ARJUNA NEET BATCH

States of Matter

LECTURE - 9

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Quick Recap

Vanderwaal gas Eq^n

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

n = mole

$a \rightarrow$ Unit $\rightarrow$ atm L^2 mol^{-2}

$b \rightarrow$ Unit $\rightarrow$ L mol^{-1}

'b' signifies Molecular size

Compressibility $\downarrow$

Chemistry Syllabus for Test-2

$\rightarrow$ Some Basic Concepts of Chemistry

Matter

Objective of today's class

COMPRESSIBILITY FACTOR. (Z)

[LIQUIDSTATE]

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Compressibility factor (Z) → It is the ratio of Volume of real gas to the Volume of ideal gas.

$$Z = \frac{V_{Real}}{V_{ideal}}$$

Acc. to ideal gas Eq^n

$$PV = nRT$$

$$V_{ideal} = \frac{nRT}{P}$$

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

For $n=1$

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

if $Z=1$

$$\boxed{PV = RT}$$

↓

u.

Z

- $Z=1$
  → Behave as ideal gas
- $Z \neq 1$
  - $Z<1$
    -ive deviation from ideal gas behaviour.
    a(↑) → greater FOA
    → Easy to Compress
    → Easy to liquify
    → Temporary gas (CH_4, NH_3, SO_2)
  - $Z>1$
    +ive deviation from ideal behaviour
    → Permanent gases (H_2, He)

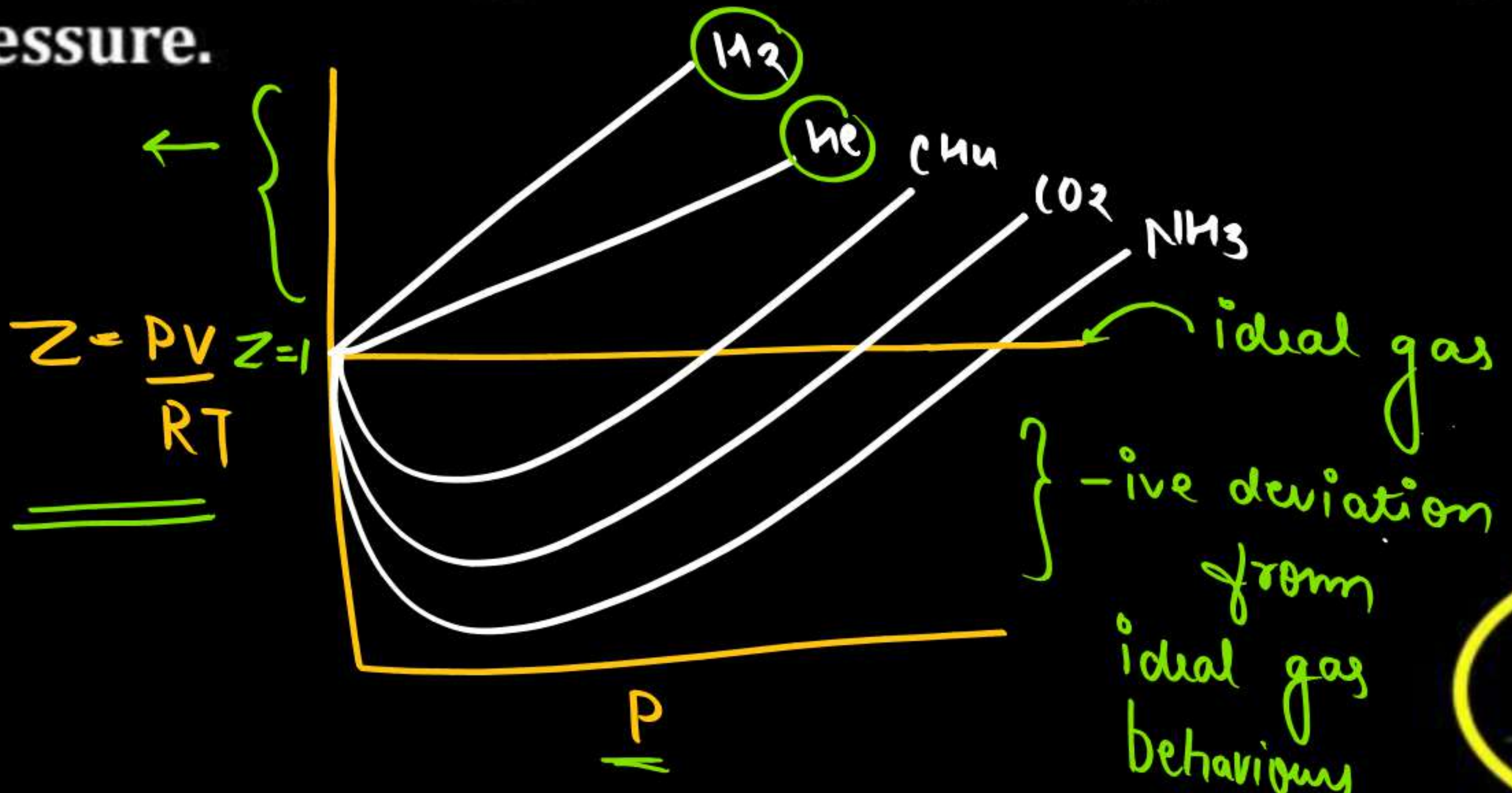
COMPRESSIBILITY FACTOR (Z)

Compressibility factor is the factor which decides the extent of derivation of real gases from the ideal gas behavior.

deviation

A graph is plotted for different gases between compressibility factor and over a range of pressure.

+ive deviation from ideal gas beho.



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Conclusion made from the graph:

(i) $z = 1$: compressibility factor is equal to unity for ideal gases at all temperature and pressure values.

Reason: For ideal gases, gas equation is $PV = nRT$. As PV is equal to nRT, so their ratio is unity.

(ii) $Z \approx 1$: compressibility factor is almost equal to unity at low pressure.

Reason: At low pressure, the interactions between the molecules is negligible, therefore there is no appreciable effect on the volume. So, $Z = 1$

(iii) Z > 1 : compressibility factor is greater than unity (positive deviation) at high pressure. Eg H_2, He (Permanent gases)

Reason: At high pressure, the gas molecules come very close to each other, so the forces of repulsion start operating between them. For ideal gases, as the pressure increases, volume decreases proportionally ($P \propto 1/V$) so the product PV remains constant. But this is not the case with real gases. In them, with increase in pressure (high pressure) the forces of repulsion does not allow the volume to decrease proportionally. Instead volume starts increasing. There by, as both pressure and volume are increased. So, the product PV also increase with pressure (PV > 1).

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

It can be better understood from the following derivation.

$$Z = \frac{PV_{real}}{nRT} \quad \text{....(i)}$$

Gas showing ideal behavior $V_{ideal} = \dfrac{nRT}{P}$ ....(ii)

Putting the value of nRT/P from equation (ii) in equation (i), we get

$$Z = \frac{V_{real}}{V_{ideal}}$$

So, the Z value depends upon the ratio of actual volume of the gas to the volume of ideal gas at that temperature.

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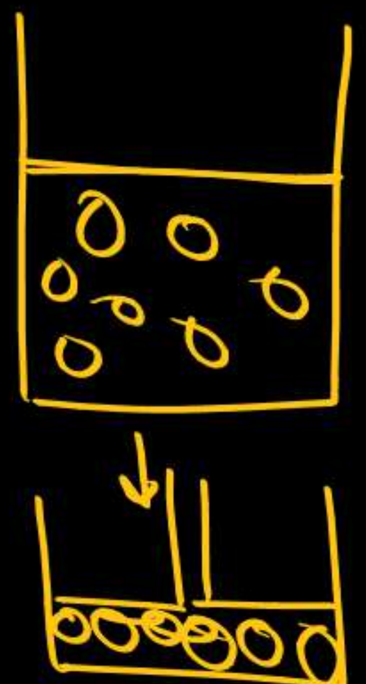
(iv) $Z < 1$: compressibility factor is less than unity $\Rightarrow 1$ (Negative deviation) when **pressure is intermediate.**

Eg: CH_4, SO_2, NH_3 (Temporary gases)

Reason: When the pressure is intermediate, then the gas molecules are at a sufficient distance to avoid repulsion. Instead, forces of attraction operate. Because of this, the molecules attract each other and they come closer. So, the volume decreases more than expected or decreases with increasing pressure. ($PV < 1$). So, the ratio of PV/nRT becomes less than 1.



At such pressure condition, the gases are more compressible due to forces of attraction.

→ Easy to Compress, Easy to liquify

So, it is concluded that behavior of gases becomes ideal when pressure is very low. Thus gases show ideal behavior when the volume occupied is large so that the volume of the molecules can be neglected in comparison to it.

	Compressibility	Pressure	Compressibility
(1)	$Z = 1$	All	Normal
(2)	$Z = 1$	Low	Negligible
(3)	$Z > 1$ H_2, He (Permanent)	High	Difficult
(4)	$Z < 1$ CH_4, NH_3, SO_2, HCl, 1 (Temporary)	Intermediate	Easy

Pressure alone does not decide the behavior of gases. Temperature is also an important factor which plays role.

Boyle Temperature or Boyle point: It is the temperature at which real gases obey ideal gas laws over an appreciable range of pressure.

$$\boxed{T_b = \frac{a}{Rb}}$$

$T_b \rightarrow$ Boyle's Temp. $R \rightarrow$ Universal gas const.

$a, b \rightarrow$ Vanderwaal const

Boyle temperature depends on the nature of the gas.

$T_b \propto a$, ↑, FOA ↑, Easy to liquify, Easy to compress

$T_b \propto \frac{1}{b}$

Effect of temperature on the compressibility factor.

(i) **Above the Boyle temperature:** Real gas show positive derivation ($Z > 1$) from ideality. This is because with increase in temperature, the molecules move far from each other. So, value increase thereby the forces of attraction between the molecules become feeble.

BELOW

(ii) **~~Above~~ the Boyle temperature:** Below the Boyle temperature 'Z' value first decreases and reaches and reaches a minimum value with increase in pressure because of forces of attraction which start operating between the molecules. Later, on further increase in pressure force of repulsion operate. So, now value of Z increases continuously.

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Conclusions:-

Gases show ideal behavior at

(i) High temperature

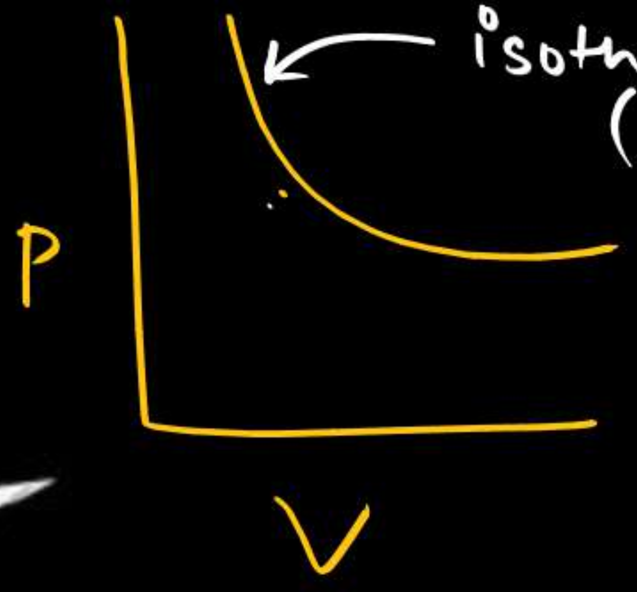
(ii) Low Pressure.

LIQUEFACTION OF GASES

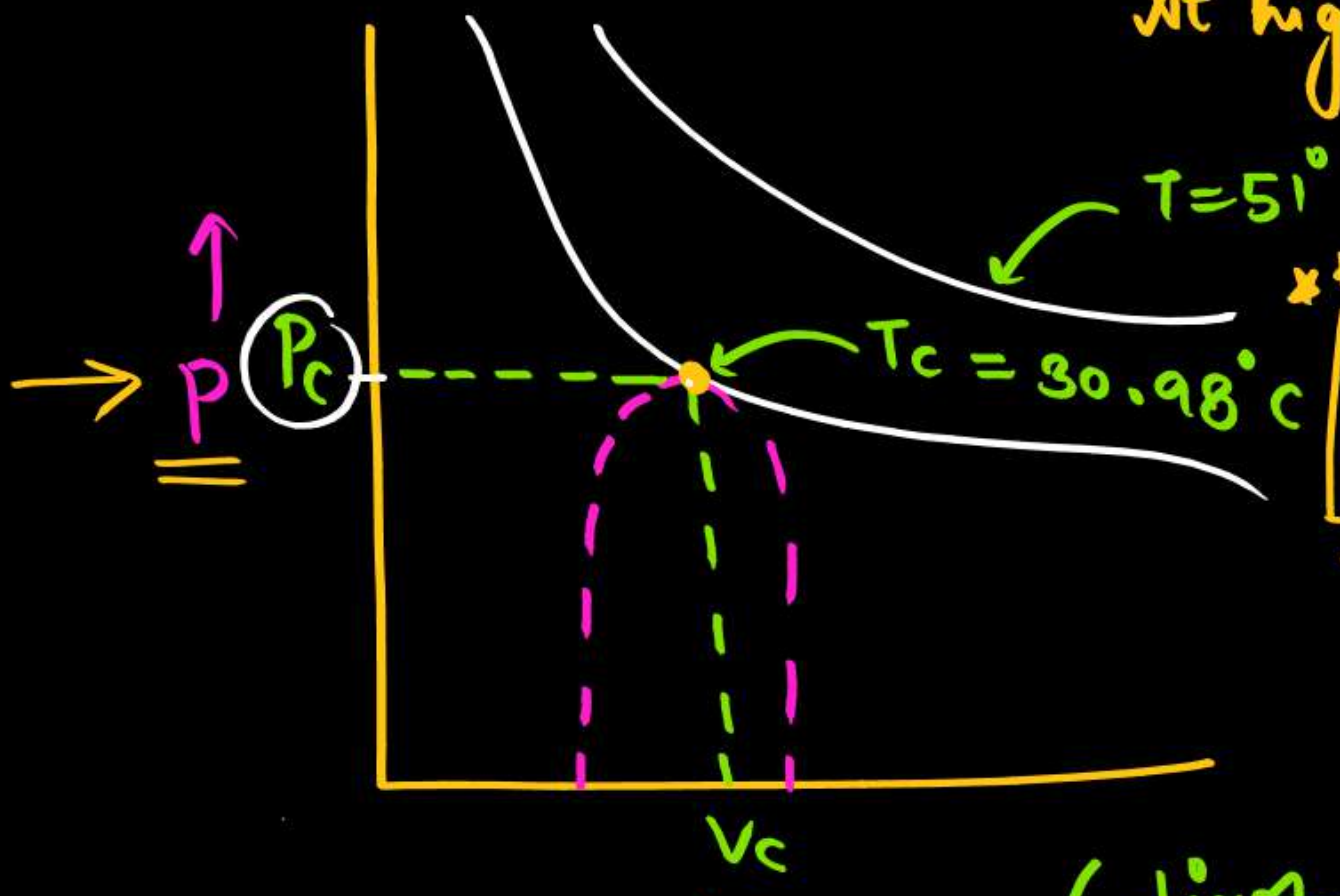
(i) Low temperature

(ii) High pressure

ANDREWS → CO_2 → Liquid state



isotherm (at const temp.)



At high temp. behave as a ideal gas

P-V isotherm for CO_2

(given by Andrews)

(first gas i.e converted into liquid state)

(i) **Critical temperature (T_C)** : Critical temperature of a gas may be defined as that temperature above when the gas cannot be liquefied however high the pressure is applied on the gas.

which

$$\boxed{T_C = \frac{8a}{27Rb}}$$

$T_C \rightarrow$ Critical temp

$R \rightarrow$ gas const.

$a, b \rightarrow$ Vanderwaal const

$T_C \propto a$

(ii) **Joule-Thomson effect** : It states that if a gas expands adiabatically under high pressure through a fine orifice into a region of low pressure results in cooling of gas.

$\rightarrow$ Joule Thomson effect observed due to adiabatic expansion of real gas, gaseous molecules spend energy against intermolecular force of attraction therefore temperature decreases.

heat = 0

(iii) Inversion temperature (T_I) : The temperature at which there is no heating or cooling of a gas occur during adiabatic expansion based on Joule Thomson effect. But the temperature below inversion temperature result in cooling of gas and above inversion temperature results in heating of gas.

$$\boxed{T_I^{\circ} = \frac{2a}{bR}}$$

$$\boxed{T_i^{\circ} = \frac{2a}{Rb}}$$

→ T_I° = Inversion temperature

→ a & b = van der waal's constant

→ R = Gas constant

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(iv) **Boyle's temperature (T_b) or Boyle's point :** The temperature at which real gases behave like ideal gas for a long range of pressure.

$$\boxed{T_b = \frac{a}{bR}}$$

$$T_b = \frac{T_I}{2}$$

or $T_I = 2T_b$

$$T_c = \frac{8a}{27Rb} \qquad T_i = \frac{2a}{Rb} \qquad T_b = \frac{a}{Rb}$$

$$\boxed{T_i > T_b > T_c}$$

(v) **Critical pressure (P_C) :** The minimum pressure that must be applied on a gas at its critical temperature just liquefy it.

$$\boxed{P_C = \frac{a}{27b^2}}$$

liquify

CRITICAL VOLUME(V_c)

(vi) The volume occupied by one mole of a gas at its critical temperature and critical pressure is known as the critical volume.

$$\boxed{V_C = 3b}$$

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Value of Compressibility FACTOR (Z) IN TERMS of P_c, T_c, & V_c

$$Z = \frac{P_c V_c}{R T_c}$$

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

$$Z = \frac{\cancel{a}}{\cancel{27}\cancel{b^2}} \times \frac{3\cancel{b} \times \cancel{27}\cancel{R}\cancel{b}}{\cancel{R} \times 8\cancel{a}}$$

$$Z = \frac{3}{8}$$

LIQUID STATE

(1)

The intermolecular forces in the liquid state molecules are stronger than in the gaseous state molecules.

(2)

Due to strong intermolecular force, there is less space between the molecules, so their density is higher than that of gases.

(3)

Liquids have definite volume.

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

$V \downarrow \quad d \uparrow$

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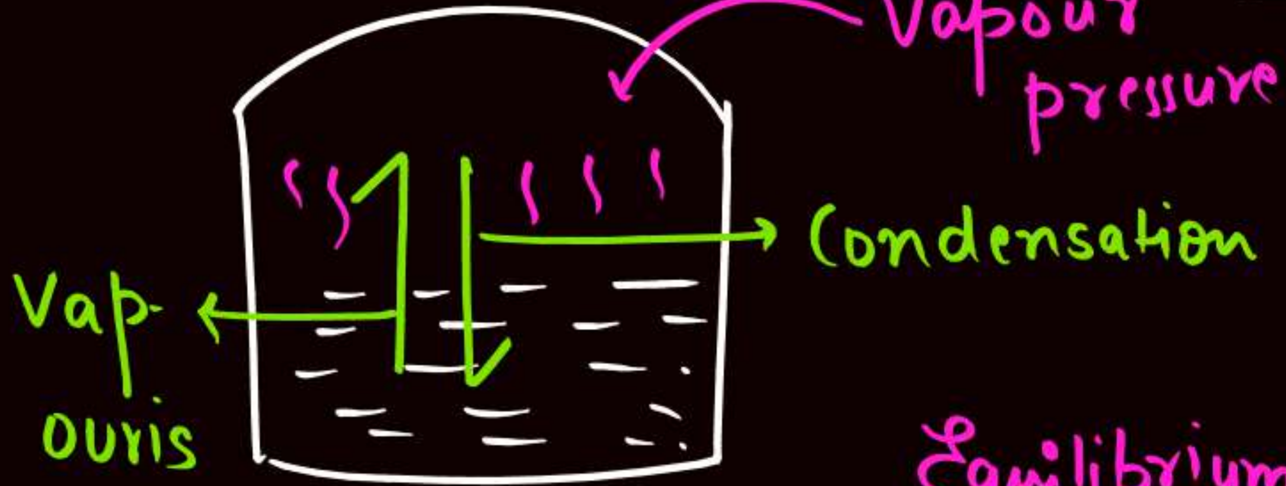
4 Viscosity

1 Vapour Pressure



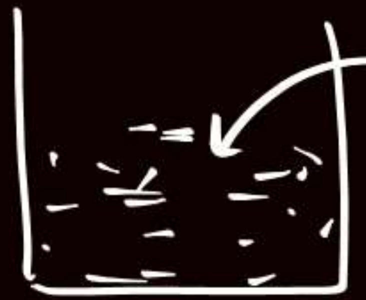
Equilibrium

Rati of vapourisation $\rightleftharpoons$ Rati of Condn.

→ 1 Intermolecu... ces $\propto \frac{1}{\text{Vapour pressure}}$

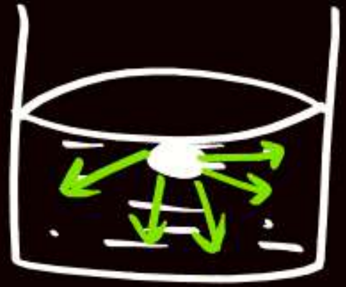
2 Temperature

2 Boiling Point



- Standard B.P.
- Normal B.P.

3 Surface Tension

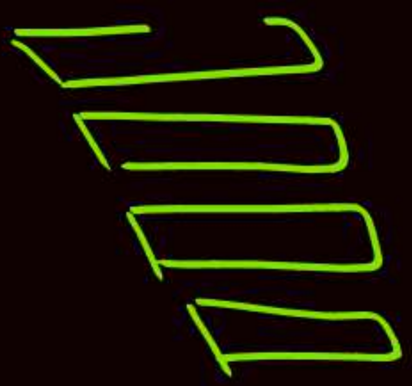


→ Intermolecular Forces $\propto$ S.T.

→ Temp $\propto \frac{1}{\text{S.T.}}$

④ Viscosity

H_2O Honey Colgate



→ Intermolecular Forces ↑ Viscosity

→ Temp ↑ Viscosity ↓

VAPOUR PRESSURE

When the rate of evaporation equal to rate of condensation i.e. equilibrium is established, the pressure exerted by the vapours of liquid on its on surface is known as $\boxed{\text{vapour pressure.}}$

The magnitude of vapour pressure depends upon the following points.

(a) Nature of liquid

(b) Temperature of liquid $\uparrow$ V.P $\uparrow$

Intermolecular Forces $\uparrow$

Vapour Pressure $\downarrow$

Rate of evaporation =

Rate of Condensation.

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Boiling Point : Boiling point of the liquid is the temperature at which the vapour pressure of the liquid is equal to the atmospheric pressure.

(i) **Normal Boiling Point :** As the atmospheric pressure varies with altitude and other conditions, the boiling points are reported at 1 atm. So, the normal boiling point of a liquid is the temperature at which the vapour pressure of the liquid is 1 atm.

e.g. Normal boiling point of water = 100°C

P = 1 atm

Standard Boiling Point: It is the temperature at which the vapour pressure of the liquid is $\boxed{1 \text{ bar.}}$

e.g. Standard boiling point of water $\boxed{= 99.6°C}$

Boiling: Boiling is a Bulk Phenomenon.

Boiling occurs in an open vessel. It cannot occur in a closed vessel.

Normal B.pt > Standard B.pt

$T = 100°C$ | $T = 99.6°C$

$P = 1$ atm | $P = 1$ bar

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Working of a Domestic Cooker : At high altitudes, the atmospheric pressure is low, so the vapour pressure of the liquid becomes equal to the atmospheric pressure at a very low temperature (i.e., much below the boiling point). That means the liquid (water) boils much before the food actually gets cooked. So, to avoid this situation, pressure cookers are used. Pressure cookers acts in away to raise the pressure, so as to increase the boiling point of a liquid (water). This results in efficient cooking of the food is shorter time.

SURFACE TENSION

This property is caused due to the strong intermolecular forces of attraction between the liquid molecules.

Surface tension may be defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid.]

Dimension = kgs^{-2}

$S.A = \frac{F}{l}$

S.I Unit = Newton per metre (Nm^{-1})

C.G.S. Unit = Dynes per centimeter (dyne cm^{-1})

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Energy

The surface required to increase the surface area of the liquid by one unit is called surface energy of the liquid.

Dimension = Jm^{-2}

Surface tension is dependent upon two factors:

(a) Nature of liquid → Intermolecular forces $\propto$ Surface tension

(b) Temperature → Temp ↑

Surface tension ↓

(a) **Nature of liquid :** Surface tension depends upon the nature of the liquid. More ~~are~~ the attractive intermolecular forces, more is the surface tension.

$\boxed{\text{Surface tension} \propto \text{Attractive Forces}}$

(b) **Temperature :** As the temperature increase, surface tension decrease. This is because, with the rise in temperature the kinetic energy of the liquid molecules increases.

$\boxed{\text{Surface tension} \propto 1/\text{Temperature}}$

Some Consequences of Surface Tension:

(a) Spherical shape of drops :

(b) Capillary Action :

(c) Liquid wets the things :

Cleansing action of soap and detergents is based on their lower surface tension. Soap solution, due to lower surface tension can penetrate into the fibre to surround the greasy substances and wash them away.

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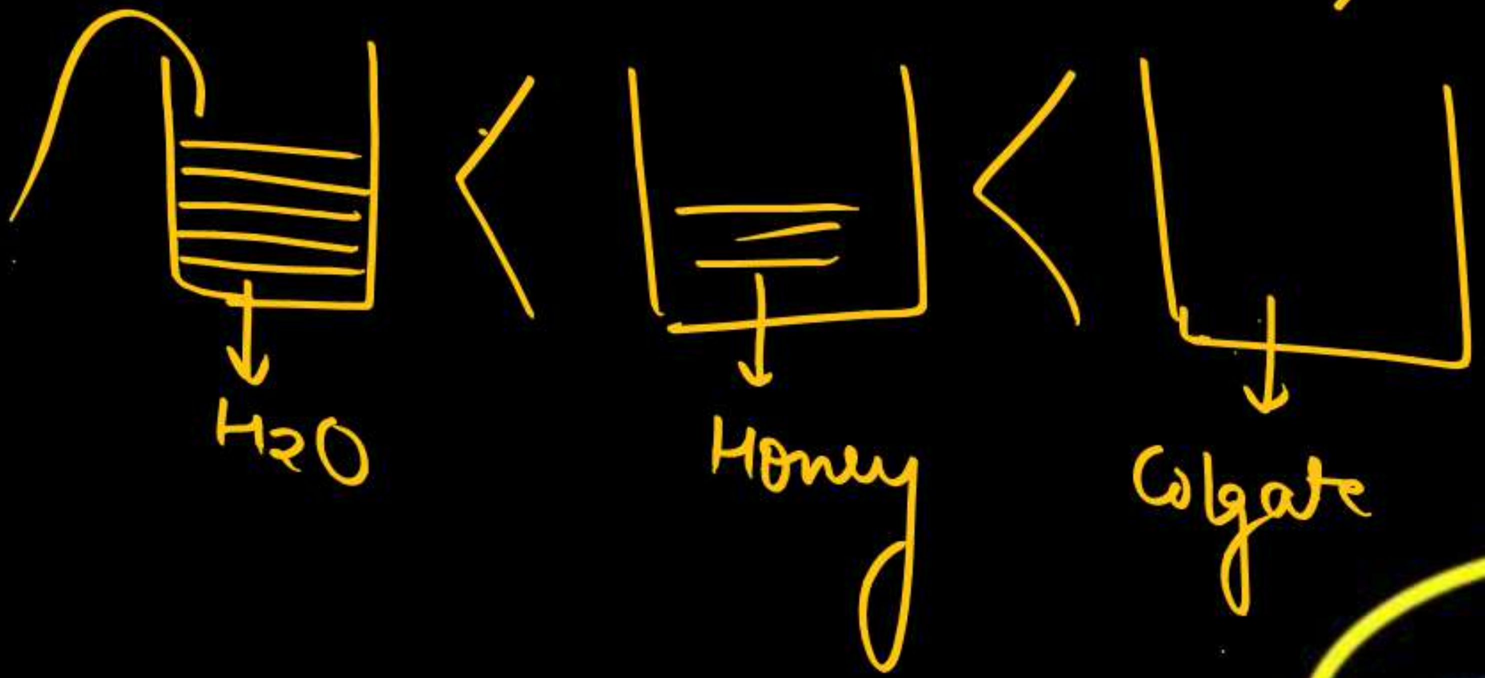
VISCOSITY

Viscosity is actually the measure of resistance to the flow of the liquid. arises

Viscosity is defined as the internal resistance to flow in liquids which aries due to the internal friction between the layers of liquid as they slip past one another while liquid flows.

? part



$$\Rightarrow \boxed{F \propto A.\frac{dv}{dz}}$$

F → force of friction

Where A = Area

$\frac{dv}{dz}$ = velocity gradient which is change of velocity with distance.

$$\boxed{F = \eta A \frac{dv}{dz}}$$

η → Viscosity Coefficient

Where 'η' is the proportionality constant. 'η' is a Greek letter (eta).

Coefficient of viscosity. Viscosity coefficient is the force when velocity gradient is unity and the area of contact is unit area.

UNITS

Imp.

S.I. unit = newton second per square metre $\boxed{(N\ s\ m^{-2})}$ = $\boxed{\text{Pascal second } (Pa\ s = 1\ kg\ m^{-1}\ s^{-1})}$

C.G.S. unit = Poise = $\boxed{1\ g\ cm^{-1}s^{-1}}$ = $10^{-1}\ kg\ m^{-1}s^{-1}$

Viscosity depends upon the following two factors:

(i) **Intermolecular attractive forces** $\propto$ Viscosity

(ii) **Temperature** $\propto$ 1/Viscosity

Viscosity in Everyday Life

Viscosity in Everyday Life:-

The effect of viscosity can be observed when we pour honey out of a jar. The honey near the jar's wall is stuck there and can't move. But even honey that's far away from the wall can't move easily. It is because the viscous force within the honey try to prevent any of the honey moving since the honey near the wall cannot move.

thanks for watching

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