States of Matter

Past Years NEET Trend



Investigation Report

TARGET EXAM	PREDICTED NO. OF MCQs	CRITICAL CONCEPTS
NEET	1-2	Maxwell's distribution of velocitiesCritical phenomenon and liquefaction of gases

Perfect Practice Plan

Topicwise Questions	Learning Plus	Multiconcept MCQs	NEET Past 10 Years Questions	Total MCQs
49	30	19	15	113

INTRODUCTION

Matter exists in three different states known as solid, liquid and gas.

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

(i)Intermolecular attractive forces.

(ii)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

INTERMOLECULAR FORCES

Inter molecular forces usually called as Vander Waal forces.

Intermolecular forces generates as the result of the neutral attractions of unlike charges or repulsions of like charges.

Intermolecular forces does not include the forces between oppositely charged ions and the forces that hold atoms of a molecule together (i.e covalent bond).

Inter molecular forces are of several types:

- (*i*) Ion dipole
- (ii) dipole dipole
- (iii) London or dispersion forces
- (iv) dipole- induced dipole
- (v) Ion induced dipole forces
- (vi) Hydrogen bond (special type of dipole-dipole forces)

(i) Ion-dipole Interactions

It is attractive force between ion and dipole of neutral molecules

Example: In water interaction between Na⁺ and H₂O (dipole molecule) or Cl⁻ and H₂O



Fig.: Ion-dipole interaction

The magnitude of interaction energy depends on the charge on the ion (z), its dipole moment (μ) and on the inverse square of the distance (r) between the ion and the dipole.

$$\therefore$$
 $E = z.\mu/r^2$

Ion - dipole forces are mainly important in aqueous solutions of ionic substances.

Example: NaCl in water

(ii) Dipole - Dipole Forces

Neutral but polar molecules experience dipole-dipole forces.

These forces are due to electrical interactions among dipoles on neighboring molecules.

The strength of a given dipole - dipole interaction depends on the sizes of the dipole moments involved.

As the molecule is more polar, dipole-dipole interactions are more and boiling point of substance will be more.

Strong dipole-dipole interactions operating between hydrogen and more electronegative fluorine, or oxygen or nitrogen is called hydrogen bond.

Dipole - Dipole interactions in solids is proportional to $\frac{1}{r^3}$

Dipole- Dipole interactions between rotating molecules is proportional to $\frac{1}{r^6}$ here 'r' is the distance between polar molecules.



More charge density towards chlorine



Fig.: Dipole-Dipole forces

(iii) London or Dispersion Forces

Instantaneous dipole on one atom can affect the electron distributions in neighbouring atoms and induce temporary dipoles in those neighbours. As a result weak attractive forces are developed known as London forces or dispersion forces.

London forces energies are in the range 1–10 kJ/mol.



Fig.: Dispersion Forces

The magnitude of London forces depends on polarisability and geometry.

A smaller molecule or atom is less polarisable and has smaller dispersion forces.

A larger molecule or heavier atom is more polarisable and has larger dispersion forces.

 F_2 has smaller dispersion forces hence it is gas where as I_2 has larger disperson forces and it is a solid.

More spread out shapes will have higher dispersion forces than those compact molecules that minimise molecular contact thus have lower dispersion forces

Example: n-pentane, b.p 2,2-dimethyl propane, b.p is 282.7K (compact chain)

London dispersion forces increases with size, molecular weight, surface area and number of electrons.

(iv) Dipole - Induced Forces

These forces are in between polar molecules and neutral molecule.

Permanent dipole of the polar molecule induces dipole on the electrically neutral molecule by deforming into electric cloud and attractive forces develop.



Fig.: Dipole-induced Forces

The magnitude of induced dipole moment also depends on the magnitude of the dipole moment of permanent dipole and polarisability of neutral molecule.

(v) Ion - induced Dipole Forces

These are the interacations between a non polar molecule polarised by an ion nearer to it. The strength of these forces depends on charge of the ion and the ease with which non polar molecule gets polarised

Example: I₂ and NO₃

(vi) Hydrogen Bond

It arises in molecules in which highly polar N–H, O–H or H–F bonds are present. H-Bonds are formed by highly electronegative elements like N, O and F. Strength of H-bond is determined by the Coulombic interaction between lone pair of e⁻ s of electronegative atom of one molecule and H-atom of other molecule. For example.

Fig.: Hydrogen bond

Thus, H-bonding is a special case of dipole-dipole interaction.

INTERMOLECULAR FORCES vs THERMAL ENERGY

Physical state of a substance depends on inter molecular forces that try to bring molecules closer.

The energy due to the motion of the atoms or molecules of the substance is called thermal energy.

Thermal energy is directly proportional to absolute temperature of the substance.

It is a measure of average kinetic energy of the molecules of the substance.

The movement of particles is called thermal motion.

When the thermal energy of the molecules is reduced by lowering the temperature the gases can be very easily liquified

If thermal energy predominates over intermolecular forces the substances would change from

$solid \rightarrow liquid \rightarrow gas$

If intermolecular forces predominate over thermal energy then substance change from

gas \rightarrow liquid \rightarrow solid

GASEOUS STATE

When the molecular forces of attraction between the particles of matter are minimum, the particles exist in gaseous state.

Properties of Gas

- Gases have neither definite shape nor definite volume
- Gases can undergo expansion and compression
- The intermolecular forces of attraction in gases are very low.
- The molecules of a gas moves with high velocities, gravitational forces have minimum influence on the velocities of gas molecules.
- Gases can expand freely into the space available to them. This property is known as **diffusion of gases**.
- Some substances which occur as solids or liquids at room temperature and atmospheric pressure exist as gases under suitable conditions. Such gases are called **vapours**.
- Gases intermix with each other and forms homogeneous mixtures.
- The standard conditions for a gas are also known as S.T.P. conditions or N.T.P. conditions.
- S.T.P conditions

Temperature	Pressure
$T = 0^{o} C$	P=1 bar
T = 273.15 K	$P = 10^5 \text{ pa}$

Parameters of gas

Mass, Volume, Pressure and temperature are the measurable properties of a gas.

(i) Mass (M)

The mass of 6.023×10^{23} molecules of a gas is known as gram molar mass.

For any gas
$$n = \frac{m}{M}$$

Where.

n = number of gram mole of gas

m = mass of gas in grams

M = molar mass of gas

(ii) volume (V)

In S.I. system, the volume of a gas is expressed in cubic meters (m^3)

1 litre = 10^3 c.c = 10^3 millilitres

1m³ = 1000 litres=10⁶ millilitres

The volume occupied by one gram mole or one mole of a gas under S.T.P. conditions is known as **gram molar volume** (GMV) or molar volume, it is 22.4 litres or 22,400 ml or 22,400 c.c.

According to latest standard values i.e., temperature is 273.15 K and pressure is 1 bar (10^5 pascals), molar volume is 22.711 lit. mol⁻¹.

The number of molecules present in 1 ml of a gas at S.T.P conditions is known as **Loschmidt number. i.e.**

 $\frac{N_A}{22400} = \frac{6.023 \times 10^{23}}{22400} = 2.688 \times 10^{19} \text{ molecules/ml}$

(iii) Pressure (P)

The pressure of a gas is defined as the force per unit area. $P = \frac{F}{A}$ The common unit of pressure is atmosphere.

1atm = 76 cm of Hg = 760 mm of Hg

The unit of high pressure is bar.

1 bar = 10^6 dynes/ cm² = 10^5 pascal = 0.987 atm

The S.I unit of pressure is pascal (Pa).

Pascal is defined as the pressure exerted when a force of 1 Newton acts on 1 square meter area

 $1 \text{ Pa} = 1 \text{ N/m}^2 = 1 \text{ kgm}^{-1} \text{ s}^{-1}$

The mass of the atmosphere pressing down on the earth's surface exerts pressure which is known as **atmospheric pressure**.

🖉 KEY NOTE

Hg is used as barometric liquid due to its high density and as it is non volatile at room temperature

(iv) Temperature (T)

The temperature of gas is measured in Celsius scale.

The temperature of a gas is also expressed in absolute scale or kelvin scale

The absolute zero is at -273.15° C or -273° C.

The temperature of a gas in absolute scale or kelvin scale is given by

 $A^{o}=C^{o}+273$ or $K=C^{o}+273$

KEY NOTE

At absolute zero

(a) Molecular motion in the gas ceases.

(b) Pressure of the gas becomes zero.

(*c*) Volume of the gas becomes zero.

(d) Kinetic energy of the gas becomes zero.

• On absolute scale, there are no negative values.

• Absolute scale is the lowest scale for gases.

GAS LAWS

The laws which relate the properties of gases like pressure (P), Volume (V), Mass (m) and temperature (T) are called gas laws.

Boyle's Laws : (Pressure Volume Relationship)

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.

At constant temperature, for a given mass of gas, the product of its volume and pressure is a constant.

$$P_1V_1 = P_2V_2$$

For a given mass of gas at constant temperature the density of the gas is proportional to the pressure acting on it.

Now since density
$$(d) = \frac{m}{v}$$

So, Boyle equation is written as

$$d = \frac{mp}{k} = k'P$$

Graphical representation of Boyle's law

The graphs plotted between the pressure and the volume of a given mass of gas at any constant temperature are called **Isotherms**.

The shape of the P vs V graph is hyperbola



At constant T, the graph P vs 1/V shows a **straight line** passing through origin.



At high pressures gases deviate from Boyle's law and under such conditions a straight line is not obtained in the graph

For a given mass of a gas at constant temperature the graph drawn between pressure (or V) and PV is a straight line parallel to the pressure (or volume) axis.



Gases obey Boyle's law at low pressure and high temperature.

TRAIN YOUR BRAIN

Q. 103 mL of carbon dioxide were collected at 27°C and 763 mm pressure. What will be its volume if the pressure is changed to 721 mm at the same temperature ?

Final

Ans. Given

ConditionsConditions $V_1 = 103 \text{ mL}$ $V_2 = ? \text{ mL}$

 $P_1 = 763 \text{ mm}$ $P_2 = 721 \text{ mm}$

By applying Boyle's Law since temperature is constant,

$$\mathbf{P}_2 \times \mathbf{V}_2 = \mathbf{P}_1 \times \mathbf{V}_1$$

Substituting the corresponding values, we have

$$721 \times V_2 = 763 \times 103$$

763×103

$$V_2 = \frac{703 \times 103}{721} = 109 \text{ mL}$$

 \therefore Volume of carbon dioxide = 109 mL

Charle's Law: (Temperature Volume Relationship)

Charle's law states that, at constant pressure, the volume of a given mass of gas is directly proportional to its absolute temperature. Mathematical expression of charles law is

 $V \propto T($ at constant P)

$$\frac{V}{T} = K$$

where, V = volume of gas

T = absolute temperature

If the volume of the gas V_1 and T_1 changes to V_2 and T_2 at constant P

$$\frac{\mathbf{V}_1}{\mathbf{T}_1} = \frac{\mathbf{V}_2}{\mathbf{T}_2}$$

At constant pressure, the volume of a given mass of a gas increases or decreases by 1/273.15 times its volume at 0 °C for every 1° C rise or fall in temperature

 $\mathbf{V}_{t} = \mathbf{V}_{0} \left(1 + \alpha t \right)$

where P is constant (This equation is Charle's-Gay Lussac equation)

where α = volume coefficient = 1/273.15

$$V_t = Volume at t \, ^{\circ}C$$

 $V_0 =$ Volume at 0 °C

Graphical representation of charle's law

Charles law graphs as known as isobar



The following graph gives relation between V & T at constant P.



____ TRAIN YOUR BRAIN __

Q. If the temp. of a particular amount of gas is increased from 27°C to 57°C, find final volume of the gas, if initial volume = 1 L and assume pressure is constant.

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

$$\frac{1}{(273+27)} = \frac{V_2}{(273+57)}$$
 So $V_2 = 1.1$ L.

Gay Lussac's Law or Amonton's Law

It is also known as PT law or Pressure Temperature Relationship.

At constant volume, the pressure of a given mass of gas is directly proportional to its absolute temperature,

$$\boxed{\underline{P} \propto T} \frac{\underline{P}}{T} = \text{constant} = K$$

Thus if pressure P₁ at T₁ changes to P₂ at T₂, then. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ Graphical representation of gay Lussac's law

The plots of P vs T are straight lines and passes through the origin



Fig.: Pressure vs temperature graph (Isochore)

At constant volume, the graph drawn in between P and T of gas is called **isochore.**

🖉 KEY NOTE

Charles law (VT-Law) not obeyed below -273°C because no substance exist as gas below the -273°C and changes to liquid

TRAIN YOUR BRAIN

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

Ans.
$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$
; $\frac{1}{T} = \frac{P_2}{2T} \implies P_2 = 2 \text{ atm}$
% increase $= \frac{P_2 - P_1}{P_1} \times 100 = \frac{1}{1} \times 100 = 100\%$

(4) Avogadro's Law (Volume-amount Relationship)

It states that equal volumes of all gases, measured under the same conditions of temperature and pressure, contains equal number of molecules or moles.

At constant temperature and pressure, the volume of a gas is proportional to the number of moles present in it.

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

$$\frac{\mathbf{V}}{\mathbf{n}} = \mathbf{K} \text{ and } \frac{\mathbf{V}_1}{\mathbf{n}_1} = \frac{\mathbf{V}_2}{\mathbf{n}_2}$$

IDEAL GAS EQUATION

A gas which obeys all the gas laws under all the conditions of temperature and pressure is called an **ideal gas.**

The ideal gas equation is obtained by combining the three gas laws. (Boyle's law, Charles law and Avogadro's law)

PV = **nRT** is ideal gas equation or perfect gas equation and also called as **equation of state**

Here n = number of moles of the gas

R = Universal gas constant

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

m = mass of the gas

M = Molecular weight of the gas

$$P = \frac{dRT}{M}$$
 here $\frac{m}{v} = d$

If T, V and P of a fixed amount of gas change from T_1 , V_1 and P_1 to T_2 , V_2 and P_2

$$\frac{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}}{\frac{P_1T_1}{P_1} = \frac{d_2T_2}{P_2}}$$
 called **combined gas law**

Numerical values of 'R'

R values depends upon the units of P,V,T.

$$R = \frac{PV}{nT}$$
, $R = \frac{latm \times 22.414 litre}{lmole \times 273 K}$

= 0.0821 lit. atm K⁻¹. mole⁻¹ = 82.1 ml. atm K⁻¹. mole⁻¹ $R = \frac{PV}{nT} = \frac{(10^5 \text{ Pa})(227 \times 10^{-3} \text{ m}^3)}{(1\text{mol})(273.15\text{ K})}$

= 8.314 Pa m³ K⁻¹ mol⁻¹ {1 Pa = 10^{-5} bar}

= 8.314×10^{-2} bar L K⁻¹ mol⁻¹

R value in different units

R-Value	Pressure unit	Volume unit
0.0821	atmosphere	litre
82.1	atmosphere	cm ³
8.314×10^{7}	dynes.cm ⁻²	cm ³
8.314	newton.m ⁻²	m ³
1.987		

• The gas constant per molecule is **known as Boltzman** constant (k).

$$\boxed{k = \frac{R}{N_A} = \frac{8.314 \times 10^7}{6.02 \times 10^{23}}}$$

= 1.38 × 10⁻¹⁶ ergs K⁻¹. molecule⁻¹

= 1.38×10^{-23} Joules K⁻¹. molecule⁻¹

Density and Molar Mass of Gaseous Substance

$$PV = nRT$$

$$\frac{n}{V} = \frac{P}{RT} \left(n = \frac{m}{M}\right)$$
So
$$\frac{m}{MV} = \frac{P}{RT}$$

$$\frac{d}{M} = \frac{P}{RT}$$
So Thus
$$M = \frac{dRT}{P}$$

Graham's Law of Diffusion

The spontaneous mixing of gases against gravitational forces to form homogeneous mixture is known as the **diffusion of gases**.

The volume of gas (V) that diffuses in unit time or the distance travelled by gas molecules per unit time is known as the **rate of diffusion (r) of** the gas

Rate of diffusion =
$$\frac{\text{Volume of the gas (V)}}{\text{Time(t)}}$$

Gases diffuses from high pressure to low pressure.

For two gases diffusing under similar conditions of temperature and pressure.

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{V_1 d_2}{V_2 d_1}} = \sqrt{\frac{M_2}{M_1}}$$

If two gases diffuses at different pressures but at similar temperature

<u>r</u> 1	<u>P</u> ₁	M_2
r ₂	P ₂ ∖	M_1

• Under similar conditions of temperature and pressure if equal volumes of two gases diffuses.

$$\frac{\mathbf{r}_1}{\mathbf{r}_2} = \frac{\mathbf{t}_2}{\mathbf{t}_1}, \quad \frac{\mathbf{t}_2}{\mathbf{t}_1} = \sqrt{\frac{\mathbf{M}_2}{\mathbf{M}_1}}$$

• For gases at different temperatures $\frac{r_1}{r_2} = \sqrt{\frac{T_1M_2}{T_2M_1}}$

TRAIN YOUR BRAIN

- **Q.** The pressure in a vessel that contained pure oxygen dropped from 2000 torr to 1500 torr in 40 min as the oxygen leaked through a small hole into a vacuum. When the same vessel was filled with another gas, the pressure dropped from 2000 torr to 1500 torr in 80 min. What is the molecular weight of the second gas?
- Ans. Change in pressure = 2000 1500 = 500 torr, time-taken = 40 min

$$rate_{1} = \frac{500}{40} \text{ torr/min; similarly } rate_{2} = \frac{500}{80} \text{ torr/min}$$
$$\frac{rate_{1}}{rate_{2}} = \sqrt{\frac{M}{32}} \implies \frac{500/40}{500/80} = \sqrt{\frac{M}{32}}$$
$$M = 128 \text{ g/mol}$$

Dalton's law of Partial Pressure

The total pressure exerted by a mixture of gases which do not chemically react is equal to the sum of the partial pressures of component gases present in the mixture at same temperature.

RT

1.
$$P = P_1 + P_2 + \dots + P_n$$

2. $P_n = (n + n + n + n)$

2.
$$I_{mix} = (I_1 + I_2 + I_3 + ...) \overline{V}$$

Partial Pressure

Pressure exerted by a gas, when it alone occupies the total volume of the mixture at the same temperature is called partial pressure.

P.P. of gas = total pressure \times mole fraction of gas

 $P_i = X_i \times P$

Mole fraction of a component

Let A and B are present in a mixture.

Mole fraction of A is $x_A = \frac{n_A}{n_A + n_B}$

Where $n_A - no.$ of moles of A

 $n_{\rm R}$ – no. of moles of B

Water insoluble gases are generally collected over water. A gas collected over water is saturated with water vapour such a gas is called moist gas.

The pressure of water vapour in a moist gas is known as aqueous tension.

$$P_{\text{moist gas}} = P_{\text{drvgas}} + \text{aqueous tension}$$

Aqueous tension is constant at given temperature and increases with increase in temperature.

🖉 KEY NOTE vapour pressure of H₂O

Limitations of Dalton's Law of Partial Pressure

Dalton's law of partial pressure is not applicable to the mixture of reacting gases

Examples:

1. CO and Cl₂ 2. NO and Cl₂

3. NO and O_2^2 4. H₂ and Cl₂

Dalton's law of partial pressure is applicable to the mixtures of non-reacting gases.

Examples:

- 1. N_2 and H_2
- 2. H_2 and O_2

TRAIN YOUR BRAIN

Q. A mixture of gases at 760 torr contains 55.0% nitrogen, 25.0% oxygen and 20.0% carbon dioxide by mole. What is the partial pressure of each gas in torr?

Ans. Let total moles of gas mixture be 100.

$$P_{N_2} = \left(\frac{n_{N_2}}{n_T}\right) \times P_T = \frac{55}{100} \times 760 = 418 \text{ torr.}$$
$$P_{O_2} = \left(\frac{n_{O_2}}{n_T}\right) \times P_T = \frac{25}{100} \times 760 = 190 \text{ torr.}$$

 $P_{CO_2} = (760 - 418 - 190) = 152$ torr.

 $P_{N_2} = 418$ torr, $P_{O_2} = 190$ torr, $P_{CO_2} = 152$ torr, total pressure = 760

Kinetic Molecular Theory of Gases

Kinetic molecular theory of gases was proposed by Maxwell, Boltzmann, Clausius etc.

Postulates of Kinetic Molecular Theory Are

Kinetic molecular theory is applicable only to ideal gases.

Gases contain large number of tiny and discrete particles called molecules.

The molecules move randomly in all the directions with high velocities.

The volume of molecules is negligible when compared to the total volume of the container.

There are no attractions or repulsions between the gas molecules.

Molecular motions are unaffected by gravity.

Pressure exerted by a gas is due to the collisions of molecules on the walls of the vessel.

The collisions between gas molecules are perfectly elastic i.e there is no loss of energy during molecular collisions.

The average kinetic energy is directly proportional to the absolute temperature of the gas

Kinetic Gas Equation

(a)
$$PV = \frac{1}{3}MC^2$$
 is called kinetic gas equation
(b) $P = \frac{1}{3}dC^2$

Where m = mass of the molecules

n = no. of molecules

$$M = Molar mass \left(\frac{m}{n}\right)$$
$$d = density$$

C = RMS velocity

The average kinetic energy per molecule of any gas is same at the same temperature

Kinetic Energy of gases

Kinetic energy of 'n' moles of gas

$$E_k = \frac{3}{2}nRT$$

The kinetic energy for one mole of the gas

$$E_{K} = \frac{3}{2}RT$$

The kinetic energy for one molecule of the gas

$$E_{K} = \frac{3}{2}kT (k = \frac{R}{N_{A}} = Boltzman constant)$$

The relation between K.E, moles and temperature.

$$\frac{\text{K.E}_1}{\text{K.E}_2} = \frac{n_1}{n_2} \times \frac{T_1}{T_2}$$

The relation between average K.E. and temperature

$$\frac{A.K.E_1}{A.K.E_2} = \frac{T_1}{T_2}$$

TRAIN YOUR BRAIN

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

Ans.
$$PV = \frac{1}{3} mN U_{rms}^2$$

 $P = ?$
 $V = 10^{-3} m^3$
 $m = 10^{-25} kg$
 $N = 10^{23}$
 $\sqrt{U^2} = 10^5 cm/sec = 10^3 m/sec$
 $\overline{U^2} = 10^6 m^2 /sec^2$
 $P \times 10^{-3} = \frac{1}{3} \times 10^{-25} \times 10^{23} \times 10^6$
 $P = \frac{1}{3} \times 10^{-2} \times 10^6 \times 10^3$
 $P = \frac{1}{3} \times 10^7 pascals$

DISTRIBUTION OF MOLECULAR VELOCITIES

Molecular velocities constantly change from a low value close to zero to a very high value.

The ratio of the number of molecules with a certain velocity to the total number of molecules always remains constant.

Maxwell-Boltzmann gave the distribution curves of molecular velocities.



Fig.: Maxwell speed Distribution plot and speeds

From the graph, (most proable velocity) $C_p <$ (average velocity) $\overline{C} <$ (RMS velocity) C

As velocity of the molecule increases, the fraction of the molecules possessing a particular velocity also increases upto a maximum value and then decrease.

As the temperature increases, the fraction of the molecules possessing low velocities decreases and the fraction of the molecules possessing high velocities increases.

Area covered under different curves is same as it represents the total number of molecules present in the gas

RMS Velocity (C)

The square root of the mean of squares of the velocities of all the molecules present in the gas at any temperature is known as the RMS velocity. It is represented by C

$$C = \sqrt{\frac{C_1^2 + C_2^2 + C_3^2 + \dots + C_n^2}{n}}$$
$$C = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3P}{d}}$$
$$C = 1.58 \times 10^4 \times \sqrt{\frac{T}{M}} \text{cm.sec}^{-1}$$

For two different gases at two different temperatures.

$$\frac{C_1}{C_2} = \sqrt{\frac{T_1}{T_2} \times \frac{M_2}{M_1}}$$

Average Velocity C

The average of the velocity of all the molecules in the gas at any temperature is known as average velocity. It is represented by \overline{C} .

$\overline{C} = \sqrt{\frac{8RT}{\pi M}} =$	$\sqrt{\frac{8\mathrm{PV}}{\pi\mathrm{M}}} = \sqrt{\frac{8\mathrm{P}}{\pi\mathrm{d}}}$
$\overline{C} = 0.9213 \times$	RMS velocity

Most Probable Velocity (C_p)

The velocity that is possessed by the maximum number of molecules present in the gas at any temperature is known as **the most prabable velocity.** It is represented by C_p

$$C_{p} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2PV}{M}} = \sqrt{\frac{2P}{d}}$$
$$C_{p} = 0.8166 \times RMS \text{ velocity}$$

For a gas at any temperature the order of velocity $C_p < \overline{C} < C$

At any temperature, order of no. of molecules.

$$n_{C_P} > n_{\overline{C}} > n_C$$

As the velocities of molecules increases, a fraction of the molecules having a particular velocity increases, reaches a maximum value and then decreases.

At high temperature, the fraction of the total molecules possessing low velocities decreases and the fraction of total molecule possessing high velocities increases.

Ratio between molecular velocities

$$C_{p}: \overline{C}: C = C_{p}: \overline{C}: C = \sqrt{2}: \sqrt{\frac{8}{\pi}}: \sqrt{3}$$
$$= \sqrt{2}: \sqrt{2.55}: \sqrt{3} = 0.8166: 0.9213: 1$$
$$\boxed{= 1: 1.128: 1.234}$$

TRAIN YOUR BRAIN

Q. Average velocity of CO_2 at the temperature T_1K and maximum possible velocity of CO_2 at the temperature $T_2 K$ is 9×10^4 cm second⁻¹. Calculate the value of T_1 and T_2 .

Ans. Average velocity =
$$\sqrt{\frac{8RT}{\pi M}}$$

and maximum possible velocity = $\sqrt{\frac{2RT}{M}}$

 $= 9 \times 10^4$ cm sec⁻¹ $= 9 \times 10^2$ m/sec

$$\therefore 9 \times 10^{2} = \sqrt{\frac{8 \times 8.314 \times T_{1}}{3.14 \times 44 \times 10^{-3}}}$$

$$9 \times 10^{2} = \sqrt{\frac{2 \times 8.314 \times T_{2}}{44 \times 10^{-3}}}$$
On solving T₁ = 1682.5 K
T₂ = 2143.4 K

BEHAVIOUR OF REAL GASES: DEVIATION FROM IDEAL GAS BEHAVIOUR

Real Gases

In perfect gas, molecules move independently of one another.

Perfect gas has energy only from kinetic energy but not any contribution from potential energy.

In real gases the intermolecular attractions are important. At large separation of molecules, the energy lowering attractions are dominant, but at short distances the energy raising repulsions dominate

The graph plotted between PV v/s P at constant temperatures is a straight line for all ideal gases, there is a significant deviation from ideal behavior.



Fig.: Graph between PV v/s P at constant temperature

For real gases the graph is not a straight line.

For gases like hydrogen and helium as the pressure increases the value of PV also increases.

Deviation from ideal behaviour can also be observed when a graph is plotted between pressure and volume.

Conclusions of the graph

At high pressure, the measured volume is more than the calculated volume.

At low pressures, measured and calculated volume approach each other.



Fig.: *Plot between P vs V to show Deviation from ideal behaviour of real gases*

Reasons for Deviation from Ideal Gas Behaviour

The two assumptions of the kinetic theory of gases do not hold good.

- 1. There is no force of attraction between the molecules of a gas
- 2. Volume of the molecules of a gas is negligibly small in comparison to the space occupied by the gas

However as temperature decreases and pressure increases, gases start to show deviations.

The Van der Waal's Equation of State

For the gas molecules the available volume for free travel is not equal to the volume of the container.

Volume correction is = V-nb

here b is proportionality constant

At low pressure, the volume is large compared with the volume excluded by the molecules i.e.

$$V >> nb \approx V$$

When two molecules, each of radius 'r' and volume (V_{mol}) = 4/3. π r² approach each other **excluded volume is = 8V_{mol}** for 2 molecules and 4V per molecule.

or

The effect of attractive interactions between molecules is to reduce the pressure that the gas exerts.

Attraction experienced by a molecule is proportional to the concentration (n/V) of molecules in the container.

$$\therefore \quad \text{Reduction in pressure } \propto \left(\frac{n}{V}\right)^2$$

 $\therefore \quad \text{Reduction pressure} = \frac{\text{an}^2}{\text{V}^2}$

where 'a' is the proportionality constant.

 \therefore van der Waals equation of state is

for 'n' moles .

$$\left[\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT\right] \quad \text{or} \quad \left[P = \frac{nRT}{V - nb} - a\left(\frac{n^2}{V^2}\right)\right]$$

for 1 mole

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

At low pressures Vander Waal's equation is

 $PV = RT - \frac{a}{V}$

🖉 KEY NOTE

Van der waal's equation of state

- At extremly low pressure PV = RT
- At high pressures PV = RT + Pb
- At high temperature PV = RT

Units of van der waal's constants

- 'a' and 'b' depend on the nature of the gas and independent of temperature.
- Units of 'a' are atm. lit². mol⁻² or N.m⁴. mol⁻².
- Units of 'b' are lit. mol⁻¹ or m³. mol⁻¹.

Significance of Van der Waal's constants

- The value of 'a' is direct measure of magnitude of attractive forces between the molecules
- Greater the value of 'a' more easily the gas is liquified.
- The value of 'b' gives an idea about the effective size of gas molecules
- Greater the value of 'b' larger the size and smaller is compressible volume
- Constant value of 'b' for over a long range of T and P indicates the molecules are incompressible.

Table : Vander Waal's constants for some common gases

Gas	(l ² atm mol ⁻²)	(l mol ⁻¹)
NH ₃	4.17	0.0371
CO ₂	3.59	0.0427
СО	1.49	0.0399
Cl ₂	6.49	0.0562

Gas	(l ² atm mol ⁻²)	(l mol ⁻¹)
H ₂	0.244	0.0266
HCl	1.36	0.0318
SO ₂	6.71	0.0564
Не	0.034	0.0237

The Compression Factor (Z)

Compression factor is the ratio of the actual molar volume of a gas to the molar volume of a perfect gas under the same conditions.

Compression factor (Z)

$$Z = \frac{\text{Molar volume of the gas}(V_m)}{\text{molar volume of perfect gas}(V_m \text{perfect})}$$

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

The molar volume of a perfect gas $(V_m \text{ perfect}) = \frac{RT}{P}$

 $Z = \frac{P.V_m}{RT}$

For ideal gas; Z = 1 at all pressures.

If Z > 1, called as positive deviation from ideal behaviour. That is molar volume of the gas is greater than that expected for a perfect gas.

(a) At low pressures some gases have Z < 1

Ex: methane, ethane, ammonia

(b) At high pressure almost all gases have Z>1 due to repulsive forces between electron clouds of molecules of gas

If Z < 1, called as negative deviation from ideal behaviour which is due to attractive forces between the molecules. That is molar volume of the gas is lower than that expected for a perfect gas

At intermediate pressure gases have Z < 1.

Ideal behaviour depends on nature of the gas and its temperature.

The temperature at which a real gas obeys ideal gas laws over a wide range of pressure is called **"Boyle's temperature"** or **Boyle's**

point (T_b).
$$T_b = \frac{a}{bR}$$

Boyle point depends on nature of the gas.

Above Boyle's temperature real gases shows positive deviations from ideality and Z>1.

Below Boyle's temperature real gases first show Z<1 with increase of pressure and reaches a minimum. On further increase of pressure 'Z' value continuously increases.

At low pressure and high temperature gases shows ideal behaviour.

LIQUEFACTION OF GASES

A gas liquifies if it is cooled below its boiling point at a given pressure.

Gases which have very low b.p are liquified by using Joule-Thomson effect.

Ideal gases do not show Joule thomson effect as there are no intermolecular forces between their molecules.

Critical Constant

Critical temperature of a gas is the highest temperature at which liquefaction of gas first occurs.

At critical temperature liquid phase passes into gaseous state imperceptibly and continuously.

A gas below the critical temperature can be liquefied by applying pressure and is called vapour of the substance.

At critical temperature, density of vapour phase equal to density of liquid phase.

Pressure required to liquefy a gas at its critical temperature is called **critical pressure** (P_c).

The volume occupied by 1 mole of the gas at critical temperature and critical pressure is critical volume (V_c).

Relationship between critcal constants and Van der Waal constants

•
$$P_{\rm C} = \frac{a}{27b^2}; V_{\rm C} = 3b; T_{\rm C} = \frac{8a}{27Rb}$$

- Reduced pressure is P/P_c
- Reduced volume is V/V_c
- Reduced temperature T/T_c

LIQUID STATE

Liquids, possess strong intermolecular forces as compared to gases

- less compressible
- more denser
- has definite volume
- can flow like gases

Vapour Pressure

The pressure exerted by vapours when there is an equilibrium state between the liquid phase and vapour phase is called equilibrium vapour pressure or **saturated vapour** pressure.

At 1 atmosphere pressure boiling temperature is called normal boiling point and at 1 bar pressure it is called **standard boiling point of the liquid.**

Standard boiling point of a liquid is slightly less than the normal boiling point since 1 bar is less than 1 atm.

Example: normal b.p. of water is 373K and standard b.p. of water is 372.6K

At high altitude as the atmospheric pressure is less, the liquids boil at low temperature.

In autoclaves water is boiled under high pressure to sterilise instruments.

In a closed vessel a liquid on heating doesn't boil but its vapour pressure increases and at critical temperature, density of liquid and density of vapour is going to be equal.

The vapaur pressure of a given liquid at two different temperature can be compared using Clausius Clapeyron equation.

$$\log \frac{P_2}{P_1} = \frac{\Delta H}{2.303 R} \left[\frac{T_2 - T_1}{T_1 T_2} \right]$$

where P_1 , P_2 vapour pressure and T_1 , T_2 are temperature, ΔH is heat of vapourisation.

Surface Tension (σ)

The force acting along the surface of a liquid at right angles to any line of 1 unit length.

The energy required to increase the surface area of the liquid by 1 unit is called surface energy.

Units for surface energy is Jm⁻²

Units for surface tension is kgs⁻² and SI unit – Nm⁻¹.

The liquid has lowest energy when the surface area is minimum . Hence liquid droplets exist in spherical shape.

Surface tension decreases with increase of temperature because of increase in kinetic energy of molecules and decrease in intermolecular forces.

🖉 KEY NOTE

• At critical temperature, surface tension becomes zero because meniscus between liquid and vapour disappears.

Viscosity

Strong intermolecular forces between the molecules of successive layers of liquid holding them together show resistance to the flow and create friction between the layers of fluid. The measure of this resistance to the flow of liquid is **viscosity**.

The regular gradation of velocity for layers in passing from one layer to the next layer is called "Laminar flow".

A force (F) is required to maintain the flow of layers is proportional

to the area (A) of contact and velocity gradient $\left(\frac{dv}{dx}\right)$.

$$F \propto A, F \propto \frac{dv}{dx}$$
$$F = \eta A. \frac{dv}{dx}$$

here η = proportionality constant or coefficient of viscosity

Viscosity coefficient is defined as the force when velocity gradient and area of contact each is unity.

 η is a measure of viscosity

SI unit for η is Nm⁻² s or Pas. C.G.S unit for η is "poise".

$$= 10^{-1} \text{ kgm}^{-1} \text{ s}^{-1}$$

As viscosity increases, liquids flow slowly.

Hydrogen bond and van der Waals forces cause high viscosity.

Glass is an extremely viscous liquid and its properties resemble to solids.

As the temperature increases, viscosity decreases since kinetic energy of molecules that overcome the intermolecular forces.

Topicwise Questions

THE GAS LAWS

1. A gas has a vapour density 11.2. The volume occupied by 1 g of the gas at NTP is:

(u) I L (b) II.2 I
$(u) \ 1 \ L \qquad (b) \ 11.2$

- (c) 22.4 L (d) 4 L
- **2.** If V_0 is the volume of a given mass of gas at 273 K at a constant pressure then according to Charle's law, the volume at 10°C will be:

(<i>a</i>) 11V ₀	(b) $\frac{1}{273}(V_0 + 10)$
(c) $V_0 + \frac{10}{273}$	(d) $\frac{283}{273}$ V ₀

- **3.** At what temperature, the volume of a gas would become zero?
 - (*a*) 0°C (*b*) 273°C (*b*) 273°C (*b*) 273°C
 - (c) -273.15°C (d) 273 K
- 4. At a constant pressure, what should be the percentage increase in the temperature in kelvin for a 10% increase in the volume:

(a)	10%	(b)	20%
(c)	5%	(<i>d</i>)	50%

- 5. According to charle's law, (here k is constant of proportionality)
 - (a) $\left(\frac{dV}{dT}\right)_{p} = K$ (b) $\left(\frac{dV}{dT}\right)_{p} = -K$ (c) $\left(\frac{dV}{dT}\right)_{p} = -\frac{K}{T}$ (d) None of these
- **6.** If the absolute temperature of a gas having volume Vcm³ is doubled and the pressure is reduced to half, the final volume will be:

(a) 0.25 V	(b) 0.050 V
(c) 2 V	(<i>d</i>) 4 V

7. An open vessel containing air is heated from 27°C to 127°C. The fraction of air originally present which goes out of it is:

<i>(a)</i>	3/4	(<i>b</i>) 1	./4
<i>(c)</i>	2/3	(d) 1	/8

8. Which of following V v/s T graph represents the behaviour of 1 mole of an ideal gas at ONE atm P?





9. A balloon is filled with H_2 at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure, the gas occupies 2.27 L volume, upto what volume can the balloon be expanded–

(a) 12.3 L	(b) 11.35 L
(c) 20.5 L	(<i>d</i>) 6 L

- **10.** The volume of a gas at 0°C is 273 mL. The volume of the gas at 27°C and at same pressure would be
 - (*a*) 573 mL (*b*) 300 mL
 - (c) 546 mL (d) 327 mL

IDEAL GAS EQUATION

- **11.** The molar volume of CO_2 is maximum at:
 - (a) NTP (b) 0° C and 2 atm
 - (*c*) 127°C and 1 atm (*d*) 273°C and 2 atm
- **12.** 0.30 g of gas was found to occupy a volume of 82.0 ml at 27°C and 3 atm pressure. The molecular mass of a gas is:
 - (*a*) 60 g (*b*) 30 g
 - (c) 90 g (d) Unpredictable
- 13. When the universal gas constant (R) is divided by Avogadro's number (N_0). Their ratio is called:
 - (a) Planck's constant (b) Rydberg's constant
 - (c) Boltzmann constant (d) van der waal's constant
- 14. When the pressure of 5 L of N₂ is doubled and its temperature is raised from 300 K to 600 K, the final volume of a gas would be:
 - (a) 10 L (b) 5 L (c) 15 L (d) 20 L
- 15. The weight of 350 ml of a diatomic gas at 0°C and 2 atm pressures is 1 gm. The weight of one atom in gram is: (N = Avogadro's number)

(c) 16 N (d) 32 N

16. Which of the following does not represent ideal gas equation?

(a) $PV = \frac{1}{2}mNu^2$	(b) $PV = nRT$
(c) $P = \frac{nRT}{V}$	(d) $PV = RT$

- **17.** Under what condition will a pure sample of ideal gas not only exhibit a pressure of 1 atm, but also a concentration of 1 mol/ litre:?
 - (a) At STP
 - (*b*) When V = 22.4 L
 - (c) When T = 12K
 - (d) Impossible under any condition
- **18.** For n moles of ideal gas, the equation of state may be written as:

(a)
$$\frac{PT}{n} = RV$$
 (b) $PV = (RT)^2$
(c) $P/T = nR/V$ (d) $PV = \frac{RT}{n}$

19. If 16 g of CH₄ and 4 g of H₂ are mixed and kept at 760 mm Hg pressure at 0°C, then volume occupied by the mixture would be

(a) 22.4 L	(b) 44.8 I
(c) 67.2 L	(<i>d</i>) 33.6 I

20. An open vessel at 27°C is heated until 1/5th part of the air in it has been expelled. If the volume of the container remains constant, what should be the temperature to which the vessel has been heated?

(a) 375 K	(<i>b</i>) 750 K
(c) 546 K	(<i>d</i>) 408 K

21. If the density of a gas A is 1.5 times that of B, then molecular mass of A is M. The molecular mass of B will be:

(a) 1.5M	(b) $\frac{M}{1.5}$
(c) 3M	$(d) \frac{M}{2}$

22. The approximate temperature at which 1 mol L^{-1} of a sample of pure ideal gas exhibits a pressure of 101.325 k Pa is:

(a) 12.2 K	(<i>b</i>) 122 K
(c) 244 K	(<i>d</i>) 300 K

- **23.** The numerical value of universal gas constant (R) depends upon:
 - (a) The nature of a gas
 - (*b*) Condition of a gas
 - (c) The units of measurement
 - (*d*) None of above
- 24. Which of the following expression is true regarding gas laws? (W = weight, M = molar mass)

(a)
$$\frac{T_1}{T_2} = \frac{M_1 W_2}{M_2 W_1}$$
 (b) $\frac{T_1}{T_2} = \frac{M_2 W_1}{M_1 W_2}$

(c)
$$\frac{T_1}{T_2} = \frac{M_1 W_1}{M_2 W_2}$$
 (d) $\frac{T_2}{T_1} = \frac{M_1 W_1}{M_2 W_2}$

- **25.** 8.2 L of an ideal gas weighs 9.0 g at 300 K and 1atm pressure. The molecular mass of the gas is:
 - (a) 9 g/mol (b) 27 g/mol
 - (c) 54 g/mol (d) 81 g/mol
- **26.** By the ideal gas law, the pressure of 0.60 moles NH_3 gas in a 3.00 L vessel at 25°C is, (given that R = 0.082 L atm mol⁻¹ K^{-1} :)
 - (a) 48.9 atm (b) 4.89 atm
 - (c) 0.489 atm (d) 489 atm
- 27. The temperature at which 28 g of N_2 will be occupying a volume of 10.0 L at 2.46 atm is:
 - (a) 299.6 K (b) 0°C (c) 273 K (d) 10°C
- 28. Density of carbon monoxide is maximum at:
 - (*a*) 2 at m and 600 K (*b*) 0.5 atm and 273 K
 - (*c*) 6 atm and 1092 K (*d*) 4 atm and 500 K
- **29.** In which one of the following does the given amount of chlorine gas exert the least pressure in a vessel of capacity 1 dm³ at 273 K?
 - (a) 8.0355 g (b) 0.071 g

(c) 6.023×10^{21} molecules (d) 0.02 mol

DALTON'S LAW OF PARTIAL PRESSURE

30. The density of a gas A is twice that of B Molecular mass of A is half of molecular mass of B. The ratio of partial pressure of A and B is:

(<i>a</i>) 1/4	(<i>b</i>) 1/2
(c) 4/1	(<i>d</i>) 2/1

31. Equal weight of CH_4 and H_2 are mixed in a container at 25°C. Fraction of total pressure exerted by hydrogen is:

(<i>a</i>) 1/2	<i>(b)</i> 1/3
(c) 1/9	(<i>d</i>) 8/9

GRAHAM'S LAW OF DIFFUSION

32. 2 g of hydrogen diffuse from a container in 10 min. How many gram of oxygen would diffuse through the same container in the same time under similar conditions?

(<i>a</i>) 0.5 g	(b) 4 g
(<i>c</i>) 6 g	(<i>d</i>) 8 g

33. If the ratio of the rates of diffusion of two gases A and B is 4 : 1, then the ratio of their densities is:

(<i>a</i>) 16 : 1	(<i>b</i>) 4 : 1
(<i>c</i>) 1 : 4	(d) 1:16

34. A certain gas takes three times as long to effuse out as helium. Its molecular mass will be:

(<i>a</i>) 27 u	(<i>b</i>) 36 u

35. Identify the pair of gases that have equal rate of diffusion:

(<i>a</i>) NO, CO	(<i>b</i>) N ₂ O, CO
$(c) N_2O, CO_2$	$(d) \operatorname{CO}_2, \operatorname{NO}_2$

KINETIC MOLECULAR THEORY OF GASES

36. The value of a gas constant R is 8.314 X. Here; X represents: (*a*) Litre atm. $K^{-1} \mod^{-1}$ (*b*) Cal $\mod^{-1} K^{-1}$

	(b) cur mor it
c) JK ⁻¹ mol ⁻¹	(<i>d</i>) None of above

37. What is the ratio of kinetic energies of 3 g of hydrogen and 4 g of oxygen at T(K)?

(<i>a</i>) 12 : 1	(<i>b</i>) 6 : 1
(<i>c</i>) 1 : 6	(<i>d</i>) 24 : 1

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38. The root mean square velocity of ideal gas at constant pressure varies with density (*d*) as:

(<i>a</i>) d^2	(<i>b</i>) d
(c) \sqrt{d}	(d) $\frac{1}{\sqrt{d}}$

39. At what temperature will the RMS velocity of SO_2 be the same as that of O_2 at 303 K?

(<i>a</i>) 403K	(<i>b</i>) 303K
(c) 606K	(<i>d</i>) 273K

40. The critical temperatures of O₂, N₂, H₂ and CO₂ are 154.3 K, 126 K, 33.2 K and 304 K. The extent of adsorption on tungsten may be highest in case of:

(<i>a</i>) H ₂	(<i>b</i>) N ₂
$(c) O_2$	$(d) \operatorname{CO}_2$

- **41.** The RMS velocity of molecules of a gas of density 4 kg m⁻³ and pressure 1.2×10^5 Nm⁻² is:
 - (a) 900 m/sec (b) 120 m/sec

(c) 600 m/s (d) 300 m/sec

42. The molecules of which of the following gases have the largest mean average K.E. at 250°C?

(a) Al	(b) CO
--------	--------

- (c) He (d) All have same average K.E.
- **43.** By what factor does the average velocity of gaseous molecule increases, when the temperature (in kelvin) is doubled?

(c) 2.8 (d) 4.0

DEVIATION FROM IDEAL GAS BEHAVIOR

44. Maximum deviation from ideal gas is expected from:

(<i>a</i>) $H_2(g)$	(<i>b</i>) N ₂
(c) CH_4	(d) $NH_3(g)$

45. The, compressibility factor (Z) of one mole of a van der waal's gas of negligible 'a' value is:

(a) 1
(b)
$$\frac{bP}{RT}$$

(c) $1 + \frac{bP}{RT}$
(d) $1 - \frac{bP}{RT}$

46. The compressibility factor for an ideal gas is:

(a) = 1	(<i>b</i>) <1
(c) > 1	$(d) \infty$

LIQUID STATE

- 47. On increasing pressure, melting point of ice:
 - (a) Decreases
 - (b) Increases
 - (c) Remain unchanged
 - (d) Change in some manner
- 48. Which of the following is correct statement?
 - (a) Surface tension of a liquid decreases with increases in temperature
 - (b) Vapour pressure of a liquid decrease with increase in temperature
 - (c) Viscosity of a liquid decrease with decrease in temperature
 - (*d*) In gravity free environment, droplets of a liquid on flat surface are slightly flattened
- **49.** When a capillary tube of diameter 0.8 mm is dipped in a liquid having density 800 kg m⁻³, then the height of liquid in the capillary tube rises to 4 cm. The surface tension of liquid is ($g = 9.8 \text{ m/s}^2$):

(a) $4.3 \times 10^{-2} \text{ Nm}^{-1}$	(b) $5.6 \times 10^{-2} \mathrm{Nm^{-1}}$
(c) $6.3 \times 10^{-2} \text{ Nm}^{-1}$	(d) $7.3 \times 10^{-2} \mathrm{Nm^{-1}}$

Learning Plus

- 1. A person living in Shimla observed that cooking food without using pressure cooker takes more time. The reason for this observation is that at high altitude:
 - (a) Pressure increases (b) Temperature decreases
 - (c) Pressure decreases (d) Temperature increases
- **2.** Which of the following property of water can be used to explain the spherical shape of rain droplets?
 - (a) Viscosity (b) Surface tension
 - (c) Critical phenomena (d) Pressure
- **3.** A plot of volume (V) versus temperature (T) for a gas at constant pressure is a straight line passing through the origin. The plots at different values of pressure are shown in figure.

Which of the following order of pressure is correct for this gas?



- **4.** The interaction energy of London force is inversely proportional to sixth power of the distance between two interacting particles but their magnitude depends upon:
 - (a) Charge of interacting particles
 - (b) Mass of interacting particles
 - (c) Polarisability of interacting particles
 - (d) Strength of permanent dipoles in the particles
- **5.** Dipole-dipole forces act between the molecules possessing permanent dipole. Ends of dipoles possess 'partial charges. The partial charge is:
 - (a) More than unit electronic charge
 - (*b*) Equal to unit electronic charge
 - (c) Less than unit electronic charge
 - (d) Double the unit electronic charge
- **6.** The pressure of a 1 : 4 mixture of dihydrogen and dioxygen enclosed in a vessel is one atmosphere. What would be the partial pressure of dioxygen?
 - (a) 0.8×10^5 atm
 - (b) 0.008 Nm^{-2}
 - (c) $8 \times 10^4 \text{ Nm}^{-2}$
 - (d) 0.25 atm
- **7.** As the temperature increases, average kinetic energy of molecules increases. What would be the effect of increase of temperature on pressure provided the volume is constant?
 - (a) Increases
 - (b) Decreases
 - (c) Remains same
 - (d) Becomes half
- **8.** Gases possess characteristic critical temperature which depends upon the magnitude of intermolecular forces between the particles. Following are the critical temperatures of some gases:

Gases	H ₂	He	0 ₂	N ₂
Critical temperature in Kelvin	33.2	5.3	154.3	126

From the above data what would be the order of liquefaction of these gases? Start writing the order from the gas liquefying first:

(a) H₂, He, O₂, N₂

- (*b*) He, O₂, H₂, N₂
- (c) N₂, O₂, He, H₂
- $(d) O_2, N_2, H_2, He$
- **9.** What is SI unit of viscosity coefficient (η) ?

(a) Pascal	$(b) \text{ Nsm}^{-2}$
(c) $\rm km^{-2} s$	$(d) \text{ Nm}^{-2}$

10. Atmospheric pressures recorded in different cities are as follows:

Cities	Shimla	Bangalore	Delhi	Mumbai
$P \text{ in } N/m^2$	1.01×10^5	1.2×10^5	1.02×10^5	1.21×10^5

Consider the above Data And Mark The Place At Which Liquid Will Boil First:

(a) Shimla	(b) Bangalore
(c) Delhi	(d) Mumbai

11. Which curve in figure represents the curve of ideal gas?



(a) Only B	(b) C and D
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- (c) E and F (d) A and B
- **12.** Increase in kinetic energy can overcome intermolecular forces of attraction. How will the viscosity of liquid be affected by the increase in temperature?
 - (a) Increase (b) No effect
 - (c) Decrease
 - (d) No regular pattern will be followed
- **13.** How does the surface tension of a liquid vary with increase in temperature?
 - (a) Remains same (b) Decreases
 - (c) Increases
 - (d) No regular pattern is followed
- 14. If both oxygen and helium gases are at the same temperature, the rate of diffusion of O_2 is very close to:
 - (a) 4 times that of He (b) 2 times that of He
 - (c) 0.35 times that of He (d) 8 times that of He
- **15.** The ratio of rates of diffusion of helium and methane under same conditions of temperature and pressure will be:

(a) 4:1	(b) 2:1
(c) 1:1	(d) 1:2

16. Average K.E. of N_2 at 27°C is E. The average K.E. of CO_2 at the same temperature will be:

(<i>a</i>) E	(<i>b</i>) 22 E
(c) $\frac{\mathrm{E}}{22}$	(d) $\frac{\mathrm{E}}{\sqrt{2}}$

- 17. The most probable velocity of a gas molecule at 298 K is 300 m/sec. Its RMS velocity (in m/s) is:
 - (*a*) 420 (*b*) 245
 - (c) 402 (d) 367
- **18.** The ratio of root mean square speed, average speed and most probable speed is
 - (a) $3\pi: 8: 2\pi$ (b) $\pi: 2\pi: 3$ (c) $2\pi: 3\pi: 8$ (d) $\pi: 2\pi: 6$
- **19.** The temperature at which real gases obey the ideal gas laws over a wide range of pressure is called:
 - (a) Critical temperature(b) Inversion temperature(c) Boyle temperature(d) Reduced temperature
- **20.** At relatively, high pressure, Van der waal's equation reduces
- to:

(a)
$$PV = RT$$
 (b) $PV = RT - a/v$

(c)
$$PV = RT - \frac{a}{V^2}$$
 (d) $PV = RT + Pb$

- **21.** In Vander waal's equation of state of the gas, the constant 'b' is a measure of:
 - (a) Intermolecular repulsions
 - (b) Intermolecular attractions
 - (c) Volume occupied by the molecules
 - (d) Intermolecular collisions per unit volume
- **22.** The real gas most closely approaches the behaviour of an ideal gas at:
 - (*a*) 15 atm and 200 K (*b*) 1 atm and 273 K
 - (c) 0.5 atm and 500 K (d) 15 atm and 500 K
- 23. A gas can be liquefied:
 - (a) At its critical temperature
 - (b) Above its critical temperature
 - (c) Below its critical temperature
 - (*d*) At 0°C
- 24. Mark out the wrong expression:
 - (a) Boyle's temperature $T_b = \frac{Ra}{b}$
 - (b) Critical pressure $P_c = \frac{a}{27b^2}$
 - (c) Critical temperature $T_c = \frac{8a}{27Rb}$
 - (d) Critical volume $V_c = 3b$
- **25.** In Van der Waal's equation for the real gas, the expression for the net force of attraction amongst the gas molecules is given by

(a) $\frac{\mathrm{an}^2}{\mathrm{V}^2}$ (b) $\mathrm{p} + \frac{\mathrm{an}^2}{\mathrm{V}^2}$

(c)
$$p - \frac{an^2}{V^2}$$
 (d) $\frac{V^2}{an^2}$

- **26.** The surface tension of which of the following liquid is maximum?
 - (a) H_2O (b) C_6H_6 (c) CH_3OH (d) C_2H_5OH
- **27.** Out of the four liquids given below, the one having lowest vapour pressure at 25°C is:
 - (a) Carbon tetrachloride (b) Benzene
 - (c) Chloroform (d) Water
- **28.** During the evaporation of liquid:
 - (a) The temperature of the liquid will rise
 - (b) The temperature of the liquid will fall
 - (c) May be rise or fall depending on the nature
 - (d) The temperature remains unaffected
- **29.** The following is a method to determine the surface tension of liquids:
 - (a) Single capillary method
 - (b) Refractometer method
 - (*c*) Polarimetry method
 - (d) Boiling point method
- **30.** Which is distilled first?
 - (a) Liquid CO_2 (b) Liquid N_2 (c) Liquid O_2 (d) Liquid H_2
- **Multiconcept MCQs**
- 1. A vessel contains H₂ & O₂ in the molar ratio 8 : 1 respectively. This mixture of gases is allowed to diffuse through a hole, find the composition of the mixture coming out of the hole:

(a) 8:1	(<i>b</i>) 16 : 1
<i>(a)</i> 0.1	(b) 10.1

(c) 32 : 1	(d) 64:1
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2. Potassium Hydroxide solutions are used to absorb CO_2 . How many litres of CO_2 at 1 atm and 22°C would be absorbed by an aqueous solution containing 15 g of KOH?

 $\begin{array}{ll} 2 \text{KOH} + \text{CO}_2 \rightarrow \text{K}_2 \text{CO}_3 + \text{H}_2 \text{O} \\ (a) \ 3.24 \text{ L} & (b) \ 1.62 \text{ L} \\ (c) \ 6.48 \text{ L} & (d) \ 0.324 \text{ L} \end{array}$

3. Two glass bulbs A (of 100 ml volume) and B (of 150 ml volume) containing same gas are connected by a small tube of negligible volume. At particular temperature the pressure in A was found to be 20 times more than that in bulb (*b*) The stopcock is opened without changing temperature. The percentage decreases in pressure will be:

(<i>a</i>) 57%	<i>(b)</i> 22%
(c) 23%	(<i>d</i>) 45%

4. A gas obeys the equation P (V–b) = RT, then which statement is/are correct:

(a) Isochoric curves have slope
$$=\frac{K}{V-h}$$

- (b) Isobaric curves have slope $=\frac{R}{P}$ and intercept b
- (c) Isochoric curves have slope $=\frac{PV}{h}$
- (d) Both (a) and (b)
- **5.** An open container of volume V contains air at temperature 300 K. The container is heated to such a temperature so that amount of gas coming out is 2/3 of the amount of gas initially present in the container. Find the temperature to which the container should be heated:

<i>(a)</i>	300 K	<i>(b)</i>	600 K
(<i>c</i>)	900 K	(d)	1200 K

6. For two gases, A and B with molecular weights M_A and M_B, it is observed that at a certain temperature T, the mean velocity

of A is equal to the root mean square velocity of (b) Thus the mean velocity of A can be made equal to the mean velocity of B, if:

- (a) A is at temperature T_1 and B at T_2 and $T_1 > T_2$
- (b) A is lowered to a temperature $T_2 < T$ while B is at T
- (c) Both A and B are lowered in temperature
- (d) Both A and B are raised in temperature
- 7. Two inflated Balloons I and II having volume 600 ml and 1500 ml at 300 K are taken as shown in diagram-



If the maximum volume of inner and outer balloons are 800 ml and 1800 ml respectively then the balloon which will burst first on gradual heating is:

- (a) Outer balloon
- (b) Inner balloon
- (c) None of them will ever burst
- (d) Both will burst simultaneously
- **8.** An unspecified quantity of an ideal gas has a volume of 30 litres at 20°C, the gas was compressed until the pressure has doubled at constant temperature & then, the temperature is raised to 100°C while the pressure is kept constant, what is the final volume the gas–

(<i>a</i>) 25 litre	(<i>b</i>) 30 L
(c) 19 L	(<i>d</i>) 45 L

- 9. The rms velocity of hydrogen is $\sqrt{7}$ times the rms velocity of nitrogen. If T is the temperature of the gas, then:
 - (a) $T_{(H_2)} = T_{(N_2)}$
 - (b) $T_{(H_2)} > T_{(N_2)}$
 - (c) $T_{(H_2)} < T_{(N_2)}$
 - (d) $T_{(H_2)} = \sqrt{7} T_{(N_2)}$
- **10.** At constant P of 0.0821 atm log V vs log T graph is plotted for 3 samples of ideal gas as shown. Value of moles of gas in these three samples is respectively:



(a)
$$3, 1, \frac{1}{2}$$
 (b) $\frac{1}{2}, 1, 3$

(c)
$$1, \frac{1}{2}, 3$$
 (d) $\frac{1}{2}, 3, 1$

11. Statement-1: According to kinetic theory of gases, volume of an ideal gas molecule and molecular attraction forces between ideal gas molecules are taken to be zero.

Statement-2: According to kinetic theory of gases, average kinetic energy of gas molecules is directly proportional to square root of the absolute temperature of the gas.

- (*a*) Statement-1 is true, statement-2 is true; statement-2 is a correct explanation for statement-1
- (b) Statement-1 is true, statement-2 is true; statement-2 is not a correct explanation for statement-1
- (c) Statement-1 is true, statement-2 is false.
- (d) Statement-1 is false, statement-2 is true.
- **12.** The compressibility factor for 1 mole of a Vander Waals gas at the Boyle's temperature is–

(a)
$$1 + \frac{b^2}{V(V-b)}$$
 (b) $1 - \frac{b}{V}$
(c) $1 + \frac{b}{V}$ (d) $1 - \frac{b^2}{V^2}$

13. A certain mass of gas A is kept in a closed container where it undergoes dimerisation, according to the reaction :

 $2A(G) \rightarrow A_2(G)$

Assuming temperature to remain constant, it was found that the partial pressure of A_2 gas after time t was one-fifth of the initial pressure in container. Then, select the correct statement(s) :

- (*a*) The ratio of initial total pressure to the total pressure at time t is 3 : 2.
- (b) The mole fraction of A_2 in the vessel after time t is 0.2.
- (c) The percentage dimerisation of A upto time t is 40%.
- (d) If each molecule of gas A weighs 10^{-22} g, then the average molecular mass of the mixture in the container after time t is about 75 amu.
- **14.** If the partition is removed the average molar mass of the sample will be (Assume ideal behaviour).

	H ₂	D ₂
	16.42 L	16.42 L
	300 K	300 K
	3 atm	6 atm
(a) $\frac{1}{2}$ g/mol	((b) $\frac{10}{3}$ g/mol
(c) $\frac{3}{2}$ g/mol	((d) $\frac{5}{3}$ g/mol

15. According to Maxwell's distribution of molecular speeds, for the below graph drawn for two different samples of gases A and B at temperature T_1 and T_2 respectively, which of the following statements is/are incorrect:



- (a) If $T_1 = T_2$, then molecular mass of gas $B(M_B)$ is greater than molecular mass of gas $A(M_A)$.
- (b) If molecular mass of gas $A(M_A)$ is equal to molecular mass of gas $B(M_B)$, then $T_1 > T_2$.
- (c) If $T_1 < T_2$, then molecular mass of gas $B(M_B)$ is necessarily less than molecular mass of gas $A(M_A)$.

(*d*) All of these

16. The distribution of molecular kinetic energy at two temperature is as shown in the following graph:



which of the following conclusion is correct?

- A. The number of molecules with energy E_a or greater is proportional to the shaded area for each temperature
- B. The number of molecules with energy E_a or less is proportional to the shaded area for each temperature
- C. The number of molecules with energy Ea is the mean of all temperature
- D. The graph follows the Maxwell-Boltzmann energy distribution law

 $(a) A,C \qquad (b) A,B \qquad (c) A,D \qquad (d) B,C$

17. For a closed container containing 100 mole of an ideal gas fitted with movable, frictionless, weightless piston operating such that pressure of gas remains constant at 8.20 atm, which graph represents correct variation of log V vs. log T where V is in litre and T in kelvin.





18. A gas is collected by downward displacement of water. Select the correct expression for p_{gas} according to figure $(d_{Hg} = 13.6 \text{ g cm}^{-3})$



19. An ideal gas of given mass is heated first in a small vessel (I) and then in a large vessel (II). A plot between p vs T are plotted for (I) and (II). The correct p-T curves in I and II conditions is.



NEET Past 10 Years Questions

 Choose the correct option for graphical representation of Boyle's law, which shows a graph of pressure vs. volume of a gas at different temperatures. (2021)



2. Choose the correct option for the total pressure (in atm.) in a mixture of 4 g O_2 and 2 g H_2 confined in a total volume of one litre at 0°C is: (2021)

[Given R = 0.082 L atm mol⁻¹K⁻¹, T = 273 K]

$(a) \ 2.602 (b) \ 25.$	18
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(<i>c</i>) 26.02	(d)	2.518
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3. A mixture of N₂ and Ar gases in a cylinder contains 7 g of N₂ and 8 g of Ar. If the total pressure of the mixture of the gases in the cylinder is 27 bar, the partial pressure of N₂ is : (2020) [Use atomic masses (in g mol⁻¹) : N = 14, AR = 40]

- 4. The minimum pressure required to compress 600 dm³ of a gas at 1 bar to 150 dm³ at 40°C is: (2020 Covid Re-NEET)
 (a) 0.2 bar
 (b) 1.0 bar
 (c) 2.5 bar
 (d) 4.0 bar
- 5. A gas at 350 K and 15 bar has molar volume 20 percent smaller than that for an ideal gas under the same conditions. The correct option about the gas under the same conditions.

The correct option about the gas and its compressibility factor (Z) is: (2019)

- (a) Z > 1 and attractive forces are dominant
- (b) Z > 1 and repulsive forces are dominant
- (c) Z < 1 and attractive forces are dominant
- (d) Z < 1 and repulsive forces are dominant
- 6. The correction factor 'a' in the ideal gas equation corresponds to (2018)
 - (a) Density of the gas molecules
 - (b) Volume of the gas molecules
 - (c) Forces of attraction between the gas molecules
 - (d) Electric field present between the gas molecules
- 7. Given, van der Waals constant for NH₃, H₂, O₂ and CO₂ are respectively 4.17, 0.244, 1.36 and 3.59, which one of the following gases is most easily liquefied? (2018)
 (a) NH₃ (b) H₂ (c) CO₂ (d) O₂
- 8. A 20 litre container at 400 K contains $CO_2(g)$ at pressure 0.4 atm and an excess of SrO (neglect the volume of solid SrO). The volume of the containers is now decreased by moving the movable piston fitted in the container. The maximum volume of the container, when pressure of CO_2 attains its maximum value, will be: (2017-Delhi)

(Given that: $SrCO_3(s) \implies SrO(s) + CO_2(g) K_p = 1.6 \text{ atm}$)

- (a) 2 litre (b) 5 litre (c) 10 litre (d) 4 litre
- 9. Equal moles of hydrogen and oxygen gases are placed in a container with a pin-hole through which both can escape. What fraction of the oxygen escapes in the time required for one-half of the hydrogen to escape? (2016-1)

(a) 3/8 (b) 1/2 (c) 1/8

10. A gas such as carbon monoxide would be most likely to obey the ideal gas law at: (2015 Re)

(d) 1/4

- (a) Low temperatures and low pressures
- (b) High temperatures and low pressures
- (c) Low temperatures and high pressures
- (d) High temperatures and high pressures
- Dipole-induced dipole interactions are present in which of the following pairs? (2013)
 - (a) H_2O and alcohol (b) Cl_2 and CCl_4
 - (c) HCl and He atoms (d) SiF_4 and He atoms
- 12. Maximum deviation from ideal gas is expected from: (2013)
 - (a) $H_2(g)$ (b) $N_2(g)$ (c) $CH_4(g)$ (d) $NH_3(g)$

- 13. 50 mL of each gas A and of gas B takes 150 and 200 seconds respectively for effusing through a pin hole under the similar conditions. If molecular mass of gas B is 36, the molecular mass of gas A will be: (2012 Pre)
 (a) 96 (b) 128 (c) 20.25 (d) 64
- 14. A certain gas takes three times as long to effuse out as helium.Its molecular mass will be:(2012 Mains)(a) 27 u(b) 36 u(c) 64 u(d) 9 u
- **15.** For real gases van Der Waals equation is written as $\left(P + \frac{an^2}{V^2}\right)$ (V - nb) = n RT

where *a* and *b* are Vander Waals constants. Two sets of gases are

(I) O_2 , CO_2 , H_2 and He

He (II) CH_4 , O_2 and H_2

The gases given in set-I in increasing order of b and gases given in set-II in decreasing order of a, are arranged below. Select the correct order from the following: (2012 Mains)

 $\begin{array}{l} (a) \ ({\rm I}) \ {\rm He} < {\rm H}_2 < {\rm CO}_2 < {\rm O}_2 \ ({\rm II}) \ {\rm CH}_4 > {\rm H}_2 > {\rm O}_2 \\ (b) \ ({\rm I}) \ {\rm O}_2 < {\rm He} < {\rm H}_2 < {\rm CO}_2 \ ({\rm II}) \ {\rm H}_2 > {\rm O}_2 > {\rm CH}_4 \\ (c) \ ({\rm I}) \ {\rm H}_2 < {\rm He} < {\rm O}_2 < {\rm CO}_2 \ ({\rm II}) \ {\rm CH}_4 > {\rm O}_2 > {\rm H}_2 \\ (d) \ ({\rm I}) \ {\rm H}_2 < {\rm O}_2 < {\rm He} < {\rm CO}_2 \ ({\rm II}) \ {\rm O}_2 > {\rm CH}_4 > {\rm H}_2 \\ \end{array}$

ANSWER KEY

				Topicwise	e Questior	IS			
1. (<i>a</i>)	2. (<i>d</i>)	3. (<i>c</i>)	4. (<i>a</i>)	5. (<i>a</i>)	6. (<i>d</i>)	7. (<i>b</i>)	8. (<i>c</i>)	9. (<i>b</i>)	10. (<i>b</i>)
11. (<i>c</i>)	12. (<i>b</i>)	13. (<i>c</i>)	14. (<i>b</i>)	15. (<i>a</i>)	16. (<i>a</i>)	17. (<i>c</i>)	18. (<i>c</i>)	19. (<i>c</i>)	20. (<i>a</i>)
21. (<i>b</i>)	22. (<i>a</i>)	23. (<i>c</i>)	24. (<i>a</i>)	25. (<i>b</i>)	26. (<i>b</i>)	27. (<i>a</i>)	28. (<i>d</i>)	29. (<i>a</i>)	30. (<i>c</i>)
31. (<i>d</i>)	32. (<i>d</i>)	33. (<i>d</i>)	34. (<i>b</i>)	35. (<i>c</i>)	36. (<i>c</i>)	37. (<i>a</i>)	38. (<i>d</i>)	39. (<i>c</i>)	40. (<i>d</i>)
41. (<i>d</i>)	42. (<i>d</i>)	43. (<i>a</i>)	44. (<i>d</i>)	45. (<i>c</i>)	46. (<i>a</i>)	47. (<i>a</i>)	48. (<i>a</i>)	49. (<i>c</i>)	
				Learn	ing Plus				
1. (<i>c</i>)	2. (<i>b</i>)	3. (<i>c</i>)	4. (<i>c</i>)	5. (<i>c</i>)	6. (<i>c</i>)	7. (<i>a</i>)	8. (<i>d</i>)	9. (<i>b</i>)	10. (<i>a</i>)
11. (<i>a</i>)	12. (<i>c</i>)	13. (<i>b</i>)	14. (<i>c</i>)	15. (<i>b</i>)	16. (<i>a</i>)	17. (<i>d</i>)	18. (<i>a</i>)	19. (<i>c</i>)	20. (<i>d</i>)
21. (<i>c</i>)	22. (<i>c</i>)	23. (<i>c</i>)	24. (<i>a</i>)	25. (<i>a</i>)	26. (<i>a</i>)	27. (<i>d</i>)	28. (<i>b</i>)	29. (<i>a</i>)	30. (<i>d</i>)
				Multicon	cept MCC	s			
1. (<i>c</i>)	2. (<i>a</i>)	3. (<i>a</i>)	4. (<i>d</i>)	5. (<i>c</i>)	6. (<i>b</i>)	7. (<i>a</i>)	8. (<i>c</i>)	9. (<i>c</i>)	10. (<i>a</i>)
11. (<i>c</i>)	12. (<i>a</i>)	13. (<i>d</i>)	14. (<i>b</i>)	15. (<i>d</i>)	16. (<i>c</i>)	17. (<i>a</i>)	18. (<i>a</i>)	19. (<i>b</i>)	
			NEE	T Past 10	Years Qu	estions			
1. (<i>c</i>)	2. (<i>b</i>)	3. (<i>b</i>)	4. (<i>d</i>)	5. (<i>c</i>)	6. (<i>c</i>)	7. (<i>a</i>)	8. (<i>b</i>)	9. (<i>c</i>)	10. (<i>b</i>)
11. (<i>c</i>)	12. (<i>d</i>)	13. (<i>c</i>)	14. (<i>b</i>)	15. (<i>c</i>)					

Solutions

Topicwise Questions

1. (*a*) The vapour density of a gas = 11.2

- \therefore Molecular weight = 2 × vapour density
- :. Molecular weight of a gas = $2 \times 11.2 = 22.4$ g
- \therefore 22.4 g of a gas occupies = 22.4 L volume

$$\therefore$$
 1 g of a gas occupied = $\frac{22.4}{22.4} \times 1 = 1L$

2. (d) According to Charle's law

$$V_{t} = V_{o}\left(\frac{273 + t}{273}\right) = V_{o} \times \frac{283}{273}$$

- **3.** (c) It becomes zero at absolute zero, i.e., -273.15° C
- 4. (a) According to Charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}; V_2 = V_1 + 0.1 V$$

$$T_2 = T_1 + \frac{x}{100} T_1$$
or
$$\frac{V_1}{T_1} = \frac{V_1 + 0.1 V}{T_1 + 0.01 x T_1}$$
or
$$\frac{V_1}{T_1} = \frac{V_1 (1 + 0.1)}{T_1 (1 + 0.01 x)}$$
or
$$1 + 0.01 x = 1.1$$
Hence; x = 10%

5. (a) According to charle's law

V = KT on differentiating at constant pressure

$$\left(\frac{dV}{dT}\right)_{P} = K$$

6. (d) Here; $T_2 = 2T_1$, and $P_2 = \frac{P_1V_1}{2} = V$ Final volume $\therefore V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2}$ $V_2 = \frac{P_1V}{T_1} \times \frac{2T_1}{P_1/2} = 4V$ 7. (b) $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Initial fraction $\frac{V_1}{V_2} = 1$ when temp. is 27°C

At 127°C, the new fraction is $\frac{V_1}{V_2} = \frac{300}{400} = \frac{3}{4}$. Air expelled = 1 3 4-3 1

8. (c) Volume of 1 mole of an ideal gas at 273 K and 1 atm is 22.42. Volume at 373 K and 1 atm P will be-

$$V = \frac{RT}{P} = \frac{0.082 \times 373}{1} = 30.6 L$$

9. (b) Apply boyles law.

$$P_1 V_1 = P_2 V_2$$

 $1 \times 2.27 = 0.2 \times V_2$
 $V_2 = 11.35 L$
So, volume at 0.2 bar P = 11.35 L

Since, The balloon bursts at 0.2 bar P, the volume of balloon should be less than 11.35 L.

- **10.** (b) $\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{273}{273} = \frac{V_2}{300} \Rightarrow V_2 = 300 \text{ mL}$
- **11.** (c) The molar volume of CO_2 at 127°C and 1 atm is

$$= \frac{nRT}{P} = \frac{1 \times 0.082 \times 400}{1} = 32.8 L$$

12. (b) PV = nRT PV = $\frac{W}{M}$ RT
 $\therefore \qquad M = \frac{WRT}{PV}$
 $= \frac{(0.30g)(0.0821L atm K^{-1}Mol^{-1})(300 K)}{(3 atm)(82.0 \times 10^{-3} L)}$
= 30 g

13. (c) It is called Boltzmann constant

14. (b)
$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

 $V_2 \text{ or } V_{\text{final}} = \frac{600 \times P_1 \times 5}{2P_1 \times 300} = 5 \text{ L}$

15. (*a*) PV = nRT

Let the atomic weight be A

$$2 \times \frac{350}{1000} = \frac{1}{2A} \times 0.0821 \times 273$$
$$A = \frac{224}{14} = 16$$
Mass of one atom = $\frac{16}{N_A}$ g

- **16.** (*a*) PV = nRT. Hence; $PV = 1/2 \text{ mNu}^2$ does not represent an ideal gas equation.
- **17.** (*c*) According to gas equation.

$$PV = nRT \text{ or } \frac{1}{T} = \frac{nR}{VP}$$
$$\frac{1}{T} = \frac{1 \times 0.082}{1 \times 1} \text{ or } T = \frac{1}{0.082} \approx 12 \text{ K}$$

18. (c) $PV = nRT \therefore \frac{P}{T} = \frac{nR}{V}$ **19.** (c) n = 1 + 2 = 3 mole \therefore Volume of 3 mole gas at NTP = 3 × 22.4 V = 67.2 L

20. (*a*) $n_1 T_1 = n_2 T_2$

$$\Rightarrow 1 \times 300 = \left(1 - \frac{1}{5}\right) \times T_2$$
$$\therefore T_2 = \frac{300 \times 5}{4} = 375 \text{ K}$$

21. (b)
$$PV = nRT$$
 or $P = \frac{n}{V}RT = \frac{dRT}{M}$
 $\frac{d_A}{M_a} = \frac{d_B}{M_B} \Rightarrow \frac{1.5}{1} = \frac{M}{M_B} \Rightarrow M_B = \frac{M}{1.5}$
22. (c) $PV = nPT$

22. (*a*)
$$PV = nRT$$

 T_2

$$\therefore$$
 T = $\frac{PV}{nR} = \frac{(1 \text{ atm}) \times (1L)}{1(\text{mol}) \times 0.082} = 12.2 \text{ K}$

23. (c) The value of gas depends upon unit of measurement.

24. (a) Ideal gas equation, PV = nRT, $T \propto \frac{1}{n}$

Where; n = number of moles =
$$\frac{\text{weight}(W)}{\text{molecular weight}(M)}$$

 $\frac{T_1}{T_1} = \frac{n_2}{m_2} = \frac{W_2 M_1}{m_1}$

$$=\frac{\mathbf{m}_2}{\mathbf{n}_1}=\frac{\mathbf{m}_2\mathbf{m}_1}{\mathbf{M}_2\mathbf{W}_1}$$

25. (*b*) Ideal gas equation is ; PV = nRT

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

∴ M = $\frac{WRT}{PV} = \frac{9.0 \times 0.082 \times 300}{1 \times 8.2} = \frac{9.0 \times 8.2 \times 3}{1 \times 8.2}$
= 27 g/mol
26. (b) PV = nRT or P = $\frac{nRT}{V}$
= $\frac{0.60 \times 0.0821 \times 298}{3} = 4.89$ atm
27. (a) 28 g N₂ = 1 mol of N₂
Applying, PV = nRT
2.46 × 10 = 1 × 0.0821 × T
T = $\frac{2.46 \times 10}{0.0821}$
T = 299.6 K
28. (d) From ideal gas equation ; d = $\frac{PM}{RT}$
∴ M and R are Constant for a particular gas
∴ d ∝ $\frac{P}{T}$, i.e.,
(a) $\frac{P}{T} = \frac{2}{600} = 0.0033$
(b) $\frac{P}{T} = \frac{0.5}{273} = 0.00183$
(c) $\frac{P}{T} = \frac{6}{1092} = 0.0080$
∴ P/T ratio is highest at 4 atm and 500 K.

: Density of CO is maximum at this condition

29. (*a*) Ideal gas equation , PV = nRT

When V and T are same, $P \propto n$

Thus; lower the number of moles of gas, lesser the pressure it will exert. Hence,

(a)
$$n = \frac{\text{weight}}{\text{molecular weight}} = \frac{8.0355}{71}$$

= 0.113 mol
(b) $n = \frac{0.071}{71} = 1 \times 10^{-3} \text{ mol}$
(c) $n = \frac{\text{number of molecular}}{N_A} = \frac{6.012 \times 10^{21}}{6.022 \times 10^{23}}$
= 0.01 mol
(d) $n = 0.02$ mol

Thus; 8.0355 g chlorine will exert the least pressure in the vessel.

$$P = \frac{WRT}{MV} = \frac{dRT}{M}$$
$$\frac{P_A}{P_B} = \frac{d_A}{d_B} \times \frac{M_B}{M_A} = \frac{2}{1} \times \frac{2}{1} = \frac{4}{1}$$

31. (d) $\frac{\text{m}}{16}$ moles of CH₄, $\frac{\text{m}}{2}$ moles of H₂

Total no. of moles =
$$\frac{m}{16} + \frac{m}{2} = \frac{9m}{16}$$
 moles
 $X_{H_2} = \frac{\frac{m}{2}}{\frac{9m}{16}} = \frac{8}{9}, P_{H_2} = \frac{8}{9} \times P_{total}$

32. (d) Rate of diffusion = $r \propto \frac{V}{t} \propto \frac{1}{\sqrt{M}}$

$$\Rightarrow \frac{V_{H_2}}{V_{O_2}} = \frac{n_{H_2}}{n_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{H_2}}}$$

Thus; $\frac{W_{H_2}}{M_{H_2}} \times \frac{M_{O_2}}{W_{H_2}} = \sqrt{\frac{M_{O_2}}{M_{H_2}}}$
 $\frac{\frac{2}{2}}{\frac{W_{32}}{2}} = \sqrt{\frac{32}{2}}, \frac{32}{w} = 4 = 8 \text{ g}$

33. (d) Rate of diffusion $\propto \frac{1}{\sqrt{d}}$ $\frac{r_{A}}{r_{B}} = \sqrt{\frac{d_{B}}{d_{A}}} \Longrightarrow \frac{4}{1} = \sqrt{\frac{d_{B}}{d_{A}}} = \frac{16}{1} = \frac{d_{B}}{d_{A}}$ $\frac{d_{A}}{d_{B}} = \frac{1}{16}$

34. (*b*) From Graham's diffusion law

$$\frac{\mathbf{r}_1}{\mathbf{r}_2} = \sqrt{\frac{\mathbf{M}_2}{\mathbf{M}_1}} \text{ or } \frac{\mathbf{V}_{\text{He}}}{\mathbf{t}_{\text{He}}} \times \frac{\mathbf{t}_2}{\mathbf{V}_2} = \sqrt{\frac{\mathbf{M}_2}{\mathbf{M}_{\text{He}}}}$$

$$\therefore \text{ Volume is same and } \mathbf{t}_2 = 3\mathbf{t} ,$$

$$\frac{3}{1} = \sqrt{\frac{\mathbf{M}_2}{4}} \Rightarrow 9 = \frac{\mathbf{M}_2}{4}$$

$$\mathbf{M}_2 = 9 \times 4 = 36 \text{ u}$$

35. (c)
$$\frac{\mathbf{r}_1}{\mathbf{r}_2} = \sqrt{\frac{\mathbf{M}_2}{\mathbf{M}_1}}$$

Molecular mass of $N_2O = 28 + 16 = 44$ Molecular mass of $CO_2 = 12 + 32 = 44$

$$\frac{\mathbf{r}_{N_2O}}{\mathbf{r}_{CO_2}} = 1$$

 $r_{N_2O} = r_{CO_2}$

36. (c) JK⁻¹ mol⁻¹
37. (a) K.E. =
$$\frac{3}{2}$$
nRT
 $\therefore \frac{(K.E.)_{H_2}}{(K.E.)_{O_2}} = \frac{n_{H_2}}{n_{O_2}}$ (at TK)
 $\frac{3}{2} \times \frac{32}{4} = 12:1$

38. (d) Root mean square velocity

$$\overline{\mathbf{u}} = \sqrt{\frac{3RT}{M}}$$

$$\sqrt{\frac{3PV}{Mn}} = \sqrt{\frac{3PV}{M \times \frac{m}{M}}} \quad \left[\because \frac{PV}{n} = RT \right]$$

$$\sqrt{\frac{3PV}{m}} = \sqrt{\frac{3P}{d}}$$
Thus; at constant pressure $\overline{\mathbf{u}} \propto \frac{1}{\sqrt{d}}$
39. (c) The RMS velocity of SO₂ = $\sqrt{\frac{3RT_{SO_2}}{M_{SO_2}}} = \sqrt{\frac{3RT_{SO_2}}{64}}$
The RMS velocity of O₂ = $\sqrt{\frac{3RT_{O_2}}{M_{O_2}}} = \sqrt{\frac{3R \times 303}{32}}$

$$\frac{3RT_{SO_2}}{64} = \frac{3R \times 303}{32}$$

$$T_{SO_2} = \frac{303 \times 64}{32} = 606 \text{ K}$$

- 40. (d) More easily liquefied gases (i.e., gases having higher critical temperature) are readily adsorbed, Van der waal's forces are stronger near the critical temperatures. So; adsorption of CO₂ will be highest.
- 41. (d) The RMS velocity of a gas $= \sqrt{\frac{3P}{d}}$ Given ; P = 1.2 × 10⁵ Nm⁻² ; d = 4kg m⁻³ \therefore V_{ms} $= \sqrt{\frac{3 \times 1.2 \times 10^5}{4}} = \sqrt{0.9 \times 10^5} = \sqrt{9 \times 10^4}$ $3 \times 10^2 = 300$ m/sec
- **42.** (*d*) K.E. of a gas is independent of mass and depends upon temperature.
- **43.** (*a*) Average velocity ; $V_{av} = \sqrt{\frac{8RT}{\pi M}}$ $V_{av} \propto \sqrt{T}$

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

when temperature is doubled; $T_2 = 2T$,

$$\frac{\mathbf{V}_1}{\mathbf{V}_2} = \sqrt{\frac{\mathbf{I}_1}{2\mathbf{T}_1}}$$
$$\mathbf{V}_2 = \sqrt{2}\mathbf{V}_1 = 1.4\,\mathbf{V}_1$$

- 44. (d) Easily liquefiable gases like NH_3 , SO_2 etc.. exhibit maximum deviation from ideal gas as for them Z<<<1. CH₄ also exhibits deviation but it is less as compared to NH_3 , but permanent gases like H_2 , N_2 cannot liquefied; thus show no deviation.
- **45.** (c) Vander waal's equation for 1 mole of a gas is

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

If a is negligible $P + \frac{a}{V^2} \approx P$ P(V-b) = RT PV - Pb = RT or $\frac{PV}{RT} - \frac{Pb}{RT} = 1$ $\frac{PV}{RT} = 1 + \frac{Pb}{RT}$

- **46.** (*a*) Compressibility factor; $Z = \frac{V_{obs}}{V_{calc}}$ For ideal gas, $V_{ob} = V_{cal}$; So Z = 1
- **47.** (*a*) Water fills a smaller volume when it is liquid rather than solid, it will go to a lower melting point, on increasing pressure, allowing more solid to become liquid.
- **48.** (*a*) Surface tension of a liquid decreases with increase in temperature. As the temperature increase, the kinetic energy of the molecules increase and hence, intermolecular attraction decreases which in turn decreases the surface tension.

49. (c)
$$T = \frac{\text{rhdg}}{2} = \frac{(4 \times 10^{-4})(0.04)(9.8) \times 800}{2}$$
$$(r = 0.4 \text{ mm} \approx 4 \times 10^{-4}\text{m})$$
$$= 4 \times 10^{-4} \times 0.04 \times 400 \times 9.8$$
$$= 4 \times 4 \times 4 \times 98 \times 10^{-5} = 16 \times 4 \times 98 \times 10^{-5}$$
$$= 6.3 \times 10^{-2} \text{ Nm}^{-1}$$

Learning Plus

- 1. (c) Pressure at the top of a mountain [or at high altitude] is low. This suggests that boiling takes place at lower temperature, due to which things take more time to boil. However, in a pressure cooker, pressure is increased and hence, boiling point increases. Thus, things comes to boil in a pressure cooker in a less period of time.
- 2. (b) The spherical shape of rain droplets is due to surface tension. The lowest energy state of a liquid will be when the surface area is minimum. Surface tension tries to decrease the surface area of the liquid to the minimum. The rain droplets are spherical because for a given volume, a sphere has minimum surface area.

3. (*c*) At a particular temperature,

pV = constant

Thus, $p_1V_1 = p_2V_2 = p_3V_3 = p_4V_4$

As
$$V_1 > V_2 > V_3 > V_4$$

Therefore, $P_1 < P_2 < P_3 < P_4$

4. (*c*) London dispersion forces operate only over very short distance. The energy of interaction varies as

1

 $\frac{1}{(\text{distance between two interating particles})^6}$

Large or more complex are the molecules, greater is the magnitude of London forces. This is obviously due to the fact that the large electron clouds are easily distorted or polarised.

Hence, greater the polarisability of the interacting particles, greater is the magnitude, of the interaction energy.

- **5.** (c) Dipole-dipole forces act between the molecules possessing permanent dipole and ends of dipoles possess 'partial charges'. Partial charges present on ends of a dipole are always less than the unit electronic charge.
- **6.** (c) Given that pressure of a 1:4 mixture of H_2 and O_2 enclosed in a vessel is one atmosphere. This suggests that the moles ratio of H_2 and O_2 is 1:4. Thus, partial pressure of dioxygen (O_2) is given as

:. Partial pressure of O_2 = Mole fraction of $O_2 \times \text{total}$ pressure of mixture

$$= \frac{4}{1+4} \times 1 = \frac{4}{5} \times 1 \text{ atm}$$

= 0.8 atm = 0.8 × 10⁵ Nm⁻² = 8 × 10⁴ Nm⁻²

- 7. (*a*) As the temperature increases, average kinetic energy of molecules increases. From Gay-Lussac's law, at constant volume, as the temperature is increased, pressure increases.
- 8. (d) Higher the critical temperature, more easily is the gas liquified. Hence, order of liquefaction starting with the gas liquefying first will be O₂, N₂, H₂, He.
 Note: Critical temperature of a gas may be defined as that temperature above which it cannot be liquified howsoever high pressure may be applied on the gas.
- **9.** (*b*) The SI unit of viscosity coefficient (η) is Nm⁻² s or Nsm⁻².

As we know that,
$$f = \eta A \frac{dv}{dx}$$

where, $f =$ force
 $\eta =$ viscosity coefficient

$$\frac{dv}{dx}$$
 = velocity gradient dx

Substitute SI units of f = N, dx = m, $A = m^2$ and $v = ms^{-1}$ in above equation, to get,

$$\eta = \frac{N \times m}{m^2 \times ms^{-1}} = Nm^{-2}s$$

Hence, the S.I unit of η is Nm⁻²s

- **10.** (*a*) Shimla has the lowest atmospheric pressure among all the four cities. Thus, at Shimla liquid will boil first because lower the atmospheric pressure, lower is the boiling point.
- (a) For ideal gas pV = constant at all pressures. Therefore, only B represents ideal gas.
- 12. (c) With increase of temperature, the kinetic energy of the molecules of liquid increases which can overcome the intermolecular forces. Hence, the liquid starts flowing. In other words the viscosity of a liquid decreases with increase in temperature. It has been found that the decrease is about 2% per degree rise in temperature.
- **13.** (*b*) The surface tension of liquids generally decreases with increase of temperature and becomes zero at the critical temperature. The decrease in surface tension with increase of temperature is obviously due to the fact that with increase of temperature, the kinetic energy of the molecules increases and therefore, the intermolecular attraction decreases.
- 14. (c) According to Graham's law of diffusion.

$$\frac{r_{o_2}}{r_{He}} = \sqrt{\frac{M_{He}}{M_{o_2}}} = \sqrt{\frac{4}{32}} = \frac{2}{5.65}$$

$$\therefore \qquad r_{o_2} = 0.35r_{He}$$

15. (b)
$$\frac{\mathbf{r}_{\text{He}}}{\mathbf{r}_{\text{CH}_4}} = \sqrt{\frac{M_{\text{CH}_4}}{M_{\text{He}}}} = \sqrt{\frac{16}{4}} = \sqrt{\frac{4}{1}} = \frac{2}{1}$$

16. (a) At a constant temperature, K.E. remains unchanged.

17. (d) Most probable velocity =
$$\sqrt{\frac{2}{3}} \times \text{RMS}$$
 velocity
 $300 = \sqrt{\frac{2}{3}} \times \text{RMS}$ velocity
RMS velocity = $\frac{\sqrt{3} \times 300}{\sqrt{2}} = 367$ m/sec
 $\sqrt{2075}$

18. (a)
$$\mathbf{v}_{\text{rms}} : \mathbf{v}_{\text{avg}} : \mathbf{v}_{\text{mp}} \Rightarrow \sqrt{\frac{3\text{RT}}{M}} : \sqrt{\frac{8\text{RT}}{\pi M}} : \sqrt{\frac{2\text{RT}}{M}}$$

19. (c) The temperature at which real gas behaves like ideal gas and obeys the gas laws over a wide range of pressure is called boyle temperature "T_b"

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

Where; "*a*" and "*b*" are van der waal's constant.

20. (*d*) At high pressures, the pressure correction for 1 mole of a gas is negligible.

$$\therefore \frac{a}{V^2} = 0$$

However; the volume correction cannot be neglected. Hence; Vander waal's equation reduces to

$$(P + 0) (V - b) = RT(for 1 mole)$$

 $PV - Pb = RT$
 $PV = RT + Pb$

21. (*c*) Vander waal's equation for one mole of a gas

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

where b is volume correction. It raised due to effect size volume.

- **22.** (*c*) At low pressure and high temperature, real gases behave ideally.
- **23.** (*c*) A gas can be liquefied only when its temperature is below its critical temperature. This is because at critical temperature no distinction exists between the gaseous and liquid state and hence; no separate phase is formed.

24. (a) Boyle's temperature;
$$T_b = \frac{a}{Rb}$$

Boyle's temperature (T_b) is the temperature at which a real gas tend to behave ideally over a wide range of pressure.

$$T_{b} = \frac{a}{R.b} = \left[\frac{\partial(PV)}{\partial P}\right]_{\text{limit}_{P} \to 0}$$

At this temperature the effect of size of molecular and intermolecular forces roughly compensate each other.

25. (a) $\frac{an^2}{V^2}$ is pressure correction.

26. (a) Surface tension of H_2O is maximum due to maximum hydrogen bonding in comparison to C_6H_6 , CH_3OH , C_2H_5OH

(Benzene does not form H-bonding).

- **27.** (*d*) The vapour pressure of water is least because of strong intermolecular H-bonding in water molecules.
- 28. (b) During evaporation, molecules beings high energy leave the surface of the liquid. As a result, average K.E. of liquid decreases K.E. ∝ T

: Temperature of liquid falls.

- **29.** (*a*) Surface tension is determined by single capillary method.
- **30.** (d) H_2 is first distilled as its has minimum boiling point.

Multiconcept MCQs

- 1. (c) $\frac{r_{H_2}}{r_{O_2}} = \frac{n_{H_2}}{n_{O_2}} \sqrt{\frac{M_{O_2}}{M_{H_2}}}$ $\frac{r_{H_2}}{r_{O_2}} = \frac{8}{1} \sqrt{\frac{32}{2}} = 32:1$
- 2. (a) Reaction given is: 2KOH + CO₂ \rightarrow K₂CO₃ + H₂O Number of moles of KOH = $\frac{15}{56}$ moles Number of moles of CO₂ required to absorb $=\frac{15}{56}$ moles of KOH $=\frac{1}{2} \times \frac{15}{56} = \frac{15}{112}$ Apply, PV = nRT (for CO_2) $1 \times V = \frac{15}{112} \times 0.0821 \times 295 = 3.24 L$ **3.** (*a*) Before opening the stopcock: $P_A \times 100 = n_A \times R \times T$...(1) $P_{B} \times 150 = n_{B} \times R \times T$...(2) After opening the stop $P \times 250 = (n_A + n_B) \times R \times T$ $250 \,\mathrm{P} = \left(\frac{100 \mathrm{P}_{\mathrm{A}}}{\mathrm{RT}} + \frac{150 \mathrm{P}_{\mathrm{B}}}{\mathrm{RT}}\right) \times \mathrm{RT}$ $250 P = 100 P_A + 150 P_B$ Given : $P_A = 20 P_B$ $250 P = 100 P_A + 150 \times \frac{P_A}{20}$

$$P = \frac{107.5}{250} P_A = 0.43 P_A$$

Percentage decrease in pressure $=\frac{P_A - 0.43P_A}{P_A} \times 100$ = 57%

4. (d) Isochoric curve is plotted between P & T. P (V-b) = RT

 $P = \frac{R}{V-b} \times T + 0$ $\downarrow \qquad \downarrow$ $y \qquad \text{m x} + C$ $Slope = \frac{R}{V-b}$ P (V-b) = RT Isobaric curve is plotted between V and T PV - Pb = RT PV = RT + Pb

$$V = \frac{RT}{P} + \frac{Pb}{P}$$
$$V = \frac{RT}{P} + b$$
$$\downarrow \qquad \downarrow \qquad \downarrow$$
$$y \quad mx + C$$
$$\therefore Slope = \frac{R}{P}, Intercept = b$$

5. (c) Since P is constant and volume of container is also constant.

$$\therefore \mathbf{n}_1 \mathbf{T}_1 = \mathbf{n}_2 \mathbf{T}_2$$
$$\mathbf{n}_1 \times 300 = \frac{1}{3} \mathbf{n}_1 \times \mathbf{T}_2$$
$$900 \text{ K} = \mathbf{T}_2$$

 $(n_2 \text{ i.e., mole left inside the container are} = n_1 - \frac{2}{3}n_1)$

6. (b) Avg vel_A = RMS_B

$$\sqrt{\frac{8RT}{\pi M_A}} = \sqrt{\frac{3RT}{M_B}}$$

$$\sqrt{\frac{8}{\pi M_A}} = \sqrt{\frac{3}{M_B}}$$
Squaring both sides –

$$\frac{8}{\pi M_A} = \frac{3}{M_B}$$

$$\frac{8}{3\pi} = \frac{M_A}{M_B}$$

$$\therefore M_B > M_A$$
We wish,
Av. vel_A = Av. vel_B

$$\sqrt{\frac{8RT_A}{M_A\pi}} \approx \sqrt{\frac{8RT_B}{M_B\pi}}$$

Since, $M_B > M_A$, \therefore either T_B should be raised or T_A be lowered to make them equal

7. (a) Apply Charle's law-

For outer balloon

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$
$$\frac{1500}{300} = \frac{1800}{T_2}$$
$$T_2 = 360K$$
Outer halloon v

Outer balloon will reach its maximum volume at 360K and burst beyond it.

For inner balloon:

$$\frac{\mathbf{V}_1}{\mathbf{T}_1} = \frac{\mathbf{V}_2}{\mathbf{T}_2}$$
$$\frac{600}{\mathbf{E}_1} = \frac{800}{\mathbf{E}_2}$$

300 Τ, T = 400K

Inner balloon will reach its maximum volume at 400K and burst beyond it. So, on gradual heating outer balloon will burst first.

8. (c) $P_i = P, V_i = 30 L, T_i = 20 + 273 = 293 K$

 $P_f = 2P, V_f = ?; T_f = 100 + 273 = 373K$

Since, the number of moles of gas are constant we can apply,

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \implies \frac{P \times 30}{293} = \frac{2P \times V}{373}$$
$$\frac{30}{293} \times \frac{373}{2} = v = 19.09 \text{ L}$$

9. (*c*) Given that:

$$\sqrt{\frac{3RT_{H_2}}{2}} = \sqrt{7}\sqrt{\frac{3RT_{N_2}}{28}} \quad \left(\text{because } U_{\text{rms}} = \sqrt{\frac{3RT}{M}}\right)$$

So $\frac{3RT_{H_2}}{2} = 7 \times \frac{3 \times RT_{N_2}}{28}$ $\therefore T_{N_2} = 2T_{H_2}$ or $T_{N_2} > T_{H_2}$
10. (a) $\frac{V}{T} = \text{constant}(k)$
log V = log T + log k where $K = \frac{nR}{P}$
For 2 graph
log $\frac{nR}{P} = 0 = \log 1$
 $\frac{nR}{P} = 1$ (R = P) or n = 1 mol
for 1 graph
log $\frac{nR}{P} = \log 3$ or n = 3 mol
for 3 graph
log $\frac{nR}{P} = -\log 2 = \log \frac{1}{2}$

$$n = \frac{1}{2}$$
 mol

11. (c) According to kinetic theory of gases, volume of an ideal gas molecule and molecular attraction forces between ideal gas molecules are taken to be zero.

> According to kinetic theory of gases, average kinetic energy of gas molecules is directly proportional to the absolute temperature of the gas.

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

$$\begin{pmatrix} P + \frac{a}{V^2} \end{pmatrix} = \frac{RT}{V-b}$$

$$P = \frac{RT}{V-b} - \frac{a}{V^2}$$
Multiplying $\frac{V'}{RT}$ on both sides-
$$\frac{PV}{RT} = \frac{RT}{V-b} \times \frac{V}{RT} - \frac{a}{V^2} \times \frac{V}{RT}$$

$$Z = \frac{V}{V-b} - \frac{a}{RTV}$$
Boyles temp., $T = \frac{a}{Rb}$

$$Z = \frac{V}{V-b} - \frac{a \times Rb}{R \times a \times V} \implies Z = \frac{V}{V-b} - \frac{b}{V}$$

$$Z = \frac{V^2 - bV + b^2}{V(V-b)} \implies Z = \frac{V(V-b) + b^2}{V(V-b)}$$

$$Z = \frac{V(V-b)}{V(V-b)} + \frac{b^2}{V(V-b)} \implies Z = 1 + \frac{b^2}{V(V-b)}$$

$$(d) \quad 2A \rightarrow A_2$$

$$t = 0 \quad n \quad 0$$

$$t = t \quad n - x \quad x/2$$

$$\therefore P \propto \text{ moles } \therefore \frac{P_i}{n} = \frac{P_f}{n - \frac{x}{2}} \quad \dots (1)$$

1

13.

ЪΤ

Also, from Dalton's law, $p_{A_2} = X_{A_2} \times P_f$ $\therefore p_{A_2}$

$$= \left(\frac{\frac{x}{2}}{n - \frac{x}{2}}\right) \times P_{f} \qquad \dots \dots (2)$$

And, $p_{A_2} = \frac{1}{5} \times P_i$ (3) From equations (1), (2) and (3) $x = \frac{2n}{5}$ $\therefore \frac{P_i}{P_f} = \frac{n}{n - \frac{x}{2}} = \frac{n}{n - \frac{n}{5}} = \frac{5}{4}$ (\mathbf{x})

$$X_{A_2} = \left(\frac{\frac{x}{2}}{n - \frac{x}{2}}\right) = \frac{n/5}{4n/5} = 0.25$$

% dimension $=\frac{x}{n} \times 100 = \frac{2n/5}{n} \times 100 = 40\%$ Mol. mass of A = $10^{-22} \times 6 \times 10^{23} = 60$ amu

$$\therefore M_{\text{avg.}} = \frac{n_A \times M_A + n_{A_2} \times M_{A_2}}{n_A + n_{A_2}}$$
$$= \frac{\frac{3n}{5} \times 60 + \frac{n}{5} \times 120}{\frac{3n}{5} + \frac{n}{5}} = 75 \text{ amu}$$

14. (b) Moles of H₂ = $\frac{3 \times 16.42}{0.0821 \times 300} = 2$ Moles of D₂ = $\frac{6 \times 16.42}{0.0821 \times 300} = 4$ Average molecular weight = $\frac{2 \times 2 + 4 \times 4}{4 + 2} = \frac{10}{3}$

15. (*d*) At constant temperature, decrease in molecular mass causes flattening of the graph. For same molecular mass of gas, increase in temperature causes flattening of the graph.

$$(V_{mp})T_1 < (V_{mp})T_2$$

- **16.** (*c*) 1. The number of molecules with energy Ea or greater is proportional to the shaded area for each temperature.
 - 2. The graph follows the Maxwell Boltzmann energy distribution law.

17. (a)
$$PV = n RT$$

 $8.20 \times V = 100 \times 0.0821 \times T$
 $V = (8.21/8.20) T$
 $\Rightarrow V = T$
 $\Rightarrow \log V = \log T$
18. (a) $P_{atm} = P_{gas} + aqueous Tertiary + h/13.6$
 $P_{gas} = P_{atm} - [aqueous tension + (h/13.6)]$
19. (b) $PV = nRT$
 $P = (nR/V)T$
 $m = nR/V$
 $tan \theta_1 = m_1 = nR/V_1$
 $\Rightarrow m_2 = nR/V_2$
 $V_1 < V_2$
 $\theta_1 > \theta_2$

NEET Past 10 Years Questions

1. (c) According to Boyle's law

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

At a given pressure,

 $V \propto T$



- 2. (b) $n_{0_2} = \frac{4}{32} = \frac{1}{8}$ mol
 - $n_{H_2} = \frac{2}{2} = 1 \mod 1$

$$n_{Total} = n_{O_2} + n_{H_2} = \frac{1}{8} + 1 = \frac{9}{8} \text{ mol}$$

$$PV = nRT$$

$$P_{Total} \times 1 = \frac{9}{8} \times 0.082 \times 273$$

$$P_{Total} = 25.18 \text{ atm}$$

3. (b) The number of moles of N_2 and Ar are:

$$n_{N_2} = \frac{\text{given mass}}{\text{Molar mass}} = \frac{7\text{g}}{28\text{g}} = 0.25 \text{ mol}$$
$$n_{Ar} = \frac{8}{40} = 0.20 \text{ mol}$$

Now, Applying Dalton's law of partial pressure

Partial pressure of $N_2 = (mole fraction of N_2) \times Total pressure$

$$p_{N_2} = \left(\frac{0.25}{(0.20 + 0.25)}\right) \times 27bar = \frac{0.25}{0.45} \times 27bar$$

 $p_{N_2} = 15 bar$

- 4. (d) As per Boyle's Law $P_1V_1 = P_2V_2$ (1 bar) × 600 dm³ = P_2 × (150 dm³) $P_2 = 4$ bar
- 5. (c) Compressibility factor(Z) = $\frac{V_{real}}{V_{ideal}}$ $\therefore V_{real} < V_{ideal}$; Hence Z < 1

When Z < 1 attractive forces are dominating and gases will liquefy easily.

- 6. (c) In real gas equation, $\left(P + \frac{an^2}{V^2}\right)(V nb) = RT$ van der waal's constant, 'a' signifies intermolecular forces of attraction.
- **7.** (*a*) Van der waal constant 'a', signifies intermolecular forces of attraction.

Higher is the value of 'a', easier will be the liquefaction of gas.

8. (b) Max. pressure of CO₂ = Pressure of CO₂ at equilibrium For reaction,

$$SrCO_3(s) \implies SrO(s) + CO_2$$

1.6 atm = maximum pressure of CO_2 Volume of container at this stage,

$$V = \frac{nRT}{P} \qquad \dots \dots (i)$$

Since container is sealed and reaction was not earlier at equilibrium

 \therefore n = constant

$$n = \frac{PV}{RT} = \frac{0.4 \times 20}{RT} \qquad \dots \dots (ii)$$

Put equation (ii) in equation (i)

$$\mathbf{V} = \left[\frac{0.4 \times 20}{\mathrm{RT}}\right] \frac{\mathrm{RT}}{1.6} = 5 \mathrm{L}$$

9. (*c*) Let the no. of moles of each gas = x

Fraction of H escaped =
$$\frac{1}{2}x$$

 $\frac{r_{o_2}}{r_{H_2}} = \sqrt{\frac{M_{H_2}}{M_{o_2}}} = \sqrt{\frac{2}{32}} = \frac{1}{4} = \frac{V_{o_2} / t}{\frac{x}{2} / t}$
 $V_{o_2} = \frac{x}{8}$

- **10.** (*b*) Gas shows ideal behaviour when temperature is very high and pressure is low because of these two conditions gas molecules are less attracted towards each other.
- **11.** (*c*) Among the 4 options, dipole-induced dipole forces of attraction are present in He atom and HCl. HCl is a polar

molecule and because of its polar effect. He gets a dipole moment (induced).

- 12. (d) Maximum deviation from ideal gas is expected from NH₃(g) because it is a polar molecule, more attractive forces (H bonds) will be present between its molecules.
- 13. (c) Acc. to Graham's diffusion Law:

$$\frac{\mathbf{r}_{1}}{\mathbf{r}_{2}} = \sqrt{\frac{\mathbf{d}_{2}}{\mathbf{d}_{1}}} = \sqrt{\frac{\mathbf{M}_{2}}{\mathbf{M}_{1}}} = \frac{\mathbf{V}_{1}}{\mathbf{t}_{1}} \times \frac{\mathbf{t}_{2}}{\mathbf{V}_{2}}$$

So;
$$\frac{V_1 / t_1}{V_2 / t_2} = \sqrt{\frac{M_2}{M_1}}$$

 $\frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}} \Rightarrow \frac{200}{150} = \sqrt{\frac{36}{M_1}} \Rightarrow M_1 = 20.25$

14. (b) Acc. to Graham's law:

$$\frac{\mathbf{r}_{1}}{\mathbf{r}_{2}} = \sqrt{\frac{\mathbf{M}_{2}}{\mathbf{M}_{1}}} = \frac{\mathbf{V}_{1} / \mathbf{t}_{1}}{\mathbf{V}_{2} / \mathbf{t}_{2}}$$

Because both are same gases: So $V_1 = V_2$ Now, we can write

$$\therefore \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M}}$$
$$\Rightarrow \frac{3t_1}{t_1} = \sqrt{\frac{M_2}{4}} \Rightarrow M_2 = 36 \text{ u}$$

15. (c) 'a' \propto intermolecular forces of attraction

'b' \propto size of molecule

(I)
$$H_2 < He < O_2 < CO_2$$

(II) $CH_4 > O_2 > H_2$