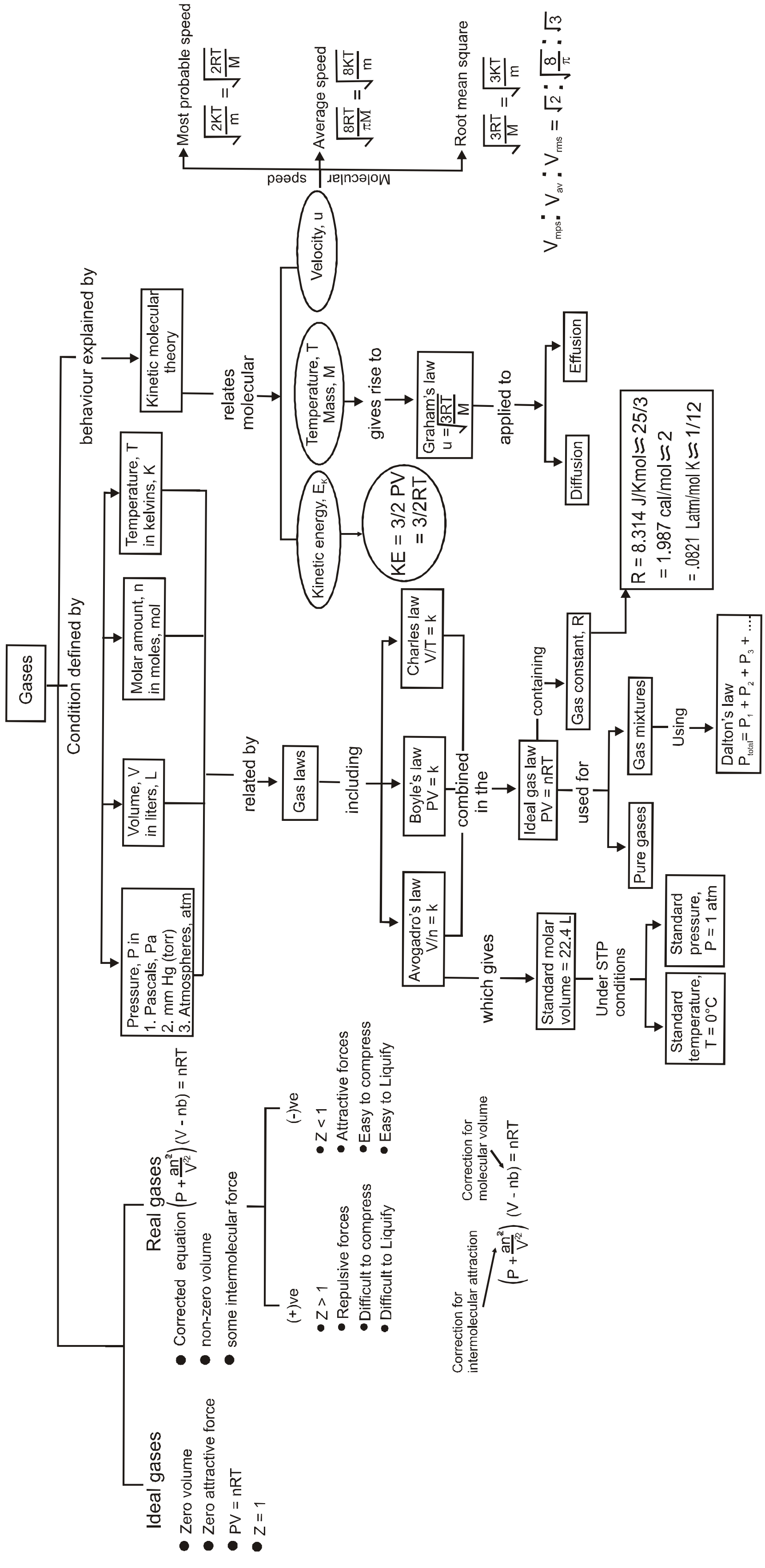
The gas laws apply to the mixture of gases as well as to pure gases According to Dalton's law of partial pressures, the total pressure exerted by a mixture of gases in a container is equal to the sum of the partial pressure of each individual gas would exert alone.

The behaviour of gases can be accounted for using a model called the kinetic-molecular theory, a group of five postulates:

1. A gas consists of tiny particles moving at random.
2. The volume of the gas particles is negligible as compared with the total volume.
3. There are no forces between particles, either attractive or repulsive.
4. Collisions of gas particles are elastic.
5. The average kinetic energy of gas particles is proportional to their absolute temperature.

The connection between temperature and kinetic energy obtained from the kinetic- molecular theory makes it possible to calculate the average speed of a gas particle at any temperature. An important practical consequence of this relationship is Graham' law, which states that the rate of a gas effusion or spontaneous passage through a pinhole in a membrane depends inversely on the square root of the molar mass of gas.

Real gases differ in their behaviour from that predicted by the ideal gas law, particularly at higher pressure, where gas particles are forced close together and intermolecular forces become significant.


Gases
Condition defined by
behaviour explained by
Kinetic molecular theory
relates molecular
Kinetic energy, E_K
Temperature, T Mass, M
Velocity, u
Molecular speed
Most probable speed
$\sqrt{\frac{2KT}{m}} = \sqrt{\frac{2RT}{M}}$
Average speed
$\sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{\pi m}}$
Root mean square
$\sqrt{\frac{3RT}{M}} = \sqrt{\frac{3KT}{m}}$
$V_{mps} : V_{av} : V_{rms} = \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3}$
KE = 3/2 PV = 3/2RT
gives rise to
Graham's law $u = \sqrt{\frac{3RT}{M}}$
applied to
Diffusion
Effusion
Pressure, P in 1. Pascals, Pa 2. mm Hg (torr) 3. Atmospheres, atm
Volume, V in liters, L
Molar amount, n in moles, mol
Temperature, T in kelvins, K
related by
Gas laws
including
Avogadro's law V/n = k
Boyle's law PV = k
Charles law V/T = k
combined in the
Ideal gas law PV = nRT
containing
Gas constant, R
R = 8.314 J/Kmol ≃ 25/3 = 1.987 cal/mol ≃ 2 = .0821 Latm/mol K ≃ 1/12
used for
Pure gases
Gas mixtures
Using
Dalton's law $P_{total} = P_1 + P_2 + P_3 + \ldots$
which gives
Standard molar volume = 22.4 L
Under STP conditions
Standard temperature, T = 0°C
Standard pressure, P = 1 atm
Ideal gases
● Zero volume
● Zero attractive force
● PV = nRT
● Z = 1
Real gases
● Corrected equation $\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$
● non-zero volume
● some intermolecular force
(+)ve
● Z > 1
● Repulsive forces
● Difficult to compress
● Difficult to Liquify
(-)ve
● Z < 1
● Attractive forces
● Easy to compress
● Easy to Liquify
Correction for intermolecular attraction
Correction for molecular volume
$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$

MISCELLANEOUS SOLVED PROBLEMS (MSPs)

1. The diameter of a bubble at the surface of a lake is 4 mm and at the bottom of the lake is 1 mm. If atmospheric pressure is 1 atm and the temperature of the lake water and the atmosphere are equal. What is the depth of the lake ?
(The density of the lake water and mercury are 1 g/ml and 13.6 g/ml respectively. Also neglect the contribution of the pressure due to surface tension)

Sol. $P_1V_1 = P_2V_2$

$\therefore$ (760 mm × 13.6 × g) $\frac{4}{3}\pi$ (4 mm/2)^3 = (760 mm × 13.6 × g + h × 1 × g) $\frac{4}{3}\pi$ (1 mm/2)^3

760 × 13.6 × 64 = (760 × 13.6 + h)

h = 64 × 760 × 13.6 – 760 × 13.6

h = 63 × 760 × 13.6 mm

$h = \frac{63 \times 760 \times 13.6}{1000 \times 1000}$ km = 0.6511 km = **651.1 m Ans.**

2. A gas is initially at 1 atm pressure. To compress it to 1/4 th of initial volume, what will be the pressure required ?

Sol. P_1 = 1 atm V_1 = V

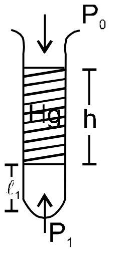
P_2 = ? V_2 = V/4

$P_1V_1 = P_2V_2$ at const. T & n

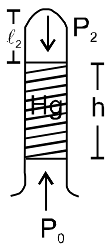
$P_2 = \frac{P_1V_1}{V_2} = \frac{1\text{ atm} \times V}{V/4}$ = **4 atm Ans.**

3. A gas column is trapped between closed end of a tube and a mercury column of length (h) when this tube is placed with its open end upwards the length of gas column is (ℓ_1), the length of gas column becomes (ℓ_2) when open end of tube is held downwards. Find atmospheric pressure in terms of height of Hg column.

case I



case II



Sol. For gas $P_1 = (P_0 + h)$ $P_2 = (P_0 – h)$

$V_1 = \pi r^2\ell_1$ $V_2 = \pi r^2\ell_2$

at const T. and moles.

$P_1V_1 = P_2V_2$; $(P_0 + h)\,\pi r^2\ell_1 = (P_0 – h)\,\pi r^2\ell_2$

$P_0\ell_2 + h\ell_1 = P_0\ell_2 – h\ell_2$

$P_0\ell_2 – P_0\ell_1 = h\ell_1 + h\ell_2$

$P_0 = \frac{h\,(\ell_1 + \ell_2)}{(\ell_2 - \ell_1)}$ cm of Hg column **Ans.**

4. If water is used in place of mercury then what should be minimum length of Barometer tube to measure normal atmospheric pressure.

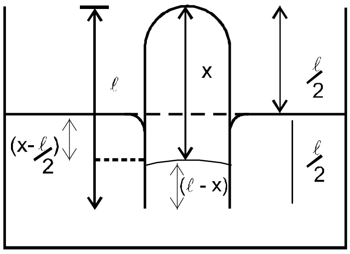
Sol. $P_{H_g} = P_{H_2O} = P_{atm}$.

0.76 m × 13.6 × g = h_{H_2O} × 1 × g;

h_{H_2O} = 0.76 × 13.6 = **10.336 m Ans.**

5.

A tube of length 50 cm is initially in open atmosphere at a pressure 75 cm of Hg. This tube is dipped in a Hg container upto half of its length. Find the level of mercury column in side the tube.



Sol. If after dipping the tube, the length of air column be x cm (situation shown in the adjoining figure).
Then by using, $P_iV_i = P_fV_f$

We have, 75 cm Hg × ℓA = P_f × x × A .... (1) (ℓ = 50 cm)

& also, P_f = 75 cm Hg + (x – $\frac{\ell}{2}$) .... (2)

(2) & (1) $\Rightarrow$ [75 + (x – 25)] × x = 75 × 50

$\Rightarrow$ x^2 + 50 x – 3750 = 0

$\therefore$ x = 41.14 or – 91.14

But, x can't be – ve ∴ $x = 41.14$

∴ Mercury column inside the tube = (50 – 41.14) cm **= 8.86 cm Ans.**

6. An open container of volume V contains air at temperature 27ºC or 300 K.The container is heated to such a temperature so that amount of gas coming out is 2/3 of **(a)** amount of gas initially present in the container. **(b)** amount of gas finally remaining in the container.
Find the temperature to which the container should be heated.

Sol. (a) Here, P & V are constant, n & T are changing. Let, initially the amount of gas present be n & temp is 27ºC or 300K. Finally amount of gas present in container = $n - \frac{2}{3}n = \left(\frac{1}{3}\times n\right)$ & final temperature be T.

Then using $n_1T_1 = n_2T_2$, we have, $n \times 300 = \frac{n}{3} \times T_2 \Rightarrow T_2 = 900K$ i.e., final temp = **900K Ans.**

(b) Let there be x moles of gas remaining in the container, $\frac{2}{3}$ of x come out

∴ $(\frac{2}{3}x + x) = n \Rightarrow \frac{5x}{3} = n \;\therefore\; x = \frac{3n}{5}$

∴ Using $n_1 T_1 = n_2 T_2$ $\quad n \times 300\text{ K} = \frac{3n}{5} \times T_2$

∴ $T_2 = 500K$

Final temperature = **500 K Ans.**

7. Find the lifting power of a 100 litre balloon filled with He at 730 mm and 25°C. (Density of air = 1.25 g /L).

Sol. Since, $PV = nRT$

$PV = \frac{W}{M}RT \quad \therefore \quad W = \frac{PVM}{RT} = \frac{730}{760} \times \frac{100\times 4}{0.082\times 298}$ g

i.e., Wt. of He = 15.72 g

Wt. of air displaced = 100 × 1.25 g/L = 125 g

∴ Lifting power of the balloon = 125 g – 15.72 g = **109.28 g Ans.**

8. A weather balloon filled with hydrogen at 1 atm and 300 K has volume equal to 12000 litres. On ascending it reaches a place where temperature is 250 K and pressure is 0.5 atm. The volume of the balloon is :

(A) 24000 litres (B) 20000 litres (C) 10000 litres (D) 12000 litres

Sol. Using $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$; $\quad \frac{1\text{ atm} \times 12000\text{ L}}{300\text{ K}} = \frac{0.5\text{ atm} \times V_2}{250\text{ K}}$

∴ $V_2 = 20{,}000$ L

Hence, Ans. (B)

9. Four one litre flasks are separately filled with the gases, O_2, F_2, CH_4 and CO_2 under the same conditions. The ratio of number of molecules in these gases :

(A) 2 : 2 : 4 : 3 (B) 1 : 1 : 1 : 1 (C) 1 : 2 : 3 : 4 (D) 2 : 2 : 3 : 4

Sol. According to Avogadro's hypothesis.
All the flasks contains same no. of molecules

∴ Ratio of no. of molecules of O_2, F_2, CH_4 & CO_2
= 1 : 1 : 1 : 1 **Hence, (B)**

10. A sample of water gas has a composition by volume of 50% H_2, 45% CO and 5% CO_2. Calculate the volume in litres at STP at which water gas which on treatment with excess of steam will produce 5 litre of H_2. The equation for the reaction is:

$$CO + H_2O \longrightarrow CO_2 + H_2$$

Sol. If x L CO in needed then volume of H_2 in water gas = $\left(\frac{x}{0.45}\times 50\%\right)$ L = $\left(\frac{x}{0.45}\times\frac{1}{2}\right)$ L = $\frac{x}{0.9}$ L

But, from equation : $CO + H_2O \longrightarrow CO_2 + H_2$
& Gay-Lussac's law, we get, that the volume of H_2
Produced = volume of CO taken.

∴ Volume of H_2 due to reaction = x L

∴ Total volume of H_2 = $\left(\frac{x}{0.9} + x\right)$ L = 5 L

⇒ $\frac{1.9x}{0.9} = 5$ L

∴ $x = \frac{0.9 \times 5}{1.9}$ ∴ Volume of water gas = $\frac{x}{0.45}$ L = $\frac{0.9 \times 5}{1.9 \times 0.45}$ L = **5.263 L Ans.**

11. The partial pressure of hydrogen in a flask containing two grams of hydrogen and 32 gm of sulphur dioxide is :

(A) 1/16th of the total pressure (B) 1/9th of the total pressure

(C) 2/3 of the total pressure (D) 1/8th of the total pressure

Sol. $n_{H_2} = \frac{2g}{2g/mol} = 1$ mol. ; $n_{SO_2} = \frac{32g}{64g/mol} = 0.5$ mol

∴ $P_{H_2} = \frac{n_{H_2}}{(n_{H_2} + n_{SO_2})} \times P_T = \frac{1}{(1+0.5)} \times P_T = \frac{2}{3} P_T$. **Hence, Ans. (C)**

12. Equal volume of two gases which do not react together are enclosed in separate vessels. Their pressures are 10 mm and 400 mm respectively. If the two vessels are joined together, then what will be the pressure of the resulting mixture (temperature remaining constant) :

(A) 120 mm (B) 500 mm (C) 1000 mm (D) 205 mm

Sol. Let, vol of containers be V & temperature be T

$P_1 = 10$mm $P_2 = 400$mm

∴ $n_1 = \frac{P_1V}{RT}$ & $n_2 = \frac{P_2V}{RT}$

∴ $n_1 + n_2 = \frac{(P_1+P_2) \times V}{RT}$

After joining two containers final vol = (V + V) = 2V (for gases)

$P_{final} = \frac{(n_1+n_2)RT}{V_{final}} = \frac{(P_1+P_2) \times V}{RT} \times \frac{RT}{2V} = \frac{(P_1+P_2)}{2} = \frac{(10+400)mm}{2} = 205$ mm.

Hence, Ans. (D)

13. 5 ml of H_2 gas diffuses out in 1 sec from a hole. Find the volume of O_2 that will diffuse out from the same hole under identical conditions in 2 sec.

Sol. Rate of diffusion of $H_2 = \frac{5ml}{1 \ sec} = 5ml/s = r_{H_2}$ (say)

∴ $r_{O_2} = r_{H_2} \times \frac{1}{4} = 5ml/s \times \frac{1}{4}$ ∴ Volume of O_2 diffused in 2.0 seconds = $\frac{5}{4} \times 2$ ml = **2.5 ml Ans.**

14. A vessel contains H_2 & O_2 in the molar ratio of 8 : 1 respectively. This mixture of gases is allowed to diffuse through a hole, find composition of the mixture coming out of the hole.

Sol. Here, $n_{H_2} : n_{O_2} = 8 : 1$ & $\frac{r_{H_2}}{r_{O_2}} = \frac{r_{H_2}}{r_{O_2}}\sqrt{\frac{M_{O_2}}{M_{H_2}}}$

⇒ $\frac{r_{H_2}}{r_{O_2}} = \frac{8}{1} \times \sqrt{\frac{32}{2}} = \frac{32}{1}$ ⇒ $\frac{(\text{no. of moles of } H_2 \text{ coming out})/\Delta t}{(\text{no. of moles of } O_2 \text{ coming out})/\Delta t} = \frac{32}{1}$

Required composition of $H_2 : O_2$ coming out = **32 : 1 Ans.**

15. If for two gases of molecular weights M_A and M_B at temperature T_A and T_B; $T_AM_B = T_BM_A$, then which property has the same magnitude for both the gases.

(A) Density (B) Pressure (C) KE per mol (D) RMS speed

Sol. Given that $T_A M_B = T_B M_A$ ⇒ $\frac{T_A}{M_A} = \frac{T_B}{M_B}$

But, r.m.s. = $\sqrt{\frac{3RT}{M}}$

$r.m.s_A = \sqrt{\frac{3RT_A}{M_A}}$ & $r.m.s_B = \sqrt{\frac{3RT_B}{M_B}}$

$r.m.s_A = r.m.s_B$ **Ans. (D)**

16. It has been considered that during the formation of earth H_2 gas was available at the earth. But due to the excessive heat on the earth this had been escaped. What was the temperature of earth during its formation? (The escape velocity is 1.1×10^6 cm/s)

Sol. Escape velocity of H_2 should be equal to average velocity of H_2.

$\therefore$ Avg velocity of $H_2 = 1.1 \times 10^6$ cm/s $= 1.1 \times 10^4$ m/s

But, avg. velocity = $\sqrt{\frac{8RT}{\pi M}}$ $\Rightarrow 1.1 \times 10^4 = \sqrt{\frac{8 \times 8.314 \times T}{\pi \times 2 \times 10^{-3}}}$ $\left(M_{H_2} = 2g = 2 \times 10^{-3} kg\right)$

$\therefore$ T = $\frac{\left(1.1 \times 10^4\right)^2 \times \pi \times 2 \times 10^{-3}}{8 \times 8.314}$ K = 11430.5 K = **11157.5ºC Ans.**

17. A gaseous mixture containing CO, methane CH_4 & N_2 gas has total volume of 40 ml. This mixture is exploded with excess of oxygen on cooling this mixture a contraction of 30 ml is observed & when this mixture is exposed to aqueous KOH a further contraction of 30 ml is observed. Find the composition of the mixture.

Sol. Let vol of CO be x mL, vol of CH_4 be y mL, vol of N_2 be z mL.

On explosion with excess of oxygen the following reactions takes place

$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$ (By Gaylussac's law of combing volume)

x mL x mL

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$

y mL y mL 2 ymL

N_2 remains unreacted

On cooling H_2O (g) liquifies hence volume reduction of 30 mL is observed

2y = 30 y = 15

But, vol of CO_2 obtained = (x + y) mL

This is absorbed in KOH & vol reduction of 30 mL is observed.

x + y = 30 $\Rightarrow$ x = 30 – y = (30 – 15) = 15

and, x + y + z = 40 $\Rightarrow$ z = 40 – x – y = 40 – 15 – 15 = 10

Composition of mixture is

Vol. of CO = **15 mL,** vol of CH_4 = **15 mL,** vol of N_2 = **10 mL** **Ans.**

CHECK LIST

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Formulae (F)

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F-10. $N_A \left(\frac{1}{2}m \quad \overline{U^2}\right) = \frac{3}{2} K N_A T = \frac{3}{2} RT$ ☐
F-11. $U_{rms} = \sqrt{\frac{3RT}{M}}$ ☐
F-12. $U_{av} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{\pi m}}$ ☐
F-13. $U_{MPS} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2KT}{m}}$ ☐

Table (Tab.)

Tab-1. Units of V, P & T ☐

Derivation (Deri.)

Deri-1. Derivation of Kinetic Gas Equation ☐