# CHAPTER 4
# GASES

* Diffusion Order:
  Gases > Liquids > Solids

* Density Order:
  Solid > Liquid > Gas
  → Density of gas is about 1000 times less than that of same mass of liquid or solid.

* Collision of gases → Perfectly elastic

* Ideal gases cannot be liquefied → According to kinetic molecular theory

* The pressure of ideal gas is greater than pressure of real gas bcz in real gas forces of attraction are present.

MCQ: At same temperature, which gas have greater average K.E

a. Helium
b. $N_2$
c. $CO_2$
✓d. All same

PRESSURE

Collision per unit area per second

ATMOSPHERIC PRESSURE

Pressure measured in barometer at sea level, 0°C (273 K)

Cross-section of barometer → $1 cm^3$

Values:

→ 1 atm
→ 76 cm Hg
→ 760 mm Hg
→ 760 torr
SI → 101325 Pa / $101325 Nm^{-2}$
→ 14.7 psi (pound square inch)

Why Mercury Is Used In Barometer?

1. It doesnot evaporate
2. The only force acting on mercury is its weight

[Mercury is used in thermometer due to thermal expansion]

rcury has:

→ Very high surface tension
→ Very low viscosity
→ Low volatility

# MOTION OF PARTICLES OF A GAS

In gases, the molecular motion is of three types:

1. Translational Motion
2. Rotational Motion
3. Vibrational Motion

## 1. TRANSLATIONAL MOTION

The motion imparted to the gaseous molecules due to their motion in all possible directions is called translational motion. In this case the entire molecules move from place to place.

## 2. ROTATIONAL MOTION

The motion imparted to the gaseous molecules as a result of net angular momentum about their centre of gravity. In this case the molecule spins like a propeller.

## 3. VIBRATIONAL MOTION.

The motion imparted to the gaseous molecules due to oscillations is called vibrational motion. In this case the molecules vibrate back and forth about the same fixed location.

A monoatomic molecule (e.g He) will show only translational motion while a diatomic ($H_2$) and polyatomic molecules ($CO_2$, $NH_3$ etc) will undergo, in addition to translational motion, the rotational and vibrational motion too

# Boyle's Law

## STATEMENT

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

## MATHEMATICALLY

$$V \propto \frac{1}{P} \quad \text{(At constant Temperature)}$$

$$PV = K$$

$$P_1 V_1 = P_2 V_2$$

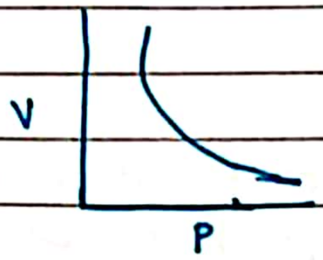
## UNIT OF k FOR BOYLE'S LAW

1. atm $dm^3$
2. Nm
3. Joule

The value of k will remain the same for the same quantity of a gas at the same temperature.

## GRAPH OF P VS. V

→ Hyperbola (due to pressure)
→ Isotherm (due to temperature)




# CHARLE'S LAW

## STATEMENT

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

## MATHEMATICALLY

$V \propto T$ (At constant Pressure and mass)

$V = kT$

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

* For a number of gases, for every degree rise in temperature, the volume of a gas increase by 1/273 of its original volume at 0°C at constant pressuro.

* New Volume at $t°C$ = Original Volume at $0°C$ + $\left[\frac{t}{273}\left(\text{Original Volume at } 0°C\right)\right]$

MCQ: If we have $5 dm^3$ gas at $100°C$. At what T; $V = 10 dm^3$

✓a. 293°C

b. 20°C

↳ This is confusing but wrong. If temperature is given in °C, 273 must be added.

MCQ #2: $V_1 = 5dm^3$ $t_1 = 10°C$ $t_2 = 20°C$

$V_2 = ?$

a) $5 dm^3$ (Not possible bcz same volume)

b) $4dm^3$ (can't reduce)

c) $10dm^3$ (can't double bcz given in °C)

✓d) $5.8 dm^3$ (so this is correct)

# ABSOLUTE TEMPERATURE

Any temperature in Kelvin scale

MCQ: Which one is absolute temperature?

✓ a) Zero Kelvin

b) -273°C

c) Both

# ABSOLUTE ZERO

→ Absolute zero may be taken ~~in~~ at any scale

→ Absolute zero is the lowest hypothetical temperature at which volume of gases should become zero.

$0K = -273°C = -459°F$

# IDEAL GAS EQUATION

$PV = nRT$

R: Universal Constant or General Gas Constant

## AT STP:

Standard Temperature (T) = 273 K

Pressure (P) = 1 atm

Amount (n) = 1 mol

Volume (V) = 22.4 $dm^3$

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

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

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

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

## SI VALUE OF R:

$R = 8.3143\ Nm\ mol^{-1} K^{-1}$

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

* $PV = nRT$
* $PM = dRT$
* $mRT = PVM$

M: Molar Volume

d: density

# CONVERSION FACTORS

$1 \, m^3 = 1000 \, dm^3$

$1 \, dm^3 = 1000 \, cm^3$

$1 \, dm^3 = 0.001 \, m^3$

$1 \, cm^3 = 0.001 \, dm^3$

$1 \, cm^3 = 10^{-5} \, m^3$

$1 \, Nm = 1 \, J$

1 calorie = 4.18 J

1 J = 0.239 cal

$1 \, J = 10^7 \, erg$

# AVOGADRO'S LAW

Volume of a gas is directly proportional to number of moles at STP.

$V \propto n$ (At constant T, P)

Equal volumes of ideal gases at the same temperature and pressure contain equal number of molecules.

$22.4\,dm^3$ of gas at STP = $6.023 \times 10^{23}$ molecules

$1\,dm^3$ of gas at STP = $\frac{6.023 \times 10^{23}}{22.4} = 2.68 \times 10^{23}$ molecules

* One molecule is approximately at a distance of 300 times its own diameter from its immediate neighbour at room temperature.

# IDEAL GASES

* Gases which obeys ideal gas law i.e $PV = nRT$ are ideal gases.

* $\frac{PV}{nRT} = 1 = Z$ → For Ideal Gases

  Z : Compressibility Factor

* Z for real gases may be greater than 1 or equal to 1.

* Z is always greater than 1 for He and $H_2$. So $H_2$ and He always show +ve deviation.

  All other gases show both type of deviation (+ve as well as -ve)

* Extent of Deviation depends on:
  1. Temperature
  2. Pressure
  3. Nature of gas

* Gases show maximum deviation at:

  High Pressure and low temperature

* Gases show ideal behaviour at:

  High Temperature and low pressure

* Helium gas behaves most like ideal gas.

* Those gases will behave more ideally having weakest attractive force.

* Real gases deviate from ideal behaviour due to the following reasons:
  1. Attractive forces exist in real gases.
  2. The volume of real gas molecule is not negligible at high pressure. and low temperature

* Deviation Order:

  $CO_2 > N_2 > H_2 > He$

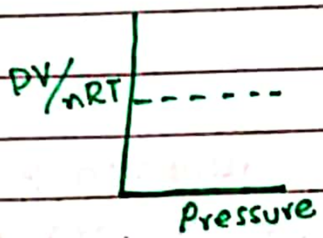
* Greater the boiling point, greater is the deviation

* Helium converts to liquid at 4 K.
* Hydrogen ($H_2$) converts to liquid at 20K

* An ideal gas cannot be liquified bcz forces operative between its molecules are negligible.

# NON- IDEAL BEHAVIOUR OF GASES

* If a graph is plotted between pressure on x-axis and PV/nRT (compressibility factor or Z) on y-axis for an ideal gas, a straight line parallel to the pressure axis is obtained.




* For real gases (non-Ideal gases), the graph is no more parallel to the pressure axis.
  At high pressure and low temperature, the graph for real gases comes closer to the expected straight line.

* If $Z = 1$, the line would be parallel to x-axis
* If $Z < 1$, then the line obtained will be below the line of ideal gas which means that there is larger decrease in volume of the gas than predicted by general gas equation due to attractive forces present among molecules.
* If $Z > 1$ then the line obtained will be above the line of an ideal gas which means that there is less decrease in volume of gas than predicted by general gas equation due to repulsive forces among molecules.

# VAN DER WAAL'S EQUATION

## PRESSURE CORRECTION:

$$P = P_i - P'$$

$P'$ : Pressure used up against intermolecular attraction

$P_i = P + P'$

$$P_i = P + \frac{a}{V^2}$$

a: co-efficient of attraction i-e attraction per unit volume and is constant for a particular real ga

Unit of a: atm $dm^6 mol^{-2}$

## VOLUME CORRECTION:

$$V_{free} = V - V_{molecules}$$

Unit of b: $dm^3 mol^{-1}$

$$V_{free} = V - b$$

where $b = 4 V_m$ (Roughly Estimated)

## VANDER WAAL'S EQUATION:

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

For 'n' moles

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

# DALTON'S LAW OF PARTIAL PRESSURES

$$P_{TOTAL} = P_1 + P_2 + P_3 \cdots$$

Dalton's law is not applicable for reactive gases.

1. PARTIAL PRESSURE FROM PERCENTAGE

Suppose Gas → A

$$P_A = \frac{\% \text{ of A (in mixture)}}{100} \times P_{TOTAL}$$

2. PARTIAL PRESSURE FROM MOLES

$$P_A = \frac{n_A}{n_{TOTAL}} \times P_{TOTAL}$$

3. PARTIAL PRESSURE FROM MOLE FRACTION

$$P_A = X_A P_{TOTAL}$$

where

$$X_A = \frac{n_A}{P_{TOTAL}}$$

# DIFFUSION AND EFFUSION

* Diffusion:
  → spontaneous
  → random
  → occur with collision

* Effusion:
  → spontaneous
  → non-random
  → no collision, escape of a gas

* Rate of Diffusion depends on:
  1. Temperature
  2. Molar mass
  3. Density
  4. Kinetic Energy

# GRAHAM'S LAW OF DIFFUSION

* Rate of Diffusion = $\frac{\text{Distance Travelled}}{\text{Time}}$

→ Relative Rate:

$$\frac{r_A}{r_B} = \frac{\text{Distance Travelled by A}}{\text{Distance Travelled by B}}$$

$$\frac{r_A}{r_B} = \sqrt{\frac{M_2}{M_1}}$$ M: Molecular mass

$$\frac{r_A}{r_B} = \sqrt{\frac{d_2}{d_1}}$$ d: Density

* Kinetic Molecular Theory predicts that the average speed depends on the molecular mass as:

$$v \propto \sqrt{\frac{1}{m}}$$

# LIQUEFACTION OF GAS

* The conversion of a gas to its liquid state by the combined efforts of lowered temperature and increased pressure is called liquefaction of gas.

→ The increased pressure is obtained by compressor.

→ It has been found that it is impossible to liquefy a gas by pressure alone if the required temperature is not obtained.

## Critical Temperature: ($T_c$)

The highest temperature at which a substance can exist as a liquid, is called its critical temperature ($T_c$)

OR

The temperature above which a gas cannot be liquified by pressure alone is called critical temperature.

## CRITICAL PRESSURE ($P_c$):

It is the pressure which is required to bring about liquefaction of a gas at its critical temperature.

## CRITICAL VOLUME ($V_c$):

The volume occupied by one mole of a gas at $P_c$ and $T_c$ is called critical volume.

# * FACTORS AFFECTING CRITICAL TEMPERATURE

1. Size of molecule
2. Shape of molecule
3. Intermolecular forces among molecules

→ The higher is the intermolecular force of attraction higher will be critical temperature and easier is the liquefaction of the gas.

<u>MCQ</u>: The critical temperature of $N_2$ is 126K and He is 5.3K. Which of the gases will liquefy first?

<u>Ans</u>: $N_2$ liquifies easily.

Intermolecular forces are high for gases which have a high critical temperature, making the change of state first.

* FOR POLAR GASES

$T_c, V_c, P_c \propto$ Intermolecular Forces

* FOR NON POLAR GASES

$T_c, V_c, P_c \propto$ Molar mass of gas

<u>MCQ</u>: Gas having highest $T_c$

a) $H_2$ b) He c) $O_2$ ✓d) $Cl_2$

# JOULE THOMSON EFFECT

* Sudden Expansion of a gas into a region of low pressure causes cooling.

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

* This effect is the basis of Linde's method of liquefaction:

→ Compression of Gas causes heating effect.
→ Expansion of gas causes cooling effect (except $H_2$ and He)
$H_2$ and He will cool upon expansion only if their initial temperatures are very low because the long-range forces in these gases are unusually weak.

NEGATIVE JOULE THOMSON EFFECT:
Instead of cooling, it causes heating on expansion.
$H_2$, He, Ne show negative Joule Thomson effect

# LINDE'S METHOD

→ Linde liquified air by this process

→ Pressure : 200atm

# PLASMA : FOURTH STATE OF MATTER

* Ionized gas mixture which consists of ions, electrons, and neutral atoms is called plasma".

→ Ionized gas cannot exist at room temperature.

→ This state of matter consist of an ionized substance but neutral as whole.

→ The free electric charges make the plasma electrically conductive so that it responds strongly to electromagnetic fields.

→ About 99% of universe is made up of plasma.

→ The sun is a 1.5 million balls of plasma. It is heated by nuclear fusion.

→ Plasma can respond to both electric and magnetic fields.

→ Natural plasma exist only at very high temperature or low temperature vacuums.