States Of Matter

A substance is said to be solid if its melting point is above room temperature under atmospheric pressure.
A liquid - If its melting point is below room temperature and boiling point is above room temperature and a gas if its boiling point is below room temperature under atmospheric pressure.
The temperature at which all the three phase exist together is called triple point.

① Solid is defined as that state of matter which possesses a definite shape and a definate volume.
② Liquid is defined as that state of matter which has a definite volume but no-definite shape, they take up the shape of the vessel in which they kept.
③ Gas - is defined as that state of matter which has neither definite shape nor definite volume.
④ Plasma state - which consists of a mixture of electrons and positively charged ions formed due to superheating of the gaseous state.
⑤ Fifth state. Consists of a supercooled solid in which the atoms lose their individual identity and condense to form a super atom.

— Intermolecular force - The force of attraction existing among the molecules of a substance are called intermolecular force. These forces are responsible for the structural feature and Physical Properties of the substance.
— Intramolecular force - The force of attraction existing among the same molecules or a polyatomic ion. It affect the chemical properties of the substance.
Types of intermolecular force.
a Dipole - Dipole interaction - These force of attraction occur among the polar molecules. The positive pole of one molecule

is attracted by the negative Pole of the other molecule. these forces are also called Keesom forces and the effect is called orientation effect.

Dipole-Dipole interaction energy between stationary Polar molecules (as present in solid) is Proportional to $1/r^3$ and between the rotating molecules is Proportional to $1/r^6$, where r is the distance between the Polar molecules.

example - HCl, PH_3, H_2S

(b) Ion-dipole interaction - this is the attraction between an ion and a Polar molecule.

The strength of this interaction depends upon the charge and size of the ion and the magnitude of dipole moment and size of the Polar molecule

(c) Ion-induced dipole interaction - A non Polar molecule may be Polarized by the Presence of an ion near it. It becomes an induced dipole. The interaction between them are called ion-induced dipole interaction.

The strength of these interactions depends upon the charge on the ion and the ease with which the non-Polar molecule gets Polarized.

(d) Dipole-induced dipole interaction - A non Polar molecule may be Polarized by the Presence of a Polar molecule near it. there by making it an induced dipole. The interaction between them are then called dipole-induced dipole interaction.

The strength will depend upon the strength of the dipole and the ease of Polarisability of the non-Polar molecule

(e) London forces or dispersion forces — At any instant of time the electron cloud of the molecule may be distorted so that an instantaneous dipole or momentary dipole is produced in which one part of the molecule is slightly more negative then the rest. The momentary dipole, induce dipole in the neighbouring molecules, these are than attracted to each other. The force of attraction between the induced momentary dipoles are called London-dispersion forces.

In case of a non Polar molecules substance, London forces are the only intermolecular forces that exist.
ex. Cl_2, CCl_4, SiH_4

— Magnitude of the London forces depend on.

i London forces increases with increasing of molecular mass.

ii Geometry of the molecule — London forces are greater in long chain than a branch chain.

n-Pentane > neo-Pentane.

— Whether a substance will exist as a solid or a liquid or a gas is depends upon.

i intermolecular force ii thermal energy

— melting Point — It is the temperature at which a solid melts to form a liquid.

— Boiling Point — It is the temperature at which a liquid change into its vapour.

— The instrument used for the measurment of atmospheric Pressure is called barometer.

— The instrument used for the measurment of the Pressure of a gas is called mano-meter.

— A standerd or normal atmospheric Pressure is defined as the Pressure exerted by a mercury column of exactly 76 cm at 0°C, this the Pressure exerted by the atmosphere at sealevel.

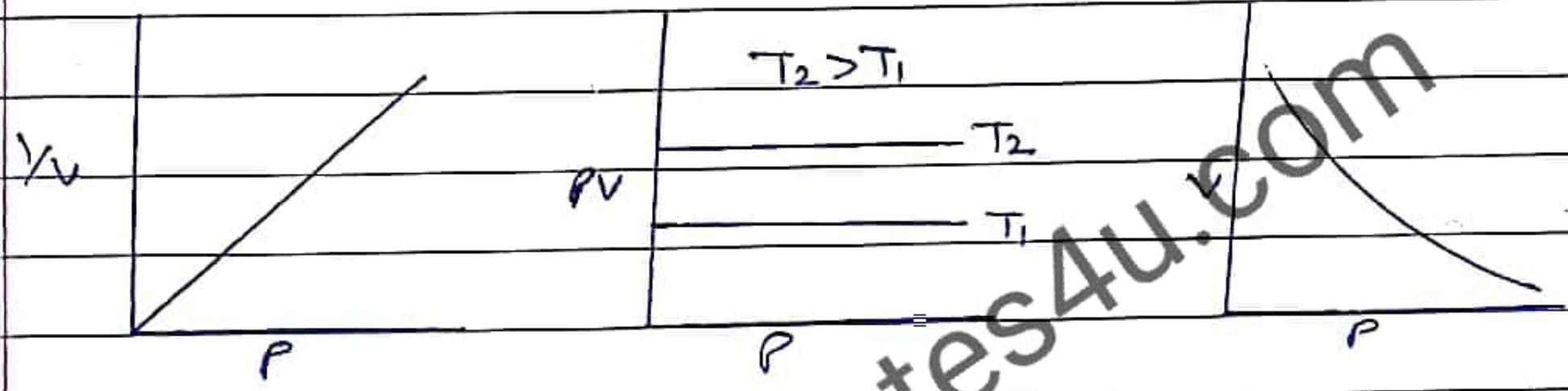
Gas laws.

1. Boyle's law – At a constant temperature the volume of a given mass of a gas is inversely proportional to its pressure.

$$V \propto \frac{1}{P} \quad \text{or} \quad V = \frac{K}{P} \quad \text{then} \quad PV = K.$$

Where K is the constant and its value depends upon the mass of the gas and temperature.

Graphical verification of Boyle's law



If 'm' is the mass of the gas and V its volume then density

$$d = \frac{m}{V}$$

By Boyle's law $V = K/P$.

$$\text{then } d = \frac{m}{K} P \quad \text{or} \quad d \propto P.$$

So At Constant temperature density is Proportional to Pressure.

2. CHARLE'S Law – At Constant Pressure, the Volume of a given mass of a gas is directly Proportional to its temperature in degree Kelvin. $V \propto T$.

$$\text{or } \frac{V}{T} = \text{Constant at Constant Pressure} = K.$$

the numerical value of K depends upon the amount of the gas taken and the Pressure

3. Gay-Lussac's Law / Amonton's law – At Constant volume, the Pressure of a given mass of a gas is directly Proportional to its temperature in degree Kelvin.

$$P \propto T \quad \text{or} \quad P = KT.$$

The numerical value of K is depends upon the amount of

the gas taken and the volume.

4. Avogadro's law — Equal volume of all gases under the same condition of temperature and pressure contain equal number of molecules.

As 1 mol of a gas contains 6.022×10^{23} molecules, this means that one mole of each gas at the same temperature and pressure should have the same volume.

- Molecular volume of a gas under different conditions.

a. Standerd temperature and Pressure condition (STP) — At 0°C or 273.15 K and one atmospheric Pressure, under these condition 1 mole of a gas occupies a volume of 22.4 Liter.

However if the standerd conditions used are 0°C and 1 bar Pressure then molar volume is higher and equal to 22.7 Lit.

b. Standerd ambient temperature and Pressure condition (SATP) — Now used in some scientific works 298.15K and 1 bar Pressure at sea level. Under these conditions the molar volume of an ideal gas is 24.789 Liter

According to Avogardo's law. n = number of moles.

$V \propto n$ or $V = kn$; m - Given mass.

But $n = \frac{m}{M}$; M - Molar mass

$\therefore V = \frac{km}{M}$ or $M = k\frac{m}{V}$ or $M = kd$.

So we can say that density of a gas is directly proportional to its molar mass.

— Ideal gas equation — The equation which gives the simultaneous effect of Pressure and temperature on the volume of a gas is known as ideal gas equation or equation of state for an ideal gas.

Method. I.

Suppose that the volume of a given mass of a gas changes from V_1 to V, when the pressure is changed from P_1 to P_2 at Constant temperature T_1

then according to Boyle's law.

$$P_2 \times V = P_1 \times V_1$$

$$V = \frac{P_1 \times V_1}{P_2} \quad \text{—(1)}$$

Now the volume V change to V_2 when the temperature is changed from T_1 to T_2 at. Constant Pressure P_2

$$\therefore \frac{V}{T_1} = \frac{V_2}{T_2}. \quad \text{or} \quad V = \frac{T_1}{T_2} \times V_2 \quad \text{—(2)}.$$

By (1) and (2). $\frac{P_1 \times V_1}{P_2} = \frac{T_1 \times V_2}{T_2}.$

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

or $\frac{PV}{T} = \text{Constant.} = K. \quad \text{—(3)}$

The value of the K depend upon the amount of the gas taken. If n is the number of moles of the gas taken then

$$K \propto n \quad \text{or} \quad K = nR. \quad \text{—(4)}$$

Where R is the Constant Called universal gas Constant and It depends on the amount of gas taken.

By (3) and (4).

$$\frac{PV}{T} = nR \quad \text{or} \quad PV = nRT.$$

Method II.

According to Boyle's law $V \propto \frac{1}{P}$ at Constant T. —(1)

According to charles law $V \propto T$ at Constant P. —(2)

According to Avogardo's law $V \propto n$ at Constant T and P. —(3).

Where n is the nomber of moles of the gas.

on Combining ①, ② and ③.

$$V \propto \frac{1}{P} \times T \times n.$$

$$V = R \times \frac{1}{P} \times T \times n.$$

$$PV = nRT.$$

A gas that obeys the ideal gas equation exactly is called ideal gas.

— Relationship between molar mass and density of a gas —

If m is the mass of the gas in grams and M is the molecular mass of the gas then.

$$n = \frac{m}{M}$$

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

$$P = \frac{m}{V} \times \frac{RT}{M} \quad \text{or} \quad M = d\frac{RT}{P}.$$

where d is the density of the gas.

— Dalton's law of Partial Pressure — If two or more gases which do not react chemically are enclosed in a vessel, the total Pressure exerted by the gaseous mixture is equal to the sum of all the Partial Pressures that each gas would exert when Present alone in the same vessel at the same temperature

Let $P_1, P_2 \ldots P_n$ are the Partial Pressure of n gaseous enclosed in a given vessel and P is the total Pressure of the gaseous mixture then By dalton law

$$P = P_1 + P_2 + \cdots + P_n.$$

— Applications of Dalton's law.

a. In the determination of Pressure of a dry gas —

Whenever a gas is Collected over water, it is moist, it is Saturated with water vapours which exerts

their own Pressure. The Pressure due to water vapours is Called aqueous tension.

If P and P' are the Pressure of the dry gas and the moist gas respectively at $t°C$ and p is the aqueous tension then. $P = P' - p$

$P_{dry\ gas} = P_{moist\ gas} - p$ Aqueous tension.

b. In the Calculation of Partial Pressure — In a mixture of non-reacting gases A, B, C etc.. If each gas is Considered to be an ideal gas, then. By $PV = nRT$.

$$P_A = n_A \frac{RT}{V} \; ; \; P_B = n_B \frac{RT}{V} \; ; \; P_C = n_C \frac{RT}{V}$$

then according to dalton's law of Partial Pressure.

$$P = P_A + P_B + P_C + \cdots$$

$$= \frac{RT}{V}(n_A + n_B + n_C + \cdots$$

$$\therefore \frac{P_A}{P} = \frac{n_A}{n_A + n_B + n_C + \cdots} = x_A \quad \text{(mole fraction)}$$

then $P_A = P \times x_A$.

Partial Pressure = Total Pressure x mole fraction.

— Diffusion — The spreading of the molecules of a gas through out the available space is Called diffusion.

— Effusion — A Process in which a gas under Pressure escape out of a fine hole or orifice in a vessel is called Effusion.

— Graham's law of diffusion/effusion — Under similar Condition of temperature and Pressure, the rate of diffusion/effusion of different gases are inversely Proportional to the square root of their densities.

For two gases having densities d_1 and d_2 and rate of.

diffusion/effusion r_1 and r_2 under similar conditions of temperature and pressure.

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

Rate of effusion/diffusion = $\frac{\text{Volume of gas effused / Diffused}}{\text{Time taken}}$.

molecular mass = 2x vapour density.

then $\frac{r_1}{r_2} = \sqrt{\frac{m_2}{m_1}}$

- If two gases taken at different Pressure then as greater Pressure greater is the number of molecules hitting per unit area, greater is the rate of diffusion

$$\frac{r_1}{r_2} = \frac{p_1}{p_2}\sqrt{\frac{d_2}{d_1}} = \frac{p_1}{p_2}\sqrt{\frac{m_2}{m_1}}$$

- Two gases undergoing diffusion at the same Pressure but at two different temperatures

$$\frac{r_1}{r_2} = \sqrt{\frac{T_1 d_2}{T_2 d_1}} = \sqrt{\frac{T_1 m_2}{T_2 m_1}}$$

- Volumes of two gases effused/diffused in the same condition in the same time are

$$\frac{r_1}{r_2} = \frac{v_1}{v_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{m_2}{m_1}}$$

- time taken for effusion/diffusion of same volume of two different gases under the same condition of temperature and Pressure are $\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{m_2}{m_1}}$

- Importance of Graham's law.

1. It helps in the seperation of gases having different density
2. It help in the seperation of Isotopes of certain elements.
3. It help to determine the density or molecular mass of an unknown gas by comparing its rate of diffusion with a known gas.

— Volume Coefficient (α_v) — At Constant temperature, the increase in volume of a gas per degree rise of temperature per cc of the gas at 0°c is called the volume Coefficient of the gas.

$$\alpha_v = \frac{V_t - V_0}{V_0 \times t}$$

V_0 = Volume at 0°c.
V_t = Volume at t°c.

For all gases value of α_v is 1/273

— Pressure Coefficient (α_p) — At Constent volume, the increase in the Pressure of a gas per degree rise of temperature divided by its Pressure at 0°c is Called Pressure Coefficient.

$$\alpha_p = \frac{P_t - P_0}{P_0 \times t}$$

P_0 = Pressure at 0°c.
P_t = Pressure at t°c.

For all gases value of α_p is 1/273

— Gases having the same moleculermass have same rate of diffusion/effusion.

— Atmolysis — The Process of seperation of two gases on the basis of their different rate of diffusion due to difference in their densities is Called atmolysis.

— Postulates of Kinetic theory of gases.

(a) Every gas is made up of a large no. of extremely small Particles Called molecules. All the molecules of a Particular gas are identical in mass and size.

(b) The molecules of a gas are separated from each other by large distance so that the actual volume of the molecules is negligible as Compared to the total volume of the gas.

(c) The distance of separation between the molecules are so large that the force of attraction or repulsion between them are negligible.

(d) The force of gravitation on the molecules is also supposed to be negligible.

(e) The molecules are supposed to be moving continuously in different directions with different velocities. Hence they keep on colliding with one another (called molecular collision) as well as on the walls of the containing vessel.

(f) The Pressure exerted on the walls of the containing vessel is due to the bombardment of the molecules on the walls of the containing vessel.

(g) The molecules are supposed to be perfectly elastic hard sphere so that no energy is wasted when the molecules collide with one another or with the walls of the vessel. The energy may be transferred from some molecules to the other on collision.

(h) Since the molecules are moving with different velocities, they possess different kinetic energy. However the average kinetic energy of the molecules of a gas is directly proportional to the absolute temperature of the gas.

— Kinetic gas equation — on the basis of the various assumptions made in the kinetic theory of gas, a mathematical equation has been derived, this equation is known as kinetic gas equation.

$$PV = \frac{1}{3} mn u^2$$

where P = Pressure exerted by gas.
V = Volume of gas.
m = mass of each molecule
n = total no. of molecules Present
u = Root mean Square speed.
M - molar mass in gram.
d = density of gas.

For 1 mole $mxn = M$.

then. $PV = \frac{1}{3} M u^2$.

$$u = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3P}{d}}$$

— Types of speed.

I Most probable speed (α) — this is the speed possessed by the maximum fraction of molecules.

II) Root mean square speed (u) — It is the square root of the mean of the square of the speeds of the molecules. If $v_1, v_2, \cdots v_n$ are the speeds of n molecules then.

$$u = \sqrt{\frac{v_1^2 + v_2^2 + \cdots + v_n^2}{n}}$$

III) Average speed (v) — It is the average of the different speeds of all the molecules.

$$\text{Average speed } (v) = \frac{v_1 + v_2 + \cdots + v_n}{n}$$

— Relationship between u, α and v.

we know $u = \sqrt{\frac{3RT}{M}}$ $\alpha = \sqrt{\frac{2RT}{M}}$ and $v = \sqrt{\frac{8RT}{\pi M}}$

then $u : \alpha : v = \sqrt{\frac{3RT}{M}} : \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{\pi M}}$

$= \sqrt{3} : \sqrt{2} : \sqrt{8/\pi} = 1.224 : 1 : 1.128$

— Deduction of Boyle law from kinetic gas equation —

According to kinetic gas equation.

$$PV = \frac{1}{3} mn u^2 = \frac{2}{3} \cdot \frac{1}{2} M u^2 \quad \because M = mn.$$

But $K.E. = \frac{1}{2} M u^2$ then $PV = \frac{2}{3} K.E.$

And According to the kinetic theory $K.E. \propto T.$ or $K.E = KT.$ where K is a Constant

then $PV = \frac{2}{3} KT.$ or $PV \propto T.$ If T is kept.

Constant than $PV =$ Constant. which is Boyle law

— Deduction of charles law by kinetic gas equation.

$$PV = \frac{2}{3} KT. \quad \text{or} \quad \frac{V}{T} = \frac{2}{3}\frac{K}{P}$$

then if P is constant than $\frac{V}{T}$ = Constant. which is charles law.

— Deduction of dalton's law of Partial Pressure by K.E.G.E.

let us Consider only two gases.

$$\text{then } PV = \frac{1}{3} mn u^2 \text{ or } P = \frac{1}{3}\frac{mn}{V} u^2.$$

for first gas $P_1 = \frac{1}{3}\frac{m_1 n_1 u_1^2}{V}$

for second gas $P_2 = \frac{1}{3}\frac{m_2 n_2 u_2^2}{V}$.

If both the gases are enclosed together in the same vessel. then total Pressure

$$P = \frac{1}{3}\frac{m_1 n_1 u_1^2}{V} + \frac{1}{3}\frac{m_2 n_2 u_2^2}{V}$$

$$P = P_1 + P_2$$

If more than two gases are Present then.

$$P = P_1 + P_2 + P_3 + \cdots\cdots\cdots$$

— Relation between Average Kinetic energy and Absolute temperature — Deduction from Kinetic theory.

According to kinetic gas equation.

$$PV = \frac{1}{3} mn u^2. = \frac{1}{2}.\frac{2}{3} mn u^2.$$

$$= \frac{1}{2}.\frac{2}{3} M u^2 \qquad mxn = M.$$

molecules mass.

$$= \frac{2}{3} K.E.$$

$$\text{or } K.E. = \frac{3}{2} PV. = \frac{3}{2} RT.$$

To get average kinetic energy per molecule, we divide both side by Avogadro number N.

Average K.E. $= \frac{3}{2}\frac{R}{N} \times T. = \frac{3}{2} kT.$

where $k = \frac{R}{N}$ is called Boltzman Constant.

then Average K.E. $\propto$ Absolute temperature.

The motion of a gas molecules due to temperature is called thermal motion.

— Ideal gas — A gas which obeys the ideal gas equation $PV = nRT$, under all conditions of temperature and pressure is called an ideal gas.

There is no gas which obeys the ideal gas equation under all condition of temperature and pressure. Hence the concept of ideal gas is only theoretical or hypothetical.

— Real gases — The gases are found to obey the gas law fairly well if the pressure is low or the temperature is heigh, such gases are known as real gases. All gases are real gases

— Compressibility factor. — The extent to which a real gas deviates from ideal behaviour can be conveniently studied in term of a quantity Z called the compressibility factor. or

Compressibility factor is the ratio of the actual molar volume of the gas to the calculated molar volume at the same temperature and pressure.

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

For an ideal gas as $PV = nRT$, $Z = 1$
For a real gas as $PV \neq nRT$. $Z \neq 1$

Case I when $Z < 1$

the gas is more compressible than expected from ideal behaviour.

Case II when $Z > 1$

the gas is less compressible than expected from ideal behaviour.

- The temperature at which a real gas behaves like an ideal gas over an appreciable pressure range is called Boyle temperature or Boyle Point.

- Causes of deviation from ideal gas.

By the kinetic theory of gases.

(I) The volume occupied by the gas molecules is negligble as compared to the total volume of the gas.

(II) The force of attraction or repulsion between the gas molecules are negligible.

these two assumptions are true only if the pressure is low or the temperature is high so that the distance between the molecules is large.

If the pressure is heigh or the temperature is low, the gas molecules comes close together, hence under these condition.

(I) The forces of attraction or repulsion between the molecule may not be negligible.

(II) The volume occupied by the gas may be so small that the volume occupied by the molecules may not be negligible

- Correction for volume - Supposed the volume occupied.

by the gas volume is V, when the molecules are moving their effective volume is 4 times the actual volume, Let us Call $b = 4v$ this volume is called Co-volume or excluded volume. then Correct volume for n mole $= V - nb$.
(because the free volume for movement of gas molecules is 3v)

(II) Correction of Pressure - A molecule lying within the vessel is attracted by other molecules on all sides but a molecule near the wall is attracted by the molecules inside, Hence it exerts less Pressure then observed Pressure is less than the ideal Pressure.

Corrected Pressure $= P + p$

the Correction term $p \propto d^2$ and $d \propto \frac{1}{V}$ for one mole.

$p \propto \frac{1}{V^2}$ or $p \propto \frac{a}{V^2}$ for one mole.

$p \propto \frac{an^2}{V^2}$ for n mole,

then Corrected Pressure $= P + \frac{an^2}{V^2}$.

where a is a Constant depend on the nature of gas.

So For ideal gas $PV = nRT$.

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

- this gas equation is called van der walls equation.
- Vanderwall Constant (a) :- If the intermolecular force is greater then the value of a is greater.
- Vanderwall Constant (b) :- It's value is a measure of the effective size of the gas molecule.

- Critical temperature - It may be defined as that temperature above which a gas Cannot be liquefied howsoever high Pressure may be applied on the gas.

— Critical Pressure - The Pressure required to liquefy the gas at the critical temperature is called critical pressure.
— Critical volume - The volume occupied by one mole of the gas at the critical temperature and the critical Pressure is called the critical volume.

— liquid - a liquid State may be considered as an intermediate State between the gas and the Solid.

Kinetic molecular theory of liquids.

(1) A liquid is made up of molecules.
(2) The molecules of the liquid are quite close together.
(3) The intermolecular force of attraction in a liquid are quite larger.
(4) The molecules of a liquid are in the State of constant rapid motion.
(5) The average kinetic energy of the molecules of a liquid is directly proportional to their absolute temperature.

— Properties
(1) Shape - No definite shape.
(2) Volume - have a definite volume.
(3) Density - much higher density than gas.
(4) Compressibility - less Compressible than gas.
(5) Diffusion - Liquids diffuse like gases but the diffusion is much Slower.
(6) Vapour Pressure - Vapour Pressure of a liquid at any temperature may be defined as the Pressure exerted by the vapour present above the liquid in equilibrium with the liquid at that temperature.

— Boiling Point - It is defined as the temperature at which the vapour Pressure of the liquid becomes equal to.

the external Pressure.

When the external Pressure is normal atmospheric Pressure the boiling Point is Called normal boiling Point.

When the external Pressure is equal to 1 bar, the boiling Point is Called Standard boiling Point.

Difference

	Evaporation.	Boiling.
1.	It takes Place Spontaneously at all temperature	It take Place at a Particuler temperature where Vapour Pressure is equal to atmospheric Pressure.
2.	Evaporation takes Place only from the Surface of the liquid.	It takes Place throughout the liquid. Boiling involves the formation of bubbles of the vapour below the Surface of the liquid.

— Heat of Vaporisation — The amount of heat required to Chang 1 mole of the liquid into its vapour at the boiling Point is Called the heat of Vaporisation of the liquid.

— Surface tension — Surface tension of a liquid is defined as the force acting at right angle to the Surface along one Centimeter length of the Surface so the unit of Surface tension is N/m^2.

— Surface energy — The work is required to be done in increase or extend the Surface area is Called Surface energy.

— Viscosity — the internal resistance to flow Possessed by a liquid is Called the Viscosity.

or.

The force of frichon which one Part of the liquid offers to another Part of the liquid is Called Viscosity.

the force of friction(f) between two layers each having area(A) cm^2 separated by a distance dx cm. and having a velocity difference dv cm/sec. then.

$$f \propto A \cdot \frac{dv}{dx} \quad \text{or} \quad f = \eta A \frac{dv}{dx}.$$

where η is a constant called viscosity constant. It may be defined as the force of friction required to maintain a velocity difference between two Parallel layers.

— Surfactants – Soaps and detergents when added to water decrease its surface tension are called surfactants

— Trouton's Rule – liquids which do not have too high boiling points, the ratio of the heat of vaporisation to the boiling point is approximately 88 J/K.

— Guldberg's rule – The boiling point of a liquid is nearly two third of its critical temperature when both are expressed on absolute scale. $T_b = \frac{2}{3} T_c$

— Effect of temperature on surface tension – (Eotvos equation)

Surface tension $(\gamma) = K(T_c - t)\left(\frac{\rho}{M}\right)^{2/3}$.

where ρ is density, M is molecular mass and t_c is the critical temperature, K is a constant, and t is the measurement temperature.

— Effect of temperature on viscosity (Arrhenius equation)

Viscosity $\eta = A e^{Ea/RT}$.

where A and Ea are constant for the given liquid and Ea is called activation energy for viscous flow.

— Critical Constants in term of Vandar walls Constants—

$$P_c = \frac{a}{27b^2} \quad ; \quad T_c = \frac{8a}{27Rb} \quad ; \quad V_c = 3b.$$

— Boyle temperature and critical temperature in term of Vardariealls Constant.

$$T_B = \frac{a}{Rb} \quad \text{and} \quad T_c = \frac{8a}{27Rb} \quad \text{so} \quad T_B > T_c$$

— Inversion temperature - Every gas has a definite temperature at which the gas shows neither heating effect nor cooling effect when allow to expand adiabatically. Above this temperature the gas show heating effect and below it show cooling effect. this temperature is called inversion temperature.

$$T_i = \frac{2a}{Rb}.$$

— Collision frequency - The number of collisions which take place in one second among the molecules present in 1cc of the gas is called Collision frequency and is denoted by Z.

— Collision number - The number of Collision which a single molecule undergo with other molecules in one second is called Collision number and is denoted by Nc

— Mean free Path - The distance travelled by a molecule between any two successive collisions is called free path. The average of these distance is called mean free path.

— Amagat's law of Partial volume - The total volume of a mixture of non-reacting gases is the sum of their Partial volumes, where Partial volume of a gas is the volume occupied by that gas at the same temperature and the Pressure of the mixture

$$V_T = V_1 + V_2 + \cdots$$

— measurment of surface tension — The instrument is called Stalgmometer, it is based upon the Principle that the surface tension of a liquid is directly proportional to the mass of the spherical drop falling from a capillary tube held vertically $\frac{\gamma_1}{\gamma_2} = \frac{m_1}{m_2}$.

— measurment of viscosity — The instrument is called Ostwald viscometer. It is based upon the Principle that for the same volume of the two liquids flowing from the same height and through the same capillary.

$\frac{\eta_1}{\eta_2} = \frac{d_1 t_1}{d_2 t_2}$ where d_1 and d_2 are densities and t_1, t_2 are time of flow.

— Clausius - Clapeyron equation.

$$\log \frac{P_2}{P_1} = \frac{\Delta H}{2.303 R} \left(\frac{T_2 - T_1}{T_1 T_2} \right)$$; ΔH = enthalpy of vaporisation of liquid.

Q. What is the difference between total kinetic energy and translational kinetic energy.

An. Total kinetic energy is the sum of translational, vibrational, and rotational kinetic energy. The total kinetic energy is equal to the translational kinetic energy for monoatomic gases.

Q. CO_2 is heavier than O_2 and N_2 but it does not form the lower layer of the atmosphere, why.

An. Gases possess the property of diffusion which is independent of the force of gravitation. Due to diffusion the gases mix into each other and remain almost uniformly distributed in the atmosphere.

Q. What would happen to the gas if the molecular collisions.

were not elastic.

An. on every collision, there would be loss of energy. As a result the molecules have slowed down and finally settle down in the vessel. and Pressure have gradually reduced to zero.

Q. what is the effect of temperature on surface tension and viscosity.

An. Both decrease with increase of temperature.

Q. Why falling liquid drops are spherical.

An. this is due to Property of surface tension. This makes the surface area minimum for a given volume. sphere has the minimum surface area.

Q. what happens if a liquid is heated to the critical temperature

An. The meniscus between the liquid and the vapour dis-appear and surface tension of the liquid becomes zero.