STATES OF MATTER



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Solids, Liquids and Gases

S. No	Solid State	Liquid State	Gaseous State
1	Definite shape, mass and volume	Definite mass and volume but shape depends on the shape of the container	Definite mass but no definite shape and volume
2	Very high density	High density but less than solids	Low density
3	Very small change in volume occurs due to change in temperature and pressure	Small change in volume occurs due to change in temperature and pressure	Considerable change in volume occurs due to change in temperature and pressure
4	Molecules or constituents of solid possess less energy	Molecules of liquid have high energy	Molecules of gas are most energetic

ELEVEN ELEMENTS THAT EXIST AS GASES

Group number	1	15	16	17	18
	Н				Не
		Ν	0	F	Ne
				CI	Ar
					Kr
					Xe
					Rn

Gaseous state-assumptions

(1)The molecules are very loosely packed

(2) Intermolecular forces negligible.

(3) Molecules move very rapidly in all directions in a random manner.

(4) The molecules collide with one another and also with walls of the container perfectly elastic collisions.

(5) The mass, volume and pressure of gases can be calculated.

Measurable Properties of Gases

Mass Nu

Number of moles =

Mass of the gas

Molecular weight of the same gas

Volume

Volume of a gas is the space occupied by its molecules and is equal to the volume of the container. 1m = 10 dm = 100 cm $1m^3 = 10^3 \text{ dm}^3 = 10^6 \text{ cm}^3$ $1L = 1 \text{ dm}^3 = 10^3 \text{ cm}^3$

Pressure

Pressure is measured in terms of **atmospheric pressure**, which is the pressure exerted by the atmospheric gases on the surface of the earth.

- 1 atm = 76 cm of mercury = 760 mm of mercury
- 1 atm = 101.325 kPa

Measurable Properties of Gases

Temperature

Centigrade or Celsius scale (°C) Fahrenheit Scale (°F) Kelvin scale (K)

100° C = 180° F

$-273.15^{\circ} \text{ C} = 0 \text{ K}$ or $0^{\circ} \text{ C} = 273.15 \text{ K}$ and $t^{\circ} \text{ C} = (t + 273.15) \text{ K}$

STP or NTP

Temperature = 0° C = 273.15 K = 273 K Pressure = 1 atm = 76 cm = 760 mm = 101.325 kPa The temperature which is same on Fahrenheit and Celsius scale is

(a) 40 (b) 90

(c) 30 (d) none

Solution: Let that temperature be x

$$\frac{C}{5} = \frac{F - 32}{9}$$

$$9C = 5(F - 32)$$

$$9x = 5x - 160$$

$$9x = 5x - 160$$

$$4x = -160$$
$$x = \frac{-160}{4}$$

X = -40

van der Waals forces

Attractive forces between neutral species (atoms or molecules, but not ions) are referred to as **van der Waals forces. van der Waals forces are Small, weak interactions between non-polar compounds** and may be of three types:

1.dipole-dipole

2. dipole/induced-dipole

3. induced-dipole/induced-dipole

1. DISPERSION FORCES OR LONDON FORCES OR INDUCED-DIPOLE/INDUCED-DIPOLE

Atoms and non polar molecules are electrically symmetrical and have no dipole moment because their electronic charge cloud is symmetrically distributed. But a dipole may develop momentarily even in such atoms and molecules.



2. Dipole - Dipole Forces

Dipole-dipole forces act between the molecules possessing permanent dipole. Ends of the dipoles possess "partial charges" and these charges are shown by Greek letter *delta* (δ). Partial charges are always less than the unit electronic charge (1.610⁻¹⁹ C). The polar molecules interact with neighbouring molecules. This interaction is stronger than the London forces but is weaker than ion-ion interaction because only partial charges are involved. The attractive force decreases with the increase of distance between the dipoles.



3. Dipole-Induced Dipole Forces

This type of attractive forces operate between the polar molecules having permanent dipole and the molecules lacking permanent dipole. Permanent dipole of the polar molecule induces dipole on the electrically neutral molecule by deforming its electronic cloud. Thus an induced dipole is developed in the other molecule. Induced dipole moment depends upon the dipole moment present in the permanent dipole and the polarisability of the electrically neutral molecule.



Boyle's Law



(at constant temperature and fixed mass)

$P_1V_1 = P_2V_2 = constant$ [at a constant temperature]





Graph of pressure, p vs. Volume, V of a gas at different temperatures.

Graph of pressure of a gas, p vs. $rac{1}{V}$

A sample of gas occupies 2 L under a pressure of 800 atmosphere. What will be its volume if the pressure is decreased to 500 atmosphere. Assume that the temperature of the gas sample does not change?

Solution:

 $V_1 = 2 L V_2 = ?$

 $P_1 = 800 \text{ atm} P_2 = 500 \text{ atm}$

According to Boyle's law $P_1V_1 = P_2V_2$

So,
$$V_2 = \frac{P_1 V_1}{P_2} = \frac{800 \times 2}{500} = 3.2 L$$

Charle's law

Charle's law: $V \propto T$ (at constant pressure and fixed mass)

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} (at constant pressure)$$



Volume vs Temperature

Gay Lussac's law At constant volume, pressure of a fixed amount of a gas varies directly with the temperature.

 $p \propto T$



Pressure vs temperature

Avogadro's law

$V \propto n$ (at constant pressure and temperature)

"equal volumes of all gases contain equal number of molecules under similar conditions of temperature and pressure".

Ideal gas equation



Avogadro's hypothesis,

Combining these, we get

$PV \propto nT$ or PV = nRT

 $V \propto n$

Ideal gas equation

The gas constant, R Units of gas constant (R)

 $R = 0.0821 \text{ L atm } \text{K}^{-1} \text{ mole}^{-1}$ R = 8.314 10⁷ ergs K⁻¹ mole⁻¹ = 8.314 JK⁻¹ mole⁻¹

R = 1.987 cal K⁻¹ mole⁻¹

 $R = \frac{8.314 \text{ eV } \text{k}^{-1} \text{ mol}^{-1}}{1.602 \times 10^{-19}}$

= 5.18910¹⁹ eVK⁻¹mol⁻¹

Do you know?

Gas constant per molecule is known as Boltzmann constant (K)

$$K = \frac{R}{N} = 1.38 \times 10^{-10} \text{ ergs/K/molecule}$$

or 1.38 × 10⁻²³ JK⁻¹ molecule⁻¹

Density and molar mass relation

Number of moles of gas : $n = \frac{PV}{RT}$

If W g of a gas of molecular mass M, occupies a volume V, under pressure P at temperature T, then $n = \frac{W}{M}$



Vapour density



Dalton's Law of Partial Pressure





$$P_{total} = (a + b + c) \frac{RT}{V}$$

$$P_{Total} = P_{He} + P_{O_2} + P_{CO_2}$$



At 27° C a cylinder of 20 litres capacity contains three gases He, O₂ and N₂ 0.502 g, 0.250 g and 1.00 g respectively. If all these gases behave ideally, calculate partial pressure of each gas as well as total pressure.

Solution:

Let number of moles of He, O_2 and N_2 be n_1 , n_2 , n_3 respectively.

 $\therefore n_1 = \frac{0.502}{4.0 \text{ g mole}^{-1}} = 0.1255 \text{ moles of He}$

 $n_2 = 0.25 \text{ g}/32.0 \text{ g mole}^{-1}$

- = 0.0078 moles of O_2
- $n_3 = 1.00 \text{ g}/28.0 \text{ mole}^{-1}$
 - = 0.0357 moles of N₂

Total number of moles in the gaseous mixture

n = 0.1255 + 0.0078 + 0.0357 = 0.169 moles



= (0.169 mole) (0.0821 L atm K⁻¹ mole⁻¹) (300 K)

20 L

= 0.208 atm

Partial pressure of helium

$$P_{He} = \frac{0.1255}{0.169} \times 0.208 = 0.1545 \text{ atm}$$

$$P_{O_2} = \frac{0.0078}{0.169} \times 0.208 = 0.0096atm$$

$$P_{N_2} = \frac{0.0357}{0.169} \times 0.208 = 0.044 \text{ atm}$$

Applications of Dalton's law of partial pressure

a. To determine the pressure of a dry gas

 $P_{dry gas} = p_{moist gas} - Aqueous Tension (at t^o C)$

b. To calculate partial pressure
 In a mixture of non-reacting gases,

Partial pressure = mole fraction x total pressure

Amagat's Law of Partial Volume

$$V = V_A + V_B + V_C + \dots$$
 (in mixture of gases)

According to Amagat's law,



Molecular mass of the mixture of gases

For example, air contains 79% nitrogen and 21% oxygen approximately. Now mass of one mole of air would be

 $(0.79 \times 28) + (0.21 \times 32) = 28.84$ gm/mole.

$$\therefore M_{mixture} = \sum M_i x_i$$

= S Molecular mass of a gas × mole fraction of a gas x_i (mole fraction of a gas) =

 $\frac{n_i}{\Sigma n_i} = \frac{\text{Moles of a gas}}{\text{Total moles of gas mixture}}$

A gas at a pressure of 5 atm is heated from 0° C to 546° C and compressed to one-third of its original volume. Hence final pressure is

(a) 10 atm (b) 30 atm

(c) 45 atm (d) 5 atm

Solution:

we know, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$$\frac{5\times V_1}{273} = \frac{P_2V_1}{3\times 819}$$

∴ P2 = 45 atm.

Hence, the answer is (c).

\clubsuit The density of O₂ at NTP will be

(a) 1.43 g/L (b) 1.45 g/L

(c) 1.55 g/L (d) 1.59 g/L

Solution: PV = nRT

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

 $\therefore PM = \rho RT$



Which of the following gases has the vapour density 14?

(a) O_2 (b) CO_2 (c) CO (d) NO

Solution: Vapour density = $\frac{\text{Molecular mass}}{2}$ If vapour density = 14 ∴ Molecular mass = 28 Hence, the answer is (c).

For a particular gas at NTP, the pressure will be __at 100° C. (a) 1.2 atm (b) 1.34 atm (c) 1.4 atm (d) 1.9 atm

Solution:

We know,
$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

 $\frac{1}{273} = \frac{P_2}{373}$

∴ P2 = 1.34 atm

Hence, the answer is (b).

Which of the following compounds is steam volatile?









The compound which forms intramolecular H-bonding has lower boiling point and hence, steam volatile.

Hence, the answer is (b).



A mixture CO and CO₂ is found to have a density of 2gL⁻¹ at 25°C and 740 torr. Find the composition of the mixture

Solution:

 $M_{\text{mixture}} = \frac{\rho RT}{P} = \frac{1.5 \times 0.082 \times 298 \times 760}{740}$ = 37.64 g mol⁻¹ Let x mole of CO and (1 - x) mole of CO₂ $\therefore x \times 28 + (1 - x) \times 44 = 37.64$ x = 0.3975%CO = 39.75 %CO₂ = 60.25
Graham's Law of Diffusion

(i) For same volume of two gases

$$\frac{r_1}{r_2} = \frac{\frac{v}{t_1}}{\frac{v}{t_2}} = \sqrt{\frac{M_2}{M_1}} \qquad \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

(ii) For two gases with **different volumes and same diffusion time**

$$\frac{r_{1}}{r_{2}} = \frac{\frac{v_{1}}{t}}{\frac{v_{2}}{t}} = \sqrt{\frac{M_{2}}{M_{1}}} \qquad \frac{v_{1}}{v_{2}} = \sqrt{\frac{M_{2}}{M_{1}}}$$

Effusion

It happens under pressure through a Small aperture.

Rate of effusion

$$r = \frac{\Delta PA}{\sqrt{2\pi RTM}}$$

A=area of aperture $\Delta P = P_i - P_o$ (For same gas) Normally considered against vacuum A balloon filled with ethylene (C₂H₄) is pricked with a sharp point & quickly dropped in a tank full of hydrogen at the same pressure. After a while the balloon will have

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(a) Shrunk(c) Completely collapsed
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(b) Enlarged(d) Remain unchanged in size

Solution

 $r \propto \frac{1}{\sqrt{M}}$

Since, molar mass of H_2 is much less than C_2H_4 . H_2 will diffuse into the balloon. Hence, answer is (b).

Kinetic molecular theory

 All gases are made up of very large number of extremely small particles called molecules.

- The actual volume of the molecules is negligible as compared to the total volume of the gas.
- The distances of separation between the molecules are so large that the forces of attraction or repulsion between them are negligible.

Kinetic molecular theory

- The molecules are in a constant state of motion in all directions. During their motion, they collide with one another and also the walls of the container.
- The molecular collisions are perfectly elastic.
- The pressure exerted by the gas is due to bombardment of the gas molecules on the walls of the container.

Average kinetic energy of molecule
∞
absolute temperature.

Kinetic Theory of Gases

The kinetic gas equation,

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

Where, m = mass of each gas molecule n' = number of gas molecules c = velocity of gas molecule

KE of one molecule =

$$\frac{1}{2}$$
mc²

Kinetic Theory of Gases

$$PV = \frac{1}{2}mc^2 \times \frac{2}{3}n'$$

$$\therefore \frac{3}{2}nRT = \frac{1}{2}mc^2n' \text{ (where } n = \text{number of moles,} \\ n' = \text{number of molecules)}$$

 $\frac{3}{2}RT = \frac{1}{2}mc^2 \cdot N_A \quad \text{(for one mole of a gas, n = 1 and n' = N_A)}$

= Average kinetic energy per mole



Average kinetic energy of one molecule =

$$\frac{3}{2} \times \frac{R}{N_A} \times T = \frac{3}{2}kT$$

k = Boltzmann constant

 $= 1.38 \times 10^{-16} \text{ ergs } \text{k}^{-1} \text{ molecule}^{-1}$

Calculate the average kinetic energy per molecule and total kinetic energy of 2 moles of an ideal gas at 25°C.

Solution:

Average KE per molecule of the gas = $\frac{3}{2} \times \frac{R}{N_A} \times T = \frac{3}{2} \times \frac{8.314}{6.023 \times 10^{23}} \times 298 = 6.17 \times 10^{-21} \text{ J}$

Average KE per mole of the gas = $\frac{3}{2}$ RT = $\frac{3}{2}$ x 8.314 x 298 = 3.72 kJ/mole

Total KE of 2 moles of the gas =
$$\frac{3}{2}$$
 RT x 2 = 7.44 kJ

Molecular velocity



Most probable velocity (C_{MP})

$$C_{MP} = \sqrt{\frac{2 \text{ RT}}{M}} = \sqrt{\frac{2 \text{ PV}}{M}} = \sqrt{\frac{2 \text{ P}}{\rho}}$$

Molecular velocity

Root mean square velocity (C_{RMS})

$$C_{\text{RMS}} = \sqrt{\frac{C_1^2 + C_2^2 + \dots + C_n^2}{N}}$$

$$C_{RMS} = \sqrt{\frac{3 RT}{M}} = \sqrt{\frac{3 PV}{M}} = \sqrt{\frac{3 P}{\rho}}$$

Interrelation of molecular velocities

 C_{Avg} : $C_{RMS} = 0.9213^{\circ}$

$$C_{MP}$$
 : $C_{RMS} = 0.8165$

 $C_{Avg:}$: $C_{MP} = 1.1286$

The gas molecules have root mean square velocity of 1000 m. What is its average velocity?

(a) 1000 m/s (b) 921.58 m/s

(c) 546 m/s (d) 960 m/s

Solution: $C_{RMS} = \sqrt{\frac{3RT}{M}} = 10^3 \text{ m/s} = ---(1)$ $C_{Avg} = \sqrt{\frac{8RT}{\pi M}} = ----(2)$ From (1) and (2) $C_{Avg} = \sqrt{\frac{8}{3\pi}} \times 10^3 = 921.5 \text{ m/s}$ ✓ Calculate the temperature at which root mean square velocity of SO₂ molecules is same as that of O₂ molecules at 27° C.

Solution:

 $C_{RMS} = \left| \frac{3RT}{M} \right|^{1/2}$ For O₂ at 27° C, C_{RMS}O₂ = $\left\{\frac{(3) (R) (300)}{32}\right\}^{1/2}$ For SO₂ at t° C, C_{RMS}SO₂ = $\left\{\frac{(3) (R) (273 + t)}{64}\right\}^{1/2}$ Since both these velocities are equal, $\left\{\frac{(3) (R) (300)}{32}\right\}^{1/2} = \left\{\frac{(3) (R) (273+t)}{64}\right\}^{1/2}$ or 600 = 273 + tor $t = 600 - 273 = 327^{\circ}$ C

Maxwell-Boltzmann velocity distribution



Maxwell Boltzmann distribution

Characteristic features of Maxwell's distribution curve

- A very small fraction of molecules has very low or very high speeds.
- The fraction of molecules possessing higher and higher speed goes on increasing till it reaches a peak. The fraction with still higher speed then goes on decreasing.
- The peak represents maximum fraction of molecules at that speed. This speed, corresponding to the peak in the curve is known as the **most probable speed**.
- On increasing temperature, the value of the most probable speed also increases.

> 20 dm³ of SO₂ diffuses through a porous partition in 60 s. What volume of O₂ will diffuse under similar conditions in 30 s?



Ans. 14.14 dm³

180 cm³ of an organic compound diffuses through a pinhole in vacuum in 15 minutes, while 120 cm³ of SO₂ under identical condition diffuses in 20 minutes. What is the molecular weight of the organic compound?



At what temperature will H₂ molecules have the same root mean square velocity as N₂ gas molecules at 27° C?

Solution:





28

Ans. 21.43 K

If a gas is expanded at constant temperature, the kinetic energy of the molecules

(a) remains same (b) will increase

(c) will decrease (d) None of these

Solution:

Since K \propto E T, the kinetic energy of molecules remains same

Hence, the answer is (a).

The relation between PV and kinetic energy of an ideal gas is PV = ____ KE



Solution:

From kinetic gas equation, $PV = \frac{1}{3}mn'c^2 = \frac{1}{2}mc^2 \times \frac{2}{3}n'$ $PV = \frac{2}{3}KE$ (for one molecule) Can we use vapour densities in place of densities in the formula?

Given
$$\frac{r_1}{r_2} = \sqrt{\frac{\rho_2}{\rho_1}}$$

Solution

Vapour density, V.D = $\frac{M}{2}$

Now, if we replace density with vapour density of two gases,

$$\frac{r_1}{r_2} = \sqrt{\frac{V.D._2}{V.D._1}} = \sqrt{\frac{\frac{M_2}{2}}{\frac{M_1}{2}}} = \sqrt{\frac{M_2}{M_1}}$$

Which is also valid.

So, we can replace density with vapour density.

Ans. Yes The ratio of root mean square velocities of SO₂ to He at 25° C is (a) 1 : 2 (b) 1 : 4 (c) 4 : 1 (d) 2 : 1 Solution: For SO₂, $C_{RMS} = \sqrt{\frac{3R \times 298}{64}}$ $C_{RMS} = \sqrt{\frac{3RT}{M}}$ 3R×298 For He, $C_{RMS} = \sqrt{\frac{3R \times 298}{4}}$ $\frac{C_{RMS}(SO_2)}{C_{RMS}(He)}$ 64 3R×298

4

$$=\sqrt{\frac{4}{64}}=1:4$$

Hence, the answer is (b).

There is no effect of _____ on the motion of gas molecules.

(a) temperature (b) pressure

(c) gravity (d) density

Solution:

According to kinetic theory of gases, velocity of gas molecules will not be affected by the gravity as these are considered to be point masses.

Hence, the answer is (c).

The root mean square velocity of an ideal gas at constant pressure varies with density (d) as

(a) d² (b) d (C)
$$\sqrt{d}$$
 (d) $\frac{1}{\sqrt{d}}$

Solution:

$$RMS = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3P}{d}}$$
$$V_{rms} \propto \frac{1}{\sqrt{d}}$$

Hence, answer is (d)

Deviation: Ideal behaviour

Gases behave ideally at low pressures and high temperatures.

For an ideal gas, PV = ZnRT and Z = 1, Z is compressibility factor, measures extent of ideal nature

If $Z \neq 1$, PV \neq nRT

These gases are real gases



Plot of pV vs p for real gas and ideal gas Plot of pressure vs volume for real gas and ideal gas

Deviation: Ideal behaviour



Reason for deviation

Near the liquefaction point there is greater deviation from ideal behaviour.

- 1. The volume of the gas molecule is negligible as compared to the total volume of the gas.
- 2. The forces of attraction between gas molecules are negligible.

van der Waals' equation

Molecule equally attracted in all directions Molecule strikes wall with

lesser velocity

van der Waals' equation

$\therefore P_{\text{Ideal}} = P_{\text{real}} + p, \text{ where}$

p = Correction of pressure

Intermolecular forces of attraction proportional to

(i) the number of molecules per unit volume in the bulk of the gas.

(ii) the number of molecules colliding per unit volume.

Hence
$$p \propto \frac{n^2}{V^2}$$
 $p = \frac{an^2}{V^2}$

$$P_{ideal} = \left(P_{real} + \frac{an^2}{V^2}\right)$$

a is van der Waals' constant

van der Waals' equation



Excluded volume for a pair of molecules = $\frac{4}{3}\pi(2r)^3 = 8 \times \frac{4}{3}\pi r^3$ V_m = Actual volume of a gas molecule. Excluded volume for one mole b = 4 × V_m × N_A. observed volume = (V – nb) (For n mole of gas)

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

Van der waals' equation for real gas Using van der Waals equation, calculate the constant 'a' when two mol of a gas confined in a four litre flask exerts a pressure of 11.0 atm at a temperature of 300 K. The value of 'b' is 0.05 litre/mol.

Solution:

van der Waals equation

 $\left(P + \frac{n^2 a}{v^2}\right) (v - nb) = nRT$

$$\left(11 + \frac{4a}{4^2}\right)\left(4 - 2 \times .05\right) = 2 \times .082 \times 300$$

 $a = 6.48 \text{ atm litre}^2 \text{ mol}^{-2}$

What is the significance of a and b?

Larger value of 'a' indicate higher intermolecular forces of attraction, i.e. liquefaction of gases become easier.

While value of 'b' indicates the effective size of the gas molecules.

Application

For extremely low pressure

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

At very low pressure, volume will be too high. $\frac{a}{v^2}$ and b can be neglected.

Hence, PV = RT (Ideal gas equation)

Application

For very high temperature

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

The volume will be extremely high

Hence negligible value of

PV = RT

Application

For hydrogen gas

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

$$PV - Pb + \frac{a}{V} - \frac{ab}{V^2} = RT$$
$$PV = RT + Pb - \frac{a}{V} + \frac{ab}{V^2}$$

Small size and less molecular mass makes $\frac{a}{V}$ and $\frac{ab}{V^2}$ negligible

PV = RT + Pb $\frac{PV}{RT} = 1 + \frac{Pb}{RT} \qquad i.e. \ Z = 1 + \frac{Pb}{RT}$
LIQUIFACTION OF GASES



Thetemperature 30.98 °C is called critical temperature (T_c) of carbon dioxide. This is the highest temperature at which liquid carbon dioxide is observed. Above this temperature it is gas. Volume of one mole of the gas at critical temperature is called **critical volume** (V_{c}) and pressure at this temperature is called critical pressure (p_c).

Critical states



At $T_B Z=1$ I.e., gases behave ideally

Every real gas has a characteristic value of T_B .

What is triple point?

Triple point is the point at which solid, liquid and vapour are in equilibrium with eachother.

Liquid State – Characteristic Properties

- The liquid state of matter has intermediate properties.
- Less orderly than solid state but more orderly than the gaseous state.

Liquid State

Liquid state is explained on the basisof kinetic theory model as follows:

- Liquids are made up of molecules.
- Molecules in a liquid are quite close to each other.
- Force of attraction between the molecules in a liquid is quite large.
- The molecules in a liquid are in a state of constant random motion.
- The average kinetic energy of the molecules in a liquid is to their absolute temperature.

Properties of Liquid

Diffusion

- There is diffusion in liquids but it is slower than in gases.
- Diffusion involves movement of molecules from higher concentration to lower concentration..

Compressibility

Liquids are relatively less compressible than gases.

Surface Tension



Fig. 5.13 Forces acting on a molecule on liquid surface and on a molecule inside the liquid Surface tension is a contractive tendency of the surface of a liquid that allows it to resist an external force. Surface tension is defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid. It is denoted by Greek letter γ.
It has dimensions of kg s⁻² and in SI unit it is

expressed as N m⁻¹.

 The energy required to increase the surface area of the liquid by one unit is defined as surface energy.
Liquid tends to rise (or fall) in the capillary because

of surface tension.

Increase in temperature increases the kinetic energy of the molecules and effectiveness of intermolecular attraction decreases, so surface tension decreases as the temperature increases.



Viscosity is a measure of resistance to flow which arises due to the internal friction between layers of fluid as they slip past one another while liquid flows. Strong intermolecular forces between molecules hold them together and resist movement of layers past one another.



If the velocity of the layer at a distance dz is changed by a value du then velocity gradient is given by the amount $\frac{du}{dz}$. A force is required to maintain the flow of layers. This force is proportional to the area of contact of layers and velocity gradient i.e.

 $F \propto A$ (A is the area of contact)

 $F \propto \frac{du}{dz}$ (where, $\frac{du}{dz}$ is velocity gradient; the change in velocity with distance)

$$F \propto A. \frac{\mathrm{du}}{\mathrm{dz}}$$

$$\Rightarrow F = \eta A \frac{\mathrm{du}}{\mathrm{dz}}$$

η ' is proportionality constant and is called **coefficient of viscosity**.

- SI unit of viscosity coefficient is 1 newton second per square metre (N s m⁻²) = pascal second (Pa s = 1kg m⁻¹s⁻¹). In cgs system the unit of coefficient of viscosity is poise.
- \geq 1 poise = 1 g cm-1s-1 = 10-1kg m⁻¹s⁻¹.
- \succ Greater the viscosity, the more slowly the liquid flows.
- Glass is an extremely viscous liquid. It is so viscous that many of its properties resemble solids.
- Viscosity of liquids decreases as the temperature rises because at high temperature molecules have high kinetic energy and can overcome the intermolecular forces.

> The value of critical coefficient is

(a)2.66	(b) 3.55
(c) 1.55	(d) 4.33

Solution:

Critical coefficient

$$=\frac{RT_{c}}{P_{c}V_{c}}=\frac{8}{3}=2.66$$

Hence, the answer is (a)

The value of van der Waals' constant 'a' for the gases O₂, N₂, NH₃, CH₄ are 1.36, 1.39, 4.17 and 2.25 litre² atm mol⁻². The gas which can be most easily liquefied is

(a) O₂

(b) N₂

(c) NH₃

(d) CH₄

Solution:

The gas with the highest value of 'a' is having highest intermolecular forces of attraction. Hence, it can be liquefied most easily.

Hence, the answer is (c).

- At the same temperature and pressure, which of the following gases will have the highest rate of diffusion?
- (a) Hydrogen
- (b) Oxygen
- (c) Methane
- (d) All will have the same rate of diffusion



Hence, the answer is (a).

The ratio between the rms velocity of H_2 at 50 K to that of O_2 at 800 K is

(a) 4 (b) 2 (c) 1 (d) $\frac{1}{4}$

$$C_{rms} = \sqrt{\frac{3RT}{M}} \qquad \therefore \quad C_{rms_1} = \sqrt{\frac{3R \times 50}{2}}$$

$$C_{rms_2} = \sqrt{\frac{3R \times 800}{32}}$$
 $\frac{C_{rms_1}}{C_{rms_2}} = \sqrt{\frac{50 \times 32}{800 \times 2}} = 1$

Hence, the answer is (c).

Three moles of CO₂ (a van der Waals' gas) occupy 10 L at 15 atm. Find out the temperature.

(Given $a = 3.59 L^2 atm mol^{-2}$, $b = 0.043 L mol^{-1}$)

According to van der Waals' gas equation,

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

$$\left(15 + \frac{3.59 \times 9}{10^2}\right) (10 - 3 \times 0.043) = 3 \times 0.082 \times T$$

∴T = 614.1 K

Ans. T = 614.1 K