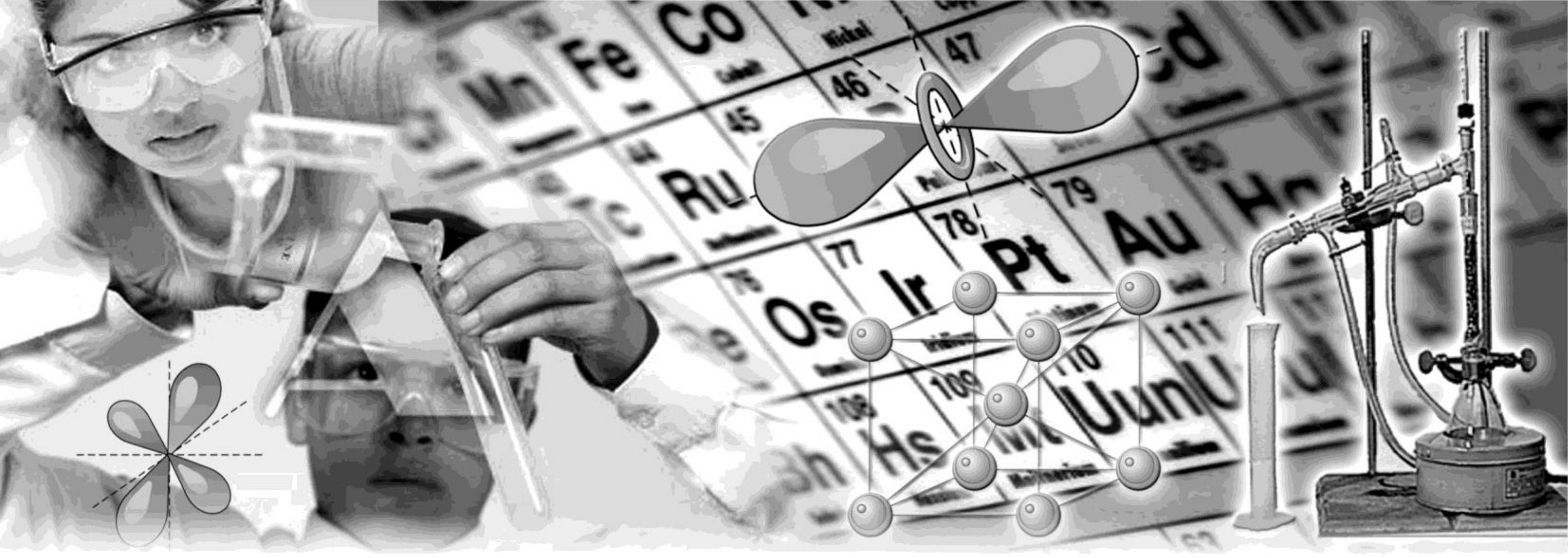


Chapter

6

Gaseous state

The state of matter in which the molecular forces of attraction between the particles of matter are minimum, is known as *gaseous state*. It is the simplest state and shows great uniformity in behaviour.

Characteristics of gases

(1) Gases or their mixtures are homogeneous in composition.

(2) Gases have very low density due to negligible intermolecular forces.

(3) Gases have *infinite expansibility and high compressibility*.

(4) Gases exert pressure.

(5) Gases possess high diffusibility.

(6) Gases do not have definite shape and volume like liquids.

(7) Gaseous molecules move very rapidly in all directions in a random manner *i.e.*, gases have highest kinetic energy.

(8) Gaseous molecules collide with one another and also with the walls of container with perfectly *elastic collisions*.

(9) Gases can be liquified, if subjected to low temperatures (below critical) or high pressures.

(10) Thermal energy of gases >> molecular attraction.

(11) Gases undergo similar change with the change of temperature and pressure. In other words, gases obey certain laws known as ***gas laws***.

Measurable properties of gases

(1) The characteristics of gases are described fully in terms of four parameters or measurable properties :

(i) The volume, V, of the gas.

(ii) Its pressure, P

(iii) Its temperature, T

(iv) The amount of the gas (*i.e.*, mass or number of moles).

(2) **Volume :** (i) Since gases occupy the entire space available to them, the measurement of volume of a gas only requires a measurement of the container confining the gas.

(ii) Volume is expressed in litres (L), millilitres (mL) or cubic centimetres (cm^3) or cubic metres (m^3).

(iii) $1L = 1000\, mL$; $1mL = 10^{-3} L$; $1L = 1\, dm^3 = 10^{-3} m^3$

$$1m^3 = 10^3\, dm^3 = 10^6\, cm^3 = 10^6\, mL = 10^3\, L$$

(3) **Mass :** (i) The mass of a gas can be determined by weighing the container in which the gas is enclosed and again weighing the container after removing the gas. The difference between the two weights gives the mass of the gas.

(ii) The mass of the gas is related to the number of moles of the gas *i.e.*

$$\text{moles of gas } (n) = \frac{\text{Mass in grams}}{\text{Molar mass}} = \frac{m}{M}$$

(4) **Temperature :** (i) Gases expand on increasing the temperature. If temperature is increased twice, the square of the velocity (v^2) also increases two times.

(ii) Temperature is measured in centigrade degree (^oC) or celsius degree with the help of thermometers. Temperature is also measured in Fahrenheit (F^o).

(iii) S.I. unit of temperature is kelvin (K) or absolute degree.

$$K = {}^oC + 273$$

(iv) Relation between F and ${}^oC$ is $\frac{{}^oC}{5} = \frac{F^o - 32}{9}$

(5) **Pressure :** (i) Pressure of the gas is the force exerted by the gas per unit area of the walls of the container in all directions. Thus, Pressure $(P) = \frac{\text{Force}(F)}{\text{Area}(A)} = \frac{\text{Mass}(m) \times \text{Acceleration}(a)}{\text{Area}(a)}$

(ii) Pressure exerted by a gas is due to kinetic energy $(KE = \frac{1}{2}mv^2)$ of the molecules. Kinetic energy of the gas molecules increases, as the temperature is increased. Thus, ***Pressure of a gas $\propto$ Temperature (T)***.

(iii) Pressure of a pure gas is measured by manometer while that of a mixture of gases by barometer.

(iv) Commonly two types of manometers are used,

(a) Open end manometer; (b) Closed end manometer

(v) The S.I. unit of pressure, the pascal (Pa), is defined as 1 newton per *metre square*. It is very small unit.

$$1 Pa = 1 Nm^{-2} = 1 kg\, m^{-1} s^{-2}$$

(vi) C.G.S. unit of pressure is dynes cm^{-2}.

(vii) M.K.S. unit of pressure is kgf/m^2. The unit kgf/cm^2 sometime called *ata* (atmosphere technical absolute).

(viii) Higher unit of pressure is *bar*, *KPa* or *MPa*.

$$1\, bar = 10^5\, Pa = 10^5\, Nm^{-2} = 100\, KNm^{-2} = 100\, KPa$$

(ix) Several other units used for pressure are,

Name	Symbol	Value
bar	*bar*	$1 bar = 10^5 Pa$
atmosphere	*atm*	$1\, atm = 1.01325 \times 10^5\, Pa$
Torr	Torr	$1\, Torr = \frac{101325}{760} Pa = 133.322\, Pa$
millimetre of mercury	mm *Hg*	$1\, mm\, Hg = 133.322\, Pa$

(x) The pressure relative to the atmosphere is called *gauge pressure*. The pressure relative to the perfect vacuum is called *absolute pressure*.

Absolute pressure = Gauge pressure + Atmosphere pressure.

(xi) When the pressure in a system is less than atmospheric pressure, the gauge pressure becomes negative, but is frequently designated and called vacuum. For example, 16 *cm* vacuum will be

$$\frac{76 - 16}{76} \times 1.013 = 0.80\, bar.$$

(xii) If '*h*' is the height of the fluid in a column or the difference in the heights of the fluid columns in the two limbs of the manometer, *d* is the density of the fluid $(Hg = 13.6 \times 10^3\, Kg/m^3 = 13.6\, g/cm^3)$ and *g* is the gravity, then pressure is given by, $P_{\text{gas}} = P_{\text{atm}} + hdg$

(xiii) Two sets of conditions are widely used as 'standard' values for reporting data.

Condition	T	P	V_m (Molar volume)
S.T.P./N.T.P.	273.15 K	1 *atm*	22.414 L
S.A.T.P*.	298.15 K	1 *bar*	24.800 L

* Standard ambient temperature and pressure.

Boyle's law

(1) In 1662, ***Robert Boyle*** discovered the first of several relationships among gas variables (P, T, V).

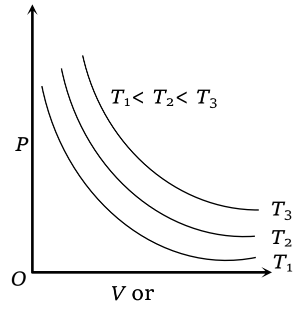
(2) It states that, "*For a fixed amount of a gas at constant temperature, the gas volume is inversely proportional to the gas pressure.*"

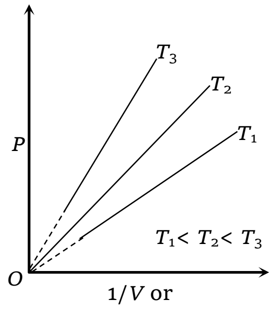
Thus, $P \propto 1/V$ at constant temperature and mass

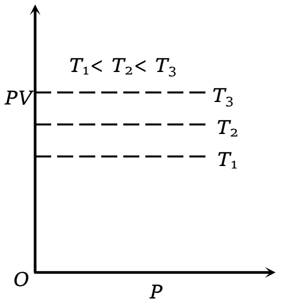
or $P = K/V$ (where K is constant)

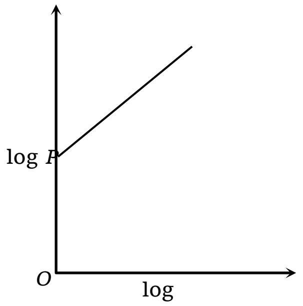
or $PV = K$ or $P_1V_1 = P_2V_2 = K$ (For two or more gases)

(3) ***Graphical representation of Boyle's law*** : Graph between P and V at constant temperature is called ***isotherm*** and is an equilateral (or rectangular) hyperbola. By plotting P versus $1/V$, this hyperbola can be converted to a straight line. Other types of isotherms are also shown below,









(4) At constant mass and temperature density of a gas is directly proportional to its pressure and inversely proportional to its volume.

Thus, $d \propto P \propto \frac{1}{V}$ $\left[\because V = \frac{\text{mass}}{d}\right]$

or $\frac{d_1}{d_2} = \frac{P_1}{P_2} = \frac{V_2}{V_1} = ... = K$

(5) At altitudes, as P is low d of air is less. That is why mountaineers carry oxygen cylinders.

Charle's law

(1) French chemist, ***Jacques Charles*** first studied variation of volume with temperature, in 1787.

(2) It states that, "*The volume of a given mass of a gas is directly proportional to the absolute temperature* $(= {}^oC + 273)$ *at constant pressure*".

Thus, $V \propto T$ at constant pressure and mass

or $V = KT = K(t({}^oC) + 273.15)$, (where k is constant),

$K = \frac{V}{T}$ or $\frac{V_1}{T_1} = \frac{V_2}{T_2} = K$ (For two or more gases)

(3) If $t = 0^oC$, then $V = V_0$

hence, $V_0 = K \times 273.15$

$\therefore K = \frac{V_0}{273.15}$

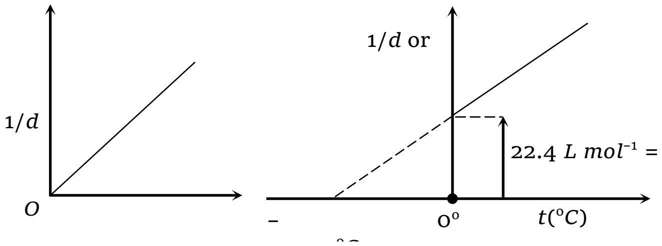
$$V = \frac{V_0}{273.15}[t + 273.15] = V_0\left[1 + \frac{t}{273.15}\right] = V_0[1 + \alpha_v t]$$

where α_v is the volume coefficient,

$$\alpha_v = \frac{V - V_0}{tV_0} = \frac{1}{273.15} = 3.661 \times 10^{-3}\ {}^oC^{-1}$$

Thus, for every 1^o change in temperature, the volume of a gas changes by $\frac{1}{273.15}\left(\approx \frac{1}{273}\right)$ of the volume at 0^oC.

(4) ***Graphical representation of Charle's law*** **:** Graph between V and T at constant pressure is called *isobar or isoplestics* and is always a straight line. A plot of V versus $t({}^oC)$ at constant pressure is a straight line cutting the temperature axis at -273.15^oC. It is the lowest possible temperature.



(5) At constant mass and pressure density of a gas is inversely proportional to it absolute temperature.

Thus, $d \propto \frac{1}{T} \propto \frac{1}{V}$ $\left[\because V = \frac{\text{mass}}{\text{d}}\right]$

or $\frac{d_1}{d_2} = \frac{T_2}{T_1} = \frac{V_2}{V_1} = ... = K$

(6) Use of hot air balloons in sports and meteorological observations is an application of Charle's law.

Gay-Lussac's law (Amonton's law)

(1) In 1802, French chemist ***Joseph Gay-Lussac*** studied the variation of pressure with temperature and extende the Charle's law so, this law is also called Charle's-Gay Lussac's law.

(2) It states that, "*The pressure of a given mass of a gas is directly proportional to the absolute temperature* $(= {}^oC + 273)$ *at constant volume.*"

Thus, $P \propto T$ at constant volume and mass

or $P = KT = K(t({}^o C) + 273.15)$ (where K is constant)

$K = \frac{P}{T}$ or $\frac{P_1}{T_1} = \frac{P_2}{T_2} = K$ (For two or more gases)

(3) If $t = 0^oC$, then $P = P_0$

Hence, $P_0 = K \times 273.15$

$\therefore$ $K = \frac{P_0}{273.15}$

$$P = \frac{P_0}{273.15}[t + 273.15] = P_0\left[1 + \frac{t}{273.15}\right] = P_0[1 + \alpha t]$$

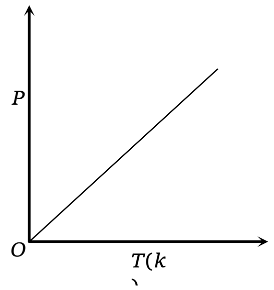
where α_P is the pressure coefficient,

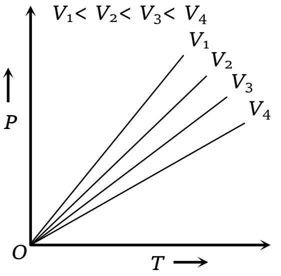
$$\alpha_P = \frac{P - P_0}{tP_0} = \frac{1}{273.15} = 3.661 \times 10^{-3}\ {}^oC^{-1}$$

Thus, for every 1^o change in temperature, the pressure of a gas changes by $\frac{1}{273.15}\left(\approx \frac{1}{273}\right)$ of the pressure at 0^oC.

(4) This law fails at low temperatures, because the volume of the gas molecules be come significant.

(5) ***Graphical representation of Gay-Lussac's law* :** A graph between P and T at constant V is called ***isochore***.





Avogadro's law

(1) According to this law, "*Equal volumes of any two gases at the same temperature and pressure contain the same number of molecules*."

Thus, $V \propto n$ (at constant T and P)

or $V = Kn$ (where K is constant)

or $\frac{V_1}{n_1} = \frac{V_2}{n_2} = ... = K$

Example, $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

$2H_2(g)$	$O_2(g)$	$2H_2O(g)$
2 moles	1 mole	2 moles
2 volumes	1 volume	2 volumes
2 litres	1 litre	2 litres
1 litre	1 / 2 litre	1 litre
1n litre	1 / 2n litre	1n litre

(2) One mole of any gas contains the same number of molecules (Avogadro's number $= 6.02 \times 10^{23}$) and by this law must occupy the same volume at a given temperature and pressure. The volume of one mole of a gas is called **molar volume, V_m** which is 22.4 $L\ mol^{-1}$ at S.T.P. or N.T.P.

(3) This law can also express as, "*The molar gas volume at a given temperature and pressure is a specific constant independent of the nature of the gas*".

Thus, $V_m =$ specific constant $= 22.4\, L\, mol^{-1}$ at S.T.P. or N.T.P.

Ideal gas equation

(1) The simple gas laws relating gas volume to pressure, temperature and amount of gas, respectively, are stated below :

Boyle's law : $P \propto \frac{1}{V}$ or $V \propto \frac{1}{P}$ (n and T constant)

Charle's law : $V \propto T$ (n and P constant)

Avogadro's law : $V \propto n$ (T and P constant)

If all the above law's combines, then

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

or $$V = \frac{nRT}{P} \quad (R = \text{ Ideal gas constant})$$

or $$PV = nRT$$

This is called ***ideal gas equation***. R is called ***ideal gas constant***. This equation is obeyed by isothermal and adiabatic processes.

(2) **Nature and values of R :** From the ideal gas equation, $R = \frac{PV}{nT} = \frac{\text{Pressure} \times \text{Volume}}{\text{mole} \times \text{Temperature}}$

$$= \frac{\frac{\text{Force}}{\text{Area}} \times \text{Volume}}{\text{mole} \times \text{Temperature}} = \frac{\text{Force} \times \text{Length}}{\text{mole} \times \text{Temperature}}$$

$$= \frac{\text{Work or energy}}{\text{mole} \times \text{Temperature}}.$$

R is expressed in the unit of work or energy $mol^{-1}\ K^{-1}$.

Since different values of R are summarised below :

$R = 0.0821\ L\, atm\, mol^{-1}\ K^{-1}$

$= 8.3143\ joule\ mol^{-1}\ K^{-1}$ (S.I. unit)

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

$= 8.3143\ KPa\, dm^3\ mol^{-1}\ K^{-1}$

$= 8.3143\ MPa\, cm^3\ mol^{-1}\ K^{-1}$

$= 5.189 \times 10^{19}\ eV\ mol^{-1}\ K^{-1}$

$= 1.99\ cal\, mol^{-1}\ K^{-1}$

(3) Gas constant, R for a single molecule is called *Boltzmann constant* (k)

$$k = \frac{R}{N} = \frac{8.314 \times 10^7}{6.023 \times 10^{23}} ergs\, mole^{-1}\ degree^{-1}$$

$$= 1.38 \times 10^{-16}\ ergs\, mol^{-1}\ degree^{-1}$$

or $1.38 \times 10^{-23}\ joule\ mol^{-1}\ degree^{-1}$

(4) **Calculation of mass, molecular weight and density of the gas by gas equation**

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

$$\left(\because n = \frac{\text{mass of the gas } (m)}{\text{Molecular weight of the gas } (M)} \right)$$

$$\therefore \quad M = \frac{mRT}{PV}$$

$$d = \frac{PM}{RT} \qquad \left(\because d = \frac{m}{V} \right)$$

or $\frac{dT}{P}=\frac{M}{R}$, $\frac{M}{R}=$ Constant

($\because$ M and R are constant for a particular gas)

Thus, $\frac{dT}{P}$ or $\frac{d_1T_1}{P_1}=\frac{d_2T_2}{T_2}$ = Constant

(For two or more different temperature and pressure)

(5) Gas densities differ from those of solids and liquids as,

(i) Gas densities are generally stated in g/L instead of g/cm^3.

(ii) Gas densities are strongly dependent on pressure and temperature as, $d \propto P \propto 1/T$

Densities of liquids and solids, do depend somewhat on temperature, but they are far less dependent on pressure.

(iii) The density of a gas is directly proportional to its molar mass. No simple relationship exists between the density and molar mass for liquid and solids.

(iv) Density of a gas at STP $=\frac{\text{molar mass}}{22.4}$

$d(N_2)$ at STP$=\frac{28}{22.4}=1.25\,g\,L^{-1}$,

$d(O_2)$ at STP $=\frac{32}{22.4}=1.43\,g\,L^{-1}$

Dalton's law of partial pressures

(1) According to this law, "*When two or more gases, which do not react chemically are kept in a closed vessel, the total pressure exerted by the mixture is equal to the sum of the partial pressures of individual gases*."

Thus, $P_{\text{total}}=P_1+P_2+P_3+.........$

Where $P_1, P_2, P_3,......$ are partial pressures of gas number 1, 2, 3 ....

(2) ***Partial pressure*** is the pressure exerted by a gas when it is present alone in the same container and at the same temperature.

Partial pressure of a gas

$$(P_1)=\frac{\text{Number of moles of the gas }(n_1)\times P_{\text{Total}}}{\text{Total number of moles }(n)\text{ in the mixture}}=\text{Mole fraction }(X_1)\times P_{\text{Total}}$$

(3) If a number of gases having volume $V_1, V_2, V_3......$ at pressure $P_1, P_2, P_3........$ are mixed together in container of volume V, then,

$$P_{\text{Total}}=\frac{P_1V_1+P_2V_2+P_3V_3.....}{V}$$

or $\quad =(n_1+n_2+n_3.....)\frac{RT}{V} \quad (\because PV=nRT)$

or $\quad =n\frac{RT}{V} \quad (\because n=n_1+n_2+n_3.....)$

(4) **Applications :** This law is used in the calculation of following relationships,

(i) Mole fraction of a gas (X_1) in a mixture of gas $=\frac{\text{Partial pressure of a gas }(P_1)}{P_{\text{Total}}}$

(ii) % of a gas in mixture $=\frac{\text{Partial pressure of a gas }(P_1)}{P_{\text{Total}}}\times 100$

(iii) Pressure of dry gas collected over water : When a gas is collected over water, it becomes moist due to water vapour which exerts its own partial pressure at the same temperature of the gas. *This partial perssure of water vapours is called aqueous tension*. Thus, $P_{\text{dry gas}}=P_{\text{moist gas}}$ or $P_{\text{Total}}-P_{\text{water vapour}}$

or $P_{\text{dry gas}}=P_{\text{moist gas}}-$ Aqueous tension (Aqueous tension is directly proportional to absolute temperature)

(iv) Relative humidity (RH) at a given temperature is given by,

$$RH=\frac{\text{Partial pressure of water in air}}{\text{Vapour pressure of water}}.$$

(5) ***Limitations* :** This law is applicable only when the component gases in the mixture do not react with each other. For example, N_2 and O_2, CO and CO_2, N_2 and Cl_2, CO and N_2 etc. But this law is not applicable to gases which combine chemically. For example, H_2 and Cl_2, CO and Cl_2, NH_3, HBr and HCl, NO and O_2 etc.

(6) Another law, which is really equivalent to the law of partial pressures and related to the partial volumes of gases is known as ***Law of partial volumes*** given by ***Amagat***. According to this law, "*When two or more gases, which do not react chemically are kept in a closed vessel, the total volume exerted by the mixture is equal to the sum of the partial volumes of individual gases.*"

Thus, $V_{\text{Total}}=V_1+V_2+V_3+......$

Where $V_1, V_2, V_3,......$ are partial volumes of gas number 1, 2, 3.....

Graham's law of diffusion and Effusion

(1) **Diffusion** is the process of spontaneous spreading and intermixing of gases to form homogenous mixture irrespective of force of gravity. While **Effusion** is the escape of gas molecules through a tiny hole such as pinhole in a balloon.

- All gases spontaneously diffuse into one another when they are brought into contact.
- Diffusion into a vacuum will take place much more rapidly than diffusion into another place.

- Both the rate of diffusion of a gas and its rate of effusion depend on its molar mass. Lighter gases diffuses faster than heavier gases. The gas with highest rate of diffusion is *hydrogen*.

(2) According to this law, "*At constant pressure and temperature, the rate of diffusion or effusion of a gas is inversely proportional to the square root of its vapour density*."

Thus, rate of diffusion $(r)\propto \frac{1}{\sqrt{d}}$ (T and P constant)

For two or more gases at constant pressure and temperature,

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

(3) Graham's law can be modified in a number of ways as,

(i) Since, 2 × vapour density (V.D.) = Molecular weight

then, $\frac{r_1}{r_2}=\sqrt{\frac{d_2}{d_1}}=\sqrt{\frac{d_2\times 2}{d_1\times 2}}=\sqrt{\frac{M_2}{M_1}}$

where, M_1 and M_2 are the molecular weights of the two gases.

(ii) Since, rate of diffusion $(r)=\frac{\text{Volume of a gas diffused}}{\text{Time taken for diffusion}}$ then,

$$\frac{r_1}{r_2}=\frac{V_1/t_1}{V_2/t_2}=\frac{w_1/t_1}{w_2/t_2}=\sqrt{\frac{d_2}{d_1}}$$

(a) When equal volume of the two gases diffuse, *i.e.* $V_1=V_2$

then, $\frac{r_1}{r_2}=\frac{t_2}{t_1}=\sqrt{\frac{d_2}{d_1}}$

(b) When volumes of the two gases diffuse in the same time, *i.e.* $t_1=t_2$

then, $\frac{r_1}{r_2}=\frac{V_1}{V_2}=\sqrt{\frac{d_2}{d_1}}$

(iii) Since, $r\propto p$ (when p is not constant)

then, $\frac{r_1}{r_2}=\frac{P_1}{P_2}\sqrt{\frac{M_2}{M_1}}$ $\left(\because r\propto\frac{P}{\sqrt{M}}\right)$

(4) Rate of diffusion and effusion can be determined as,

(i) Rate of diffusion is equal to distance travelled by gas per unit time through a tube of uniform cross-section.

(ii) Number of moles effusing per unit time is also called rate of diffusion.

(iii) Decrease in pressure of a cylinder per unit time is called rate of effusion of gas.

(iv) The volume of gas effused through a given surface per unit time is also called rate of effusion.

(5) **Applications :** Graham's law has been used as follows,

(i) To determine vapour densities and molecular weights of gases.

(ii) To prepare Ausell's marsh gas indicator, used in mines.

(iii) ***Atmolysis* :** The process of separation of two gases on the basis of their different rates of diffusion due to difference in their densities is called atmolysis. It has been applied with success for the separation of isotopes and other gaseous mixtures.

Kinetic theory of gases

(1) Kinetic theory was developed by Bernoulli, Joule, Clausius, Maxwell and Boltzmann etc. and represents ***dynamic particle or microscopic model*** for different gases since it throws light on the behaviour of the particles (atoms and molecules) which constitute the gases and cannot be seen. Properties of gases which we studied earlier are part of macroscopic model.

(2) **Postulates**

(i) Every gas consists of a large number of small particles called molecules moving with very high velocities in all possible directions.

(ii) The volume of the individual molecule is negligible as compared to the total volume of the gas.

(iii) Gaseous molecules are *perfectly elastic* so that there is no net loss of kinetic energy due to their collisions.

(iv) The effect of gravity on the motion of the molecules is negligible.

(v) Gaseous molecules are considered as *point masses* because they do not posses potential energy. So the attractive and repulsive forces between the gas molecules are negligible.

(vi) The pressure of a gas is due to the continuous bombardment on the walls of the containing vessel.

(vii) At constant temperature the average K.E. of all gases is same.

(viii) The average K.E. of the gas molecules is directly proportional to the absolute temperature.

(3) **Kinetic gas equation :** On the basis of above postulates, the following gas equation was derived,

$$PV=\frac{1}{3}mnu^2_{rms}$$

where, P = pressure exerted by the gas

V = volume of the gas

m = average mass of each molecule

n = number of molecules

u = root mean square (RMS) velocity of the gas.

(4) **Calculation of kinetic energy**

We know that,

K.E. of one molecule $= \frac{1}{2} mu^2$

K.E. of n molecules $= \frac{1}{2} mnu^2 = \frac{3}{2} PV$

$(\because\ PV = \frac{1}{3} mnu^2)$

$n = 1$, Then K.E. of 1 mole gas $= \frac{3}{2} RT$

$(\because PV = RT)$

$$= \frac{3}{2} \times 8.314 \times T = 12.47\ T\ Joules\ .$$

$$= \frac{\text{Average K.E. per mole}}{N(\text{Avogadro number})} = \frac{3}{2}\frac{RT}{N} = \frac{3}{2} KT$$

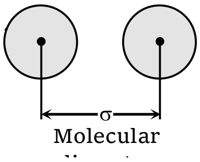
$$\left(K = \frac{R}{N} = \text{Boltzmann constant} \right)$$

This equation shows that K.E. of translation of a gas depends only on the absolute temperature. This is known as *Maxwell generalisation*. Thus average K.E. ∝ T.

If $T = 0K$ (*i.e.*, $-273.15^o C$) then, average K.E. = 0. Thus, absolute zero (0K) is the temperature at which molecular motion ceases.

Molecular collisions

(1) The closest distance between the centres of two molecules taking part in a collision is called *molecular or collision diameter* (σ). The molecular diameter of all the gases is nearly same lying in the order of 10^{-8} *cm*



(2) The number of collisions taking place in unit time per unit volume, called *collision frequency* (z).

(i) The number of collision made by a single molecule with other molecules per unit time are given by, $Z_A = \sqrt{2}\pi\sigma^2 u_{\text{av.}} n$

where n is the number of molecules per unit molar volume,

$$n = \frac{\text{Avogadro number}(N_0)}{V_m} = \frac{6.02 \times 10^{23}}{0.0224} m^{-3}$$

(ii) The total number of bimolecular collision per unit time are given by, $Z_{AA} = \frac{1}{\sqrt{2}} \pi\sigma^2 u_{\text{av.}} n^2$

(iii) If the collisions involve two unlike molecules, the number of bimolecular collision are given by,

$$Z_{AB} = \sigma_{AB}^2 \left[8\pi RT \frac{(M_A + M_B)}{M_A M_B} \right]^{1/2}$$

where, $\sigma_{AB} = \frac{\sigma_A + \sigma_B}{2}$

M_A, M_B are molecular weights $(M = mN_0)$

(iv) (a) At particular temperature; $Z \propto p^2$

(b) At particular pressure; $Z \propto T^{-3/2}$

(c) At particular volume; $Z \propto T^{1/2}$

(3) During molecular collisions a molecule covers a small distance before it gets deflected. *The average distance travelled by the gas molecules between two successive collision is called mean free path (λ).*

$$\lambda = \frac{\text{Average distance travelled per unit time}(u_{\text{av}})}{\text{No. of collisions made by single molecule per unit time }(Z_A)} .$$

$$= \frac{u_{\text{av}}}{\sqrt{2}\pi\sigma^2 u_{\text{avr.}} n} = \frac{1}{\sqrt{2}\pi n \sigma^2}$$

(4) Based on kinetic theory of gases mean free path, $\lambda \propto \frac{T}{P}$. Thus,

(i) Larger the size of the molecules, smaller the mean free path, *i.e.*, $\lambda \propto \frac{1}{(\text{radius})^2}$

(ii) Greater the number of molecules per unit volume, smaller the mean free path.

(iii) Larger the temperature, larger the mean free path.

(iv) Larger the pressure, smaller the mean free path.

(5) Relation between collision frequency (Z) and mean free path (λ) is given by, $Z = \frac{u_{rms}}{\lambda}$

Molecular speeds or velocities

(1) At any particular time, in the given sample of gas all the molecules do not possess same speed, due to the frequent molecular collisions with the walls of the container and also with one another, the molecules move with ever changing speeds and also with ever changing direction of motion.

(2) According to Maxwell, at a particular temperature the distribution of speeds remains constant and this distribution is referred to as the *Maxwell-Boltzmann distribution* and given by the following expression,

$$\frac{dn_0}{n} = 4\pi \left(\frac{M}{2\pi RT} \right)^{3/2} .e^{-Mu^2/2RT} .u^2 dc$$

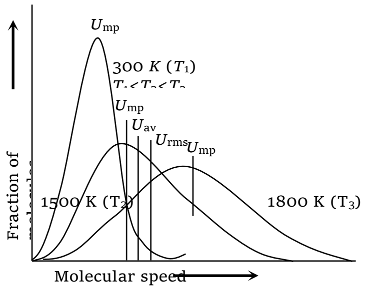
where, $dn_0 =$ Number of molecules out of total number of molecules n, having velocities between c and $c+dc$, $dn_0/n =$ Fraction of the total number of molecules, M = molecular weight, T = absolute temperature. The exponential factor $e^{-Mu^2/2RT}$ is called *Boltzmann factor*.

(3) Maxwell gave distribution curves of molecular speeds for CO_2 at different temperatures. Special features of the curve are :

(i) Fraction of molecules with two high or two low speeds is very small.

(ii) No molecules has zero velocity.

(iii) Initially the fraction of molecules increases in velocity till the peak of the curve which pertains to most probable velocity and thereafter it falls with increase in velocity.



(4) **Types of molecular speeds or Velocities**

(i) ***Root mean square velocity*** (u_{rms}) **:** It is the square root of the mean of the squares of the velocity of a large number of molecules of the same gas.

$$u_{rms} = \sqrt{\frac{u_1^2 + u_2^2 + u_3^2 + ...u_n^2}{n}}$$

$$u_{rms} = \sqrt{\frac{3PV}{(mN_0) = M}} = \sqrt{\frac{3RT}{(mN_0) = M}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3P}{d}}$$

where k = Boltzmann constant $= \frac{R}{N_0}$

(a) For the same gas at two different temperatures, the ratio of *RMS* velocities will be, $\frac{u_1}{u_2} = \sqrt{\frac{T_1}{T_2}}$

(b) For two different gases at the same temperature, the ratio of *RMS* velocities will be, $\frac{u_1}{u_2} = \sqrt{\frac{M_2}{M_1}}$

(c) *RMS* velocity at any temperature t^oC may be related to its value at S.T.P. as, $u_t = \sqrt{\frac{3P(273+t)}{273d}}$.

(ii) ***Average velocity*** (v_{av}) **:** It is the average of the various velocities possessed by the molecules.

$$v_{av} = \frac{v_1 + v_2 + v_3 + ...v_n}{n} \; ; \; v_{av} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8kT}{\pi m}}$$

(iii) ***Most probable velocity*** (α_{mp})**:** It is the velocity possessed by maximum number of molecules of a gas at a given temperature.

$$\alpha_{mp} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2PV}{M}} = \sqrt{\frac{2P}{d}}$$

(5) **Relation between molecular speeds or velocities,**

(i) Relation between u_{rms} and v_{av}: $v_{av} = 0.9213 \times u_{rms}$

or $u_{rms} = 1.085 \times v_{av}$

(ii) Relation between α_{mp} and u_{rms}: $\alpha_{mp} = 0.816 \times u_{rms}$

or $u_{rms} = 1.224 \times \alpha_{mp}$

(iii) Relation between α_{mp} and v_{av}: $v_{av} = 1.128 \times \alpha_{mp}$

(iv) Relation between α_{mp}, v_{av} and u_{rms}

$$\alpha_{mp} \; : \; v_{av} \; : \; u_{rms}$$

$$\sqrt{\frac{2RT}{M}} \; : \; \sqrt{\frac{8RT}{\pi M}} \; : \; \sqrt{\frac{3RT}{M}}$$

$$\sqrt{2} \; : \; \sqrt{\frac{8}{\pi}} \; : \; \sqrt{3}$$

$$1.414 \; : \; 1.595 \; : \; 1.732$$

$$1 \; : \; 1.128 \; : \; 1.224$$

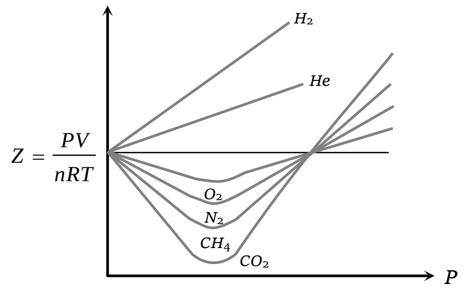
i.e., $\alpha_{mp} < v_{av} < u_{rms}$

Real and Ideal gases

(1) Gases which obey gas laws or ideal gas equation $(PV = nRT)$ at all temperatures and pressures are called *ideal or perfect gases*. Almost all gases deviate from the ideal behaviour *i.e.*, *no gas is perfect* and the concept of perfect gas is only theoretical.

(2) Gases tend to show ideal behaviour more and more as the temperature rises above the boiling point of their liquefied forms and the pressure is lowered. Such gases are known as *real or non ideal gases*. Thus, a "*real gas is that which obeys the gas laws under low pressure or high temperature*".

(3) The deviations can be displayed, by plotting the *P*-*V* isotherms of real gas and ideal gas.



(4) It is difficult to determine quantitatively the deviation of a real gas from ideal gas behaviour from the *P-V* isotherm curve as shown above. Compressibility factor *Z* defined by the equation,

$$PV = ZnRT \text{ or } Z = PV/nRT = PV_m/RT$$

is more suitable for a quantitative description of the deviation from ideal gas behaviour.

(5) Greater is the departure of *Z* from unity, more is the deviation from ideal behaviour. Thus, when

(i) $Z = 1$, the gas is ideal at all temperatures and pressures. In case of N_2, the value of *Z* is close to 1 at $50^o C$. This temperature at which a real gas exhibits ideal behaviour, for considerable range of pressure, is known as *Boyle's temperature or Boyle's point* (T_B).

(ii) $Z > 1$, the gas is less compressible than expected from ideal behaviour and shows positive deviation, usual at high *P i.e.* $PV > RT$.

(iii) $Z < 1$, the gas is more compressible than expected from ideal behaviour and shows negative deviation, usually at low *P i.e.* $PV < RT$.

(iv) $Z > 1$ for H_2 and *He* at all pressure *i.e.*, always shows positive deviation.

(v) The most easily liquefiable and highly soluble gases (NH_3, SO_2) show larger deviations from ideal behaviour *i.e.* $Z << 1$.

(vi) Some gases like CO_2 show both negative and positive deviation.

(6) **Causes of deviations of real gases from ideal behaviour :** The ideal gas laws can be derived from the kinetic theory of gases which is based on the following two important assumptions,

(i) The volume occupied by the molecules is negligible in comparison to the total volume of gas.

(ii) The molecules exert no forces of attraction upon one another. It is because neither of these assumptions can be regarded as applicable to real gases that the latter show departure from the ideal behaviour.

Vander Waal's equation

(1) To rectify the errors caused by ignoring the intermolecular forces of attraction and the volume occupied by molecules, Vander Waal (in 1873) modified the ideal gas equation by introducing two corrections,

(i) Volume correction (ii) Pressure correction

(2) Vander Waal's equation is obeyed by the real gases over wide range of temperatures and pressures, hence it is called *equation of state for the real gases.*

(3) The Vander Waal's equation for *n moles* of the gas is,

$$\underbrace{\left(P + \frac{n^2 a}{V^2}\right)}_{\substack{\text{Pressure correction}\\ \text{for molecular attraction}}} \quad \underbrace{[V - nb]}_{\substack{\text{Volume correction for}\\ \text{finite size of molecules}}} \quad = nRT$$

a and *b* are Vander Waal's constants whose values depend on the nature of the gas. Normally for a gas $a >> b$.

(i) ***Constant a*** : It is a indirect measure of magnitude of attractive forces between the molecules. Greater is the value of *a*, more easily the gas can be liquefied. Thus the easily liquefiable gases (like $SO_2 > NH_3 > H_2S > CO_2$) have high values than the permanent gases (like $N_2 > O_2 > H_2 > He$).

Units of '*a*' are : *atm.* $L^2 \, mol^{-2}$ or *atm.* $m^6 mol^{-2}$ or $N m^4 \, mol^{-2}$ (S.I. unit).

(ii) ***Constant b*** **:** Also called co-volume or excluded volume,

$$b = 4N_0 v = 4N_0 \left(\frac{4}{3}\pi r^3\right)$$

It's value gives an idea about the effective size of gas molecules. Greater is the value of *b*, larger is the size and smaller is the compressible volume. As *b* is the effective volume of the gas molecules, the constant value of *b* for any gas over a wide range of temperature and pressure indicates that the *gas molecules are incompressible*.

Units of '*b*' are : $L \, mol^{-1}$ or $m^3 \, mol^{-1}$ (S.I. unit)

(iii) The two Vander Waal's constants and Boyle's temperature (T_B) are related as,

$$T_B = \frac{a}{bR}$$

(4) **Vander Waal's equation at different temperature and pressures**

(i) ***When pressure is extremely low*** : For one mole of gas,

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT \text{ or } PV = RT - \frac{a}{V} + Pb + \frac{ab}{V^2}$$

(ii) ***When pressure is extremely high*** **:** For one mole of gas,

$$PV = RT + Pb \, ; \; \frac{PV}{RT} = 1 + \frac{Pb}{RT} \text{ or } Z = 1 + \frac{Pb}{RT}$$

where *Z* is compressibility factor.

(iii) ***When temperature is extremely high*** **:** For one mole of gas,

$$PV = RT.$$

(iv) ***When pressure is low*** **:** For one mole of gas,

$$\left(P+\frac{a}{V^2}\right)(V-b)=RT \text{ or } PV=RT+Pb-\frac{a}{V}+\frac{ab}{V^2}$$

$$\text{or } \frac{PV}{RT}=1-\frac{a}{VRT} \text{ or } Z=1-\frac{a}{VRT}$$

(v) ***For hydrogen*** **:** Molecular mass of hydrogen is small hence value of '*a*' will be small owing to smaller intermolecular force. Thus the terms $\frac{a}{V}$ and $\frac{ab}{V^2}$ may be ignored. Then Vander Waal's equation becomes,

$$PV=RT+Pb \text{ or } \frac{PV}{RT}=1+\frac{Pb}{RT}$$

$$\text{or } Z=1+\frac{Pb}{RT}$$

In case of hydrogen, compressibility factor is always greater than one.

(5) **Merits of Vander Waal's equation**

(i) The Vander Waal's equation holds good for real gases upto moderately high pressures.

(ii) The equation represents the trend of the isotherms representing the variation of *PV* with *P* for various gases.

(iii) From the Vander Waal's equation it is possible to obtain expressions of Boyle's temperature, critical constants and inversion temperature in terms of the Vander Waal's constants '*a*' and '*b*'.

(iv) Vander Waal's equation is useful in obtaining a 'reduced equation of state' which being a general equation of state has the advantage that a single curve can be obtained for all gases when the equation if graphically represented by plotting the variables.

(6) **Limitations of Vander Waal's equation**

(i) This equation shows appreciable deviations at too low temperatures or too high pressures.

(ii) The values of Vander Waal's constants *a* and *b* do not remain constant over the entire ranges of *T* and *P*, hence this equation is valid only over specific range of *T* and *P*.

(7) **Other equations of state :** In addition to Vander Waal's equation, there are also equations of state which have been used to explain real behaviour of gases are,

(i) ***Clausius equation*** **:** $\left[P+\frac{a}{T(V+c)^2}\right](V-b)=RT$. Here 'c' is another constant besides *a*, *b* and *R*.

(ii) ***Berthelot equation*** **:** $\left(P+\frac{a}{TV^2}\right)(V-b)=RT$.

(iii) ***Wohl equation*** **:** $P=\frac{RT}{(V-b)}-\frac{a}{V(V-b)}+\frac{c}{V^2}$

(iv) ***Dieterici equation*** **:** $P=\frac{RT}{V-b}.e^{-a/RTV}$.

The expression is derived on the basis of the concept that molecules near the wall will have higher potential energy than those in the bulk.

(v) ***Kammerlingh Onnes equation*** **:** It is the most general or satisfactory expression as equation expresses *PV* as a power series of *P* at a given temperature.

$$PV=A+BP+CP^2+DP^3+......$$

Here coefficients *A*, *B*, *C* etc. are known as first, second and third etc. virial coefficients.

(a) Virial coefficients are different for different gases.

(b) At very low pressure, first virial coefficient, *A* = *RT*.

(c) At high pressure, other virial coefficients also become important and must be considered.

The critical state

(1) A state for every substance at which the vapour and liquid states are indistinguishable is known as *critical state*. It is defined by critical temperature and critical pressure.

(2) **Critical temperature (T_c)** of a gas is that temperature above which the gas cannot be liquified however large pressure is applied. It is given by, $T_c=\frac{8a}{27Rb}$

(3) **Critical pressure (P_c)** is the minimum pressure which must be applied to a gas to liquify it at its critical temperature. It is given by, $P_c=\frac{a}{27b^2}$

(4) **Critical volume (V_c)** is the volume occupied by one mole of the substance at its critical temperature and critical pressure. It is given by, $V_c=3b$

(5) **Critical compressibility factor (Z_c)** is given by, $Z_c=\frac{P_cV_c}{RT_c}=\frac{3}{8}=0.375$

A gas behaves as a Vander Waal's gas if its critical compressibility factor (Z_c) is equal to 0.375. A substance is the gaseous state below T_c is called vapour and above T_c is called gas.

Degrees of freedom of a gaseous molecule

(1) The motion of atoms and molecules is generally described in terms of the degree of freedom which they possess.

(2) The *degrees of freedom* of a molecule are defined as the independent number of parameters

required to describe the state of the molecule completely.

(3) When a gaseous molecule is heated, the energy supplied to it may bring about three kinds of motion in it, these are,

(i) The translational motion (ii) The rotational motion (iii) The vibrational motion.

This is expressed by saying that the molecule possesses translational, rotational and vibrational degrees of freedom.

(4) For a molecule made up of N atoms, total degrees of freedom = 3N. Further split up of these is as follows,

	Translational	Rotational	Vibrational
For linear molecule :	3	2	3N – 5
For non-linear molecule :	3	3	3N – 6

Specific and Molar heat capacity of gases

(1) **Specific heat (or specific heat capacity)** of a substance is the quantity of heat (in *calories*, *joules*, *kcal*, or *kilo joules*) required to raise the temperature of 1*g* of that substance through 1^oC. It can be measured at constant pressure (c_p) and at constant volume (c_v).

(2) **Molar heat capacity** of a substance is the quantity of heat required to raise the temperature of 1 *mole* of the substance by 1^oC.

∴ Molar heat capacity = Specific heat capacity × Molecular weight, *i.e.*,

$$C_v = c_v \times M \text{ and } C_p = c_p \times M .$$

(3) Since gases upon heating show considerable tendency towards expansion if heated under constant pressure conditions, an additional energy has to be supplied for raising its temperature by 1^oC relative to that required under constant volume conditions, *i.e.*,

$C_p > C_v$ or $C_p = C_v + \text{Work done on expansion, } P\Delta V (= R)$

where, $C_p =$ molar heat capacity at constant pressure; $C_v =$ molar heat capacity at constant volume.

(4) **Some useful relations of C_p and C_v**

(i) $C_p - C_v = R = 2\,calories = 8.314\,J$

(ii) $C_v = \frac{3}{2}R$ (for monoatomic gas) and $C_v = \frac{3}{2} + x$ (for *di* and polyatomic gas), where x varies from gas to gas.

(iii) $\frac{C_p}{C_v} = \gamma$ (Ratio of molar capacities)

(iv) For monoatomic gas $C_v = 3\,calories$ whereas, $C_p = C_v + R = 5\,calories$

(v) For monoatomic gas, $(\gamma) = \frac{C_p}{C_v} = \frac{\frac{5}{2}R}{\frac{3}{2}R} = 1.66$.

(vi) For diatomic gas $(\gamma) = \frac{C_p}{C_v} = \frac{\frac{7}{2}R}{\frac{5}{2}R} = 1.40$

(vii) For triatomic gas $(\gamma) = \frac{C_p}{C_v} = \frac{8R}{6R} = 1.33$

Liquefaction of gases

(1) A gas may be liquefied by cooling or by the application of high pressure or by the combined effect of both. The first successful attempt for liquefying gases was made by ***Faraday***.

(2) Gases for which the intermolecular forces of attraction are small such as H_2, N_2, *Ar* and O_2, have low values of T_c and cannot be liquefied by the application of pressure are known as “permanent gases” while the gases for which the intermolecular forces of attraction are large, such as polar molecules NH_3, SO_2 and H_2O have high values of T_c and can be liquefied easily.

(3) **Methods of liquefaction of gases :** The modern methods of cooling the gas to or below their T_c and hence of liquefaction of gases are done by Linde's method and Claude's method.

(i) **Linde's method :** This process is based upon ***Joule-Thomson effect*** which states that “When a gas is allowed to expand adiabatically from a region of high pressure to a region of extremely low pressure, it is accompained by cooling.”

(ii) **Claude's method :** This process is based upon the principle that when a gas expands adiabatically against an external pressure (as a piston in an engine), it does some external work. Since work is done by the molecules at the cost of their kinetic energy, the temperature of the gas falls causing cooling.

(iii) By adiabatic demagnetisation.

(4) **Uses of liquefied gases :** Liquefied and gases compressed under a high pressure are of great importance in industries.

(i) Liquid ammonia and liquid sulphur dioxide are used as refrigerants.

(ii) Liquid carbon dioxide finds use in soda fountains.

(iii) Liquid chlorine is used for bleaching and disinfectant purposes.

(iv) Liquid air is an important source of oxygen in rockets and jet-propelled planes and bombs.

(v) Compressed oxygen is used for welding purposes.

(vi) Compressed helium is used in airships.

(5) **Joule-Thomson effect :** When a real gas is allowed to expand adiabatically through a porous plug or a fine hole into a region of low pressure, it is accompanied by cooling (except for hydrogen and helium which get warmed up).

Cooling takes place because some work is done to overcome the intermolecular forces of attraction. As a result, the internal energy decreases and so does the temperature.

Ideal gases do not show any cooling or heating because there are no intermolecular forces of attraction i.e., they do not show Joule-Thomson effect.

During Joule-Thomson effect, enthalpy of the system remains constant.

Joule-Thomson coefficient. $\mu = (\partial T / \partial P)_H$.

For cooling, $\mu = +ve$ (because dT and dP will be $-ve$)

For heating $\mu = -ve$ (because $dT = +ve, dP = -ve$).

For no heating or cooling $\mu = 0$ (because $dT = 0$).

(6) **Inversion temperature :** It is the temperature at which gas shows neither cooling effect nor heating effect *i.e.,* Joule-Thomson coefficient $\mu = 0$. Below this temperature, it shows cooling effect and above this temperature, it shows heating effect.

Any gas like H_2, He etc, whose inversion temperature is low would show heating effect at room temperature. However, if these gases are just cooled below inversion temperature and then subjected to Joule-Thomson effect, they will also undergo cooling.

Tips & Tricks

- If the number of molecules present in 1 c.c. of the gas or vapour at S.T.P., then that is called loschmidt number. Its value is 2.687×10^{19} per c.c.
- $CO_2 > SO_2 > SO_3 > PCl_3$ is order of rate of diffusion.
- Vapour density is independent of temperature and has no unit while absolute density is dependent of temperature and has unit of gm^{-1}
- The isotherms of CO_2 were first studied by Andrews.
- 1 *Cal* = 4.2 *Joule*, 1 *Kcal* = 4200 *Joule*
- The gas which has least mean free path has maximum value of *a*, is easily liquefied and has maximum value of T_b.
- $T_c < T_b < T_i$
- For critical constants, compression factor Z is < 1.
- Boyle's law and Avogadro's law are applicable under limiting condition. This limiting condition is $P \rightarrow 0$.
- $T_c = 0.296\ T_b$; $T_i = 6.75\ T_c$
- Mean free path increases if H_2 is replaced by *He*.

OT Ordinary Thinking

Objective Questions

Characteristics and Measurable properties of gases

1. Which one of the following statements is not correct about the three states of matter *i.e.* solid, liquid and gaseous

(a) Molecules of a solid possess least energy whereas those of a gas possess highest energy

(b) The density of solid is highest whereas that of gases is lowest

(c) Gases like liquids possess definite volumes

(d) Molecules of a solid possess vibratory motion

2. The temperature and pressure at which ice, liquid water and water vapour can exist together are

(a) $0^oC, 1\,atm$ (b) $2^oC, 4.7\,atm$

(c) $0^oC, 4.7\,mm$ (d) $-2^oC, 4.7\,mm$

3. Which of the following is true about gaseous state

(a) Thermal energy = Molecular attraction

(b) Thermal energy >> Molecular attraction

(c) Thermal energy << Molecular attraction

(d) Molecular forces >> Those in liquids

4. Kinetic energy of molecules is highest in

(a) Gases (b) Solids

(c) Liquids (d) Solutions

5. Which of the following statement is correct

(a) In all the three states the molecules possess random translational motion

(b) Gases cannot be converted into solids without passing through liquid state

(c) One of the common property of liquids and gases is viscosity

(d) According to Boyle's law V/P is constant at constant T

6. A volume of 1 m^3 is equal to

(a) $1000\ cm^3$ (b) $100\ cm^3$

(c) $10\ dm^3$ (d) $10^6\ cm^3$

7. Which one of the following is not a unit of pressure

(a) Newton (b) Torr

(c) Pascal (d) Bar

8. 1^oC rise in temperature is equal to a rise of

(a) 1^oF (b) $9/5^oF$

(c) $5/9^oF$ (d) 33^oF

9. Which of the following relations for expressing volume of a sample is not correct

(a) $1L = 10^3\,ml$ (b) $1\,dm^3 = 1\,L$

(c) $1\,L = 10^3 m^3$ (d) $1L = 10^3 cm^3$

10. One atmosphere is numerically equal to approximately

(a) 10^6 dynes cm^{-2} (b) 10^2 dynes cm^{-2}

(c) 10^4 dynes cm^{-2} (d) 10^8 dynes cm^{-2}

11. 2*gm* of O_2 at 27^oC and 760*mm* of *Hg* pressure has volume **[BCECE 2005]**

(a) 1.5 *lit.* (b) 2.8 *lit.*

(c) 11.2 *lit.* (d) 22.4 *lit.*

12. Pressure of a gas in a vessel can be measured by

(a) Barometer (b) Manometer

(c) Stalgometer (d) All the baove

13. Volume occupied by a gas at one atmospheric pressure and 0^oC is *V mL*. Its volume at 273 *K* will be

[Bihar MADT 1982]

(a) *V ml* (b) *V*/2 *ml*

(c) 2 *V* (d) None of these

14. Which one of the following statements is wrong for gases

[CBSE PMT 1999]

(a) Gases do not have a definite shape and volume

(b) Volume of the gas is equal to the volume of the container confining the gas

(c) Confined gas exerts uniform pressure on the walls of its container in all directions

(d) Mass of the gas cannot be determined by weighing a container in which it is enclosed

15. Which of the following exhibits the weakest intermolecular forces **[AIIMS 2000]**

(a) NH_3 (b) HCl

(c) He (d) H_2O

16. N_2 is found in a litre flask under $100\,kPa$ pressure and O_2 is found in another 3 litre flask under $320\,kPa$ pressure. If the two flasks are connected, the resultant pressures is

[Kerala PMT 2004]

(a) 310 *kPa* (b) 210 *kPa*

(c) 420 *kPa* (d) 365 *kPa*

(e) 265 *kPa*

Ideal gas equation and Related gas laws

1. If *P*, *V*, *T* represent pressure, volume and temperature of the gas, the correct representation of Boyle's law is

[BIT Ranchi 1988]

(a) $V \propto \frac{1}{T}$ (at constant P) (b) $PV = RT$

(c) $V \propto 1/P$ (at constant T) (d) $PV = nRT$

2. At constant temperature, in a given mass of an ideal gas **[CBSE PMT 1991]**

(a) The ratio of pressure and volume always remains constant

(b) Volume always remains constant

(c) Pressure always remains constant

(d) The product of pressure and volume always remains constant

3. Air at sea level is dense. This is a practical application of **[Kerala CEE 2000]**

(a) Boyle's law (b) Charle's law

(c) Avogadro's law (d) Dalton's law

4. If 20 cm^3 gas at 1 *atm*. is expanded to 50 cm^3 at constant *T*, then what is the final pressure**[CPMT 1988]**

(a) $20 \times \frac{1}{50}$ (b) $50 \times \frac{1}{20}$

(c) $1 \times \frac{1}{20} \times 50$ (d) None of these

5. Which of the following statement is false**[BHU 1994]**

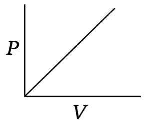
(a) The product of pressure and volume of fixed amount of a gas is independent of temperature

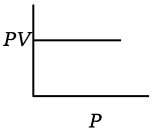
(b) Molecules of different gases have the same K.E. at a given temperature

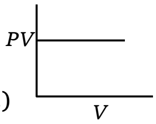
(c) The gas equation is not valid at high pressure and low temperature

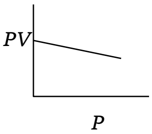
(d) The gas constant per molecule is known as Boltzmann constant

6. Which of the following graphs represent Boyle's law

(a) 

(b) 

(c) 

(d) 

7. Densities of two gases are in the ratio 1 : 2 and their temperatures are in the ratio 2 : 1, then the ratio of their respective pressures is **[BHU 2000]**

(a) 1 : 1 (b) 1 : 2

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

8. At constant pressure, the volume of fixed mass of an ideal gas is directly proportional to**[EAMCET 1985]**

(a) Absolute temperature (b) Degree centigrade

(c) Degree Fahrenheit (d) None

9. Which of the following expression at constant pressure represents Charle's law **[AFMC 1990]**

(a) $V \propto \frac{1}{T}$ (b) $V \propto \frac{1}{T^2}$

(c) $V \propto T$ (d) $V \propto d$

10. Use of hot air balloons in sports and meteorological obsevations is an application of**[Kerala ME**

(a) Boyle's law (b) Newtonic law

(c) Kelvin's law (d) Charle's law

11. A 10 *g* of a gas at atmospheric pressure is cooled from $273^o C$ to $0^o C$ keeping the volume constant, its pressure would become

(a) 1/2 *atm* (b) 1/273 *atm*

(c) 2 *atm* (d) 273 *atm*

12. Pressure remaining the same, the volume of a given mass of an ideal gas increases for every degree centigrade rise in temperature by definite fraction of its volume at **[CBSE PMT 1989]**

(a) $0^o C$ (b) Its critical temperature

(c) Absolute zero (d) Its Boyle temperature

13. A certain sample of gas has a volume of 0.2 *litre* measured at 1 *atm*. pressure and $0^o C$. At the same pressure but at $273^o C$, its volume will be**[EAMCET 1992, 9**

(a) 0.4 *litres* (b) 0.8 *litres*

(c) 27.8 *litres* (d) 55.6 *litres*

14. 400 cm^3 of oxygen at $27^o C$ were cooled to $-3^o C$ without change in pressure. The contraction in volume will be

(a) 40 cm^3 (b) 30 cm^3

(c) 44.4 cm^3 (d) 360 cm^3

15. The pressure *p* of a gas is plotted against its absolute temperature *T* for two different constant volumes, V_1 and V_2. When $V_1 > V_2$, the

(a) Curves have the same slope and do not intersect

(b) Curves must intersect at some point other than $T = 0$

(c) Curve for V_2 has a greater slope than that for V_1

(d) Curve for V_1 has a greater slope than that for V_2

16. Two closed vessels of equal volume containing air at pressure P_1 and temperature T_1 are connected to each other through a narrow tube. If the temperature in one of the vessels is now maintained at T_1 and that in the other at T_2, what will be the pressure in the vessels

(a) $\frac{2P_1T_1}{T_1+T_2}$ (b) $\frac{T_1}{2P_1T_2}$

(c) $\frac{2P_1T_2}{T_1+T_2}$ (d) $\frac{2P_1}{T_1+T_2}$

17. "One gram molecule of a gas at N.T.P. occupies 22.4 *litres*." This fact was derived from **[CPMT 1981, 1995]**

(a) Dalton's theory

(b) Avogadro's hypothesis

(c) Berzelius hypothesis

(d) Law of gaseous volume

18. In a closed flask of 5 *litres*, 1.0 *g* of H_2 is heated from 300 to 600 *K*. which statement is not correct **[CBSE PMT 1991]**

(a) Pressure of the gas increases

(b) The rate of collision increases

(c) The number of moles of gas increases

(d) The energy of gaseous molecules increases

19. Which one of the following statements is false **[Manipal PMT 1991]**

(a) Avogadro number $= 6.02 \times 10^{21}$

(b) The relationship between average velocity $(\bar{v})$ and root mean square velocity (u) is $\bar{v} = 0.9213\, u$

(c) The mean kinetic energy of an ideal gas is independent of the pressure of the gas

(d) The root mean square velocity of the gas can be calculated by the formula $(3RT/M)^{1/2}$

20. The compressibility of a gas is less than unity at STP. Therefore **[IIT 2000]**

(a) $V_m > 22.4$ *litres* (b) $V_m < 22.4$ *litres*

(c) $V_m = 22.4$ *litres* (d) $V_m = 44.8$ *litres*

21. In the equation of sate of an ideal gas $PV = nRT$, the value of the universal gas constant would depend only on **[KCET 2005]**

(a) The nature of the gas

(b) The pressure of the gas

(c) The units of the measurement

(d) None of these

22. In the ideal gas equation, the gas constant *R* has the dimensions of **[NCERT 1982]**

(a) *mole-atm* K^{-1} (b) *litre mole*

(c) *litre-atm* K^{-1} $mole^{-1}$ (d) *erg* K^{-1}

23. In the equation $PV = nRT$, which one cannot be the numerical value of *R* **[BIT 1987]**

(a) $8.31 \times 10^7\, erg\, K^{-1} mol^{-1}$

(b) $8.31 \times 10^7\, dyne\ cm\, K^{-1} mol^{-1}$

(c) $8.31\, JK^{-1} mol^{-1}$

(d) $8.31\, atm.\, K^{-1} mol^{-1}$

24. Which one of the following indicates the value of the gas constant *R* **[EAMCET 1989]**

(a) 1.987 cal K^{-1} mol^{-1} (b) 8.3 cal K^{-1} mol^{-1}

(c) 0.0821 *lit* K^{-1} mol^{-1} (d) 1.987 Joules K^{-1} mol^{-1}

25. The constant *R* is **[Orissa 1990]**

(a) Work done per molecule

(b) Work done per degree absolute

(c) Work done per degree per mole

(d) Work done per mole

26. Select one correct statement. In the gas equation, $PV = nRT$ **[CBSE PMT 1992]**

(a) *n* is the number of molecules of a gas

(b) *V* denotes volume of one mole of the gas

(c) *n* moles of the gas have a volume *V*

(d) *P* is the pressure of the gas when only one mole of gas is present

27. The correct value of the gas constant *R* is close to **[CBSE PMT 1992]**

(a) 0.082 *litre-atmopshere K*

(b) 0.082 *litre-atmosphere* $K^{-1}\, mol^{-1}$

(c) 0.082 *litre-* $atmosphere^{-1}\, K\, mole^{-1}$

(d) 0.082 $litre^{-1}\, atmosphere^{-1}\, K\, mol$

28. S.I. unit of gas constant *R* is **[CPMT 1994]**

(a) 0.0821 *litre atm* K^{-1} $mole^{-1}$

(b) 2 *calories* K^{-1} $mole^{-1}$

(c) 8.31 *joule* K^{-1} $mole^{-1}$

(d) None

29. Gas equation $PV = nRT$ is obeyed by **[BHU 2000]**

(a) Only isothermal process (b)Only adiabatic process

(c) Both (a) and (b) (d) None of these

30. For an ideal gas number of moles per litre in terms of its pressure *P*, gas constant *R* and temperature *T* is **[AIEEE 2002]**

(a) PT/R (b) PRT

(c) P/RT (d) RT/P

31. If two moles of an ideal gas at 546 *K* occupy a volume of 44.8 *litres*, the pressure must be **[NCERT 1981; JIPMER 1991]**

(a) 2 *atm* (b) 3 *atm*

(c) 4 *atm* (d) 1 *atm*

32. How many moles of *He* gas occupy 22.4 *litres* at 30^oC and one atmospheric pressure **[KCET 1992]**

(a) 0.90 (b) 1.11

(c) 0.11 (d) 1.0

33. Volume of 0.5 *mole* of a gas at 1 *atm*. pressure and $273\,K$ is **[EAMCET 1992]**

(a) 22.4 *litres* (b) 11.2 *litres*

(c) 44.8 *litres* (d) 5.6 *litres*

34. At 0^oC and one *atm* pressure, a gas occupies 100 *cc*. If the pressure is increased to one and a half-time and temperature is increased by one-third of absolute temperature, then final volume of the gas will be

[DCE 2000]

(a) 80 *cc* (b) 88.9 *cc*
(c) 66.7 *cc* (d) 100 *cc*

35. Correct gas equation is **[CBSE PMT 1989; CPMT 1991]**

(a) $\frac{V_1T_2}{P_1} = \frac{V_2T_1}{P_2}$ (b) $\frac{P_1V_1}{P_2V_2} = \frac{T_1}{T_2}$
(c) $\frac{P_1T_2}{V_1} = \frac{P_2V_2}{T_2}$ (d) $\frac{V_1V_2}{T_1T_2} = P_1P_2$

36. Two separate bulbs contain ideal gases *A* and *B*. The density of gas *A* is twice that of gas *B*. The molecular mass of *A* is half that of gas *B*. The two gases are at the same temperature. The ratio of the pressure of *A* to that of gas *B* is

[BHU 1994]

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

37. 16 *g* of oxygen and 3 *g* of hydrogen are mixed and kept at 760 *mm* pressure and 0^oC. The total volume occupied by the mixture will be nearly **[Vellore CMC 1991]**

(a) 22.4 *litres* (b) 33.6 *litres*
(c) 448 *litres* (d) 44800 *ml*

38. Pure hydrogen sulphide is stored in a tank of 100 *litre* capacity at 20^oC and 2 *atm* pressure. The mass of the gas will be **[CPMT 1989]**

(a) 34 *g* (b) 340 *g*
(c) 282.4 *g* (d) 28.24 *g*

39. At N.T.P. the volume of a gas is found to be 273 *ml*. What will be the volume of this gas at 600 *mm Hg* and 273^oC

[CPMT 1992]

(a) 391.8 *mL* (b) 380 *ml*
(c) 691.6 *ml* (d) 750 *ml*

40. One *litre* of a gas weighs 2 *g* at 300 *K* and 1 *atm* pressure. If the pressure is made 0.75 *atm*, at which of the following temperatures will one *litre* of the same gas weigh one *gram*

[CBSE PMT 1992]

(a) 450 *K* (b) 600 *K*
(c) 800 *K* (d) 900 *K*

41. A wheather balloon filled with hydrogen at 1 *atm* and 27^oC has volume equal to 12000 *litres*. On ascending it reaches a place where the temperature is -23^oC and pressure is 0.5 *atm*. The volume of the balloon is

[CBSE PMT 1991]

(a) 24000 *litres* (b) 20000 *litres*
(c) 10000 *litres* (d) 12000 *litres*

42. The density of a gas at 27^oC and 1 *atm* is *d*. Pressure remaining constant at which of the following temperatures will its density become 0.75 *d* **[CBSE PMT 1992]**

(a) 20^oC (b) 30^oC
(c) 400 *K* (d) 300 *K*

43. A sample of gas occupies 100 *ml* at 27^oC and 740 *mm* pressure. When its volume is changed to 80 *ml* at 740 *mm* pressure, the temperature of the gas will be

[Vellore CMC 1991]

(a) 21.6^oC (b) 240^oC
(c) -33^oC (d) 89.5^oC

44. The total pressure exerted by a number of non-reacting gases is equal to the sum of the partial pressures of the gases under the same conditions is known as **[CPMT 1986]**

(a) Boyle's law (b) Charle's law
(c) Avogadro's law (d) Dalton's law

45. “Equal volumes of all gases at the same temperature and pressure contain equal number of particles.” This statement is a direct consequence of **[Kerala MEE 2002]**

(a) Avogadro’s law (b) Charle's law
(c) Ideal gas equation (d) Law of partial pressure

46. If three unreactive gases having partial pressures P_A, P_B and P_C and their moles are 1, 2 and 3 respectively then their total pressure will be **[CPMT 1994]**

(a) $P = P_A + P_B + P_C$ (b) $P = \frac{P_A + P_B + P_C}{6}$
(c) $P = \frac{\sqrt{P_A + P_B + P_C}}{3}$ (d) None

47. Dalton's law of partial pressure will not apply to which of the following mixture of gases **[Bihar MADT 1981]**

(a) H_2 and SO_2 (b) H_2 and Cl_2
(c) H_2 and CO_2 (d) CO_2 and Cl_2

48. Which of the following mixtures of gases does not obey Dalton's law of partial pressure

[CBSE PMT 1996: Kerala PMT 2000]

(a) O_2 and CO_2 (b) N_2 and O_2
(c) Cl_2 and O_2 (d) NH_3 and HCl

49. To which of the following gaseous mixtures is Dalton's law not applicable

(a) $Ne + He + SO_2$ (b) $NH_3 + HCl + HBr$
(c) $O_2 + N_2 + CO_2$ (d) $N_2 + H_2 + O_2$

50. Equal amounts of two gases of molecular weight 4 and 40 are mixed. The pressure of the mixture is 1.1 *atm*. The partial pressure of the light gas in this mixture is

[CBSE PMT 1991]

(a) 0.55 *atm* (b) 0.11 *atm*
(c) 1 *atm* (d) 0.12 *atm*

51. Rate of diffusion of a gas is **[IIT 1985; CPMT 1987]**

(a) Directly proportional to its density
(b) Directly proportional to its molecular mass
(c) Directly proportional to the square root of its molecular mass
(d) Inversely proportional to the square root of its molecular mass

52. Which of the following gas will have highest rate of diffusion

[Pb. CET Sample paper 1993; CPMT 1990]

(a) NH_3 (b) N_2

(c) CO_2 (d) O_2

53. Which of the following relationship is correct, where r is the rate of diffusion of a gas and d is its density **[CPMT 1994]**

(a) $r \propto \sqrt{1/d}$ (b) $r \propto \sqrt{d}$

(c) $r = d$ (d) $r \propto d$

54. According to Grahman's law at a given temperature, the ratio of the rates of diffusion r_A / r_B of gases A and B is given by **[IIT 1998]**

(a) $(P_A / P_B)(M_A / M_B)^{1/2}$

(b) $(M_A / M_B)(P_A / P_B)^{1/2}$

(c) $(P_A / P_B)(M_B / M_A)^{1/2}$

(d) $(M_A / M_B)(P_B / P_A)^{1/2}$

(where P and M are the pressures and molecular weights of gases A and B respectively)

55. The ratio of the rate of diffusion of a given element to that of helium is 1.4. The molecular weight of the element is

[Kerala PMT 1990]

(a) 2 (b) 4

(c) 8 (d) 16

56. A gas diffuse 1/5 times as fast as hydrogen. Its molecular weight is **[CPMT 1992; Bihar CEE 1982]**

(a) 50 (b) 25

(c) $25\sqrt{2}$ (d) $50\sqrt{2}$

57. The molecular weight of a gas which diffuses through a porous plug at 1/6th of the speed of hydrogen under identical conditions is **[EAMCET 1990]**

(a) 27 (b) 72

(c) 36 (d) 48

58. Molecular weight of a gas that diffuses twice as rapidly as the gas with molecular weight 64 is **[EAMCET 1994]**

(a) 16 (b) 8

(c) 64 (d) 6.4

59. The densities of hydrogen and oxygen are 0.09 and 1.44 $g\ L^{-1}$. If the rate of diffusion of hydrogen is 1 then that of oxygen in the same units will be **[RPMT 1994]**

(a) 4 (b) 1/4

(c) 16 (d) 1/16

60. If rate of diffusion of A is 5 times that of B, what will be the density ratio of A and B **[AFMC 1994]**

(a) 1/25 (b) 1/5

(c) 25 (d) 4

61. The densities of two gases are in the ratio of 1 : 16. The ratio of their rates of diffusion is **[CPMT 1995]**

(a) 16 : 1 (b) 4 : 1

(c) 1 : 4 (d) 1 : 16

62. At constant volume and temperature conditions, the rate of diffusion D_A and D_B of gases A and B having densities ρ_A and ρ_B are related by the expression **[IIT 1993]**

(a) $D_A = \left[D_B \cdot \frac{\rho_A}{\rho_B}\right]^{1/2}$ (b) $D_A = \left[D_B \cdot \frac{\rho_B}{\rho_A}\right]^{1/2}$

(c) $D_A = D_B\left(\frac{\rho_A}{\rho_B}\right)^{1/2}$ (d) $D_A = D_B\left(\frac{\rho_B}{\rho_A}\right)^{1/2}$

63. Atmolysis is a process of

(a) Atomising gas molecules

(b) The breaking of atoms to sub-atomic particles

(c) Separation of gases from their gaseous mixture

(d) Changing of liquids to their vapour state

64. A bottle of ammonia and a bottle of dry hydrogen chloride connected through a long tube are opened simultaneously at both ends, the white ammonium chloride ring first formed will be **[IIT 1988]**

(a) At the centre of the tube

(b) Near the hydrogen chloride bottle

(c) Near the ammonia bottle

(d) Throughout the length of the tube

65. Which of the following pairs will diffuse at the same rate through a porous plug **[EAMCET 1990]**

(a) $CO,\ NO_2$ (b) NO_2, CO_2

(c) NH_3, PH_3 (d) $NO\ C_2H_6$

66. If 4 *g* of oxygen diffuse through a very narrow hole, how much hydrogen would have diffused under identical conditions **[CPMT 1971]**

(a) 16 *g* (b) 1 *g*

(c) 1/4 *g* (d) 64 *g*

67. A gas diffuse at a rate which is twice that of another gas B. The ratio of molecular weights of A to B is **[EAMCET 1986]**

(a) 1.0 (b) 0.75

(c) 0.50 (d) 0.25

68. Two grams of hydrogen diffuse from a container in 10 *minutes*. How many *grams* of oxygen would diffuse through the same container in the same time under similar conditions **[MNR 1980]**

(a) 0.5 *g* (b) 4 *g*

(c) 6 *g* (d) 8 *g*

69. The rate of diffusion of methane at a given temperature is twice that of X. The molecular weight of X is

[MNR 1995; Kerala CEE 2001]

(a) 64.0 (b) 32.0

(c) 40.0 (d) 80.0

70. X *ml* of H_2 gas effuses through a hole in a container in 5 *seconds*. The time taken for the effusion of the same volume of the gas specified below under identical condition is

[IIT 1996]

(a) 10 *seconds* : *He* (b) 20 *seconds* : O_2
(c) 25 *seconds* : *CO* (d) 55 *seconds* : CO_2

71. At what temperature, the rate of effusion of N_2 would be 1.625 times that of SO_2 at 50^oC **[CBSE PMT 1996]**

(a) 110 K (b) 173 K
(c) 373 K (d) 273 K

72. Given the reaction $C(s)+H_2O(l)\rightarrow CO(g)+H_2(g)$ calculate the volume of the gases produced at STP from 48.0 *g* of carbon

(a) 179.2 L (b) 89.6 L
(c) 44.8 L (d) 22.4 L

73. 4.4 *g* of a gas at STP occupies a volume of 2.24 *L*, the gas can be **[Haryana CEET 2000]**

(a) O_2 (b) CO
(c) NO_2 (d) CO_2

74. Under what conditions will a pure sample of an ideal gas not only exhibit a pressure of 1 *atm* but also a concentration of 1 *mole* $litre-1$

$(R = 0.082\ litre\ atm\ mol^{-1}\ \deg^{-1})$ **[CBSE PMT 1993]**

(a) At STP
(b) When $V = 22.4\ litres$
(c) When $T = 12\ K$
(d) Impossible under any conditions

75. There are 6.02×10^{22} molecules each of N_2, O_2 and H_2 which are mixed together at 760 *mm* and 273 *K*. The mass of the mixture in grams is **[Pb. PMT 1997]**

(a) 6.2 (b) 4.12
(c) 3.09 (d) 7

76. Volume of 4.4 *g* of CO_2 at NTP is **[Pb. CET 1997]**

(a) 22.4 L (b) 44.8 L
(c) 2.24 L (d) 4.48 L

77. The energy of an ideal gas depends only on its

(a) Pressure (b) Volume
(c) Number of moles (d) Temperature

78. A bottle of cold drink contains 200 *ml* liquid in which CO_2 is 0.1 molar. Suppose CO_2 behaves like an ideal gas, the volume of the dissolved CO_2 at STP is **[CBSE PMT 1991]**

(a) 0.224 *litre* (b) 0.448 *litre*
(c) 22.4 *litre* (d) 2.24 *litre*

79. The vapour density of a gas is 11.2. The volume occupied by 11.2 *g* of this gas at N.T.P. is **[MNR 1982; CBSE PMT 1991]**

(a) 1 L (b) 11.2 L
(c) 22.4 L (d) 20 L

80. A pre-weighed vessel was filled with oxygen at N.T.P. and weighted. It was then evacuated, filled with SO_2 at the same temperature and pressure, and again weighted. The weight of oxygen will be **[NCERT 1989]**

(a) The same as that of SO_2
(b) $\frac{1}{2}$ that of SO_2
(c) Twice that of SO_2
(d) One fourth that of SO_2

81. Five grams each of the following gases at 87^oC and 750 *mm* pressure are taken. Which of them will have the least volume **[MNR 1991]**

(a) *HF* (b) *HCl*
(c) *HBr* (d) *HI*

82. Who among the following scientists has not done any important work on gases **[Bihar MADT 1980]**

(a) Boyle (b) Charles
(c) Avogadro (d) Faraday

83. A closed vessel contains equal number of nitrogen and oxygen molecules at a pressure of *P mm*. If nitrogen is removed from the system then the pressure will be

[MP PMT 1985]

(a) P (b) $2P$
(c) $P/2$ (d) P^2

84. If the four tubes of a car are filled to the same pressure with N_2, O_2, H_2 and *Ne* separately, then which one will be filled first **[Manipal PMT 2001]**

(a) N_2 (b) O_2
(c) H_2 (d) *Ne*

85. Which of the following gas mixture is not applicable for Dalton's law of partial pressure **[Pb. CET 20**

(a) SO_2 and Cl_2 (b) CO_2 and N_2
(c) CO and CO_2 (d) CO and N_2

86. At what pressure a quantity of gas will occupy a volume of $60\,ml$, if it occupies a volume of $100\,ml$ at a pressure of $720\,mm$? (while temperature is constant) : **[Pb. CET 2000]**

(a) $700\ mm$ (b) $800\ mm$
(c) $100\ mm$ (d) $1200\ mm$

87. At constant temperature and pressure which gas will diffuse first H_2 or O_2 ? **[Pb. CET 2000]**

(a) Hydrogen
(b) Oxygen
(c) Both will diffuse in same time
(d) None of the above

88. When a jar containing gaseous mixture of equal volumes of CO_2 and H_2 is placed in a solution of sodium hydroxide, the solution level will **[Pb. CET 2001]**

(a) Rise (b) Fall
(c) Remain constant (d) Become zero

89. At S.T.P. $1g\,CaCO_3$ on decomposition gives CO_2

[Pb. CET 2000]

(a) 22.4 litre (b) 2.24 litre
(c) 0.224 litre (d) 11.2 litre

90. At NTP, the density of a gas, whose molecular weight is 45 is **[Pb. CET 2001, 03]**
(a) 44.8 gm/litre (b) 11.4 gm/litre
(c) 2 gm/ litre (d) 3 gm/litre

91. What is the ratio of diffusion rate of oxygen and hydrogen **[Pb. CET 2003]**
(a) 1 : 4 (b) 4 : 1
(c) 1 : 8 (d) 8 : 1

92. The maximum number of molecules is present in **[CBSE PMT 2004]**
(a) 0.5 g of H_2 gas (b) 10 g of O_2 gas
(c) 15 L of H_2 gas at STP(d) 5 L of N_2 gas at STP

93. One litre oxygen gas at STP will weigh**[Pb. CET 2004]**
(a) 1.43 g (b) 2.24 g
(c) 11.2 g (d) 22.4 g

94. How will you separate mixture of two gases**[AFMC 2004]**
(a) Fractional distillation technique
(b) Grahams law of diffusion technique
(c) Osmosis
(d) Chromatography

95. The rate of diffusion of hydrogen gas is **[MH CET 2003; Pb. CET 2000]**
(a) 1.4 times to He gas (b) Same as He gas
(c) 5 times to He gas (d) 2 times to He gas

96. Hydrogen diffuses six times faster than gas A. The molar mass of gas A is **[KCET 2004]**
(a) 72 (b) 6
(c) 24 (d) 36

97. At what pressure will a quantity of gas, which occupies $100\,ml$ at a pressue of $720\,mm$, occupy a volume of $84\,ml$ **[DPMT 2004]**
(a) 736.18 mm (b) 820.20 mm
(c) 784.15 mm (d) 857.14 mm

98. Containers A and B have same gases. Pressure, volume and temperature of A are all twice that of B, then the ratio of number of molecules of A and B are **[AFMC 2004]**
(a) 1 : 2 (b) 2
(c) 1 : 4 (d) 4

99. A mixture of NO_2 and N_2O_4 has a vapour density of 38.3 at $300\,K$. What is the number of moles of NO_2 in $100\,g$ of the mixture **[Kerala PMT 2004]**
(a) 0.043 (b) 4.4
(c) 3.4 (d) 3.86
(e) 0.437

100. A cylinder of 5 litres capacity, filled with air at NTP is connected with another evacuated cylinder of 30 litres of capacity. The resultant air pressure in both the cylinders will be **[BHU 2004]**
(a) 10.8 cm of Hg (b) 14.9cm of Hg
(c) 21.8 cm of Hg (d) 38.8 cm of Hg

101. A certain mass of gas occupies a volume of 300 c.c. at 27C and 620 mm pressure. The volume of this gas at 47^oC and 640 mm pressure will be **[MH CET 20**
(a) 400 c.c. (b) 510 c.c.
(c) 310 c.c. (d) 350 c.c.

102. What will be the volume of the mixture after the reaction?
$\underset{4\text{ litre}}{NH_3} + \underset{1.5\text{ litre}}{HCl} \rightarrow \underset{\text{(solid)}}{NH_4Cl}$ **[BVP 2004]**
(a) 0.5 litre (b) 1 litre
(c) 2.5 litre (d) 0.1 litre

103. The pressure and temperature of $4\,dm^3$ of carbon dioxide gas are doubled. Then the volume of carbon dioxide gas would be **[KCET 2004]**
(a) $2\,dm^3$ (b) $3dm^3$
(c) $4\,dm^3$ (d) $8\,dm^3$

104. If the absolute temperature of an ideal gas become double and pressure become half, the volume of gas would be **[Kerala CET 2005]**
(a) Remain unchange (b) Will be double
(c) Will be four time (d) will be half
(e) Will be one fourth

105. At what temperature, the sample of neon gas would be heated to double of its pressure, if the initial volume of gas is/are reduced to 15% at 75^oC **[Kerala CET 2005]**
(a) 319^oC (b) 592^oC
(c) 128^oC (d) 60^oC
(e) 90^oC

106. Equation of Boyle's law is **[DPMT 2005]**
(a) $\frac{dP}{p} = -\frac{dV}{V}$ (b) $\frac{dP}{P} = +\frac{dV}{V}$
(c) $\frac{d^2P}{P} = -\frac{dV}{dT}$ (d) $\frac{d^2P}{P} = +\frac{d^2V}{dT}$

Kinetic molecular theory of gases and Molecular collisions

1. Postulate of kinetic theory is **[EAMCET 1980]**
(a) Atom is indivisible
(b) Gases combine in a simple ratio
(c) There is no influence of gravity on the molecules of a gas
(d) None of the above

2. According to kinetic theory of gases,**[EAMCET 1980]**
(a) There are intermolecular attractions
(b) Molecules have considerable volume
(c) No intermolecular attractions
(d) The velocity of molecules decreases after each collision

3. In deriving the kinetic gas equation, use is made of the root mean square velocity of the molecules because it is **[Bihar MADT 1980]**

(a) The average velocity of the molecules

(b) The most probable velocity of the molecules

(c) The square root of the average square velocity of the molecules

(d) The most accurate form in which velocity can be used in these calculations

4. Kinetic energy of a gas depends upon its **[Bihar MADT 1982]**

(a) Molecular mass (b) Atomic mass

(c) Equivalent mass (d) None of these

5. The kinetic theory of gases perdicts that total kinetic energy of a gaseous assembly depends on **[NCERT 1984]**

(a) Pressure of the gas

(b) Temperature of the gas

(c) Volume of the gas

(d) Pressure, volume and temperature of the gas

6. According to kinetic theory of gases, the energy per mole of a gas is equal to **[EAMCET 1985]**

(a) 1.5 RT (b) RT

(c) 0.5 RT (d) 2.5 RT

7. Internal energy and pressure of a gas per unit volume are related as **[CBSE PMT 1993]**

(a) $P = \frac{2}{3}E$ (b) $P = \frac{3}{2}E$

(c) $P = \frac{1}{2}E$ (d) $P = 2E$

8. The translational kinetic energy of an ideal gas depends only on its

(a) Pressure (b) Force

(c) Temperature (d) Molar mass

9. Helium atom is two times heavier than a hydrogen molecule at 298 K, the average kinetic energy of helium is **[IIT 1982]**

(a) Two times that of a hydrogen molecule

(b) Same as that of a hydrogen molecule

(c) Four times that of a hydrogen molecule

(d) Half that of a hydrogen molecule

10. Which of the following is valid at absolute zero **[Pb. CET 1985]**

(a) Kinetic energy of the gas becomes zero but the molecular motion does not become zero

(b) Kinetic energy of the gas becomes zero and molecular motion also becomes zero

(c) Kinetic energy of the gas decreases but does not become zero

(d) None of the above

11. The average $K.E.$ of an ideal gas in *calories* per mole is approximately equal to **[EAMCET 1989]**

(a) Three times the absolute temperature

(b) Absolute temperature

(c) Two times the absolute temperature

(d) 1.5 times the absolute temperature

12. According to kinetic theory of gases, for a diatomic molecule **[MNR 1991]**

(a) The pressure exerted by the gas is proportional to the mean velocity of the molecules

(b) The pressure exerted by the gas is proportional to the root mean square velocity of the molecules

(c) The root mean square velocity is inversely proportional to the temperature

(d) The mean translational kinetic energy of the molecules is proportional to the absolute temperature

13. At STP, 0.50 *mol* H_2 gas and 1.0 *mol He* gas **[CBSE PMT 1993, 2000]**

(a) Have equal average kinetic energies

(b) Have equal molecular speeds

(c) Occupy equal volumes

(d) Have equal effusion rates

14. Which of the following expressions correctly represents the relationship between the average molar kinetic energy, $\overline{K.E.}$, of CO and N_2 molecules at the same temperature **[CBSE PMT 2000]**

(a) $\overline{KE}_{CO} = \overline{KE}_{N_2}$

(b) $\overline{KE}_{CO} > \overline{KE}_{N_2}$

(c) $\overline{KE}_{CO} < \overline{KE}_{N_2}$

(d) Cannot be predicted unless the volumes of the gases are given

15. Indicate the correct statement for a 1-L sample of $N_2(g)$ and $CO_2(g)$ at 298 K and 1 *atm* pressure

(a) The average translational KE per molecule is the same in N_2 and CO_2

(b) The *rms* speed remains constant for both N_2 and CO_2

(c) The density of N_2 is less than that of CO_2

(d) The total translational KE of both N_2 and CO_2 is the same

16. With increase of pressure, the mean free path **[Pb. CET 1985]**

(a) Decreases (b) Increases

(c) Does not change (d) Becomes zero

17. Which one of the following statements is NOT true about the effect of an increase in temperature on the distribution of molecular speeds in a gas **[AIEEE 2005]**

(a) The most probable speed increases

(b) The fraction of the molecules with the most probable speed increases

(c) The distribution becomes broader

(d) The area under the distribution curve remains the same as under the lower temperature

18. If P, V, M, T and R are pressure, volume, molar mass, temperature and gas constant respectively, then for an ideal gas, the density is given by **[CBSE PMT 1989, 91]**

(a) $\frac{RT}{PM}$ (b) $\frac{P}{RT}$

(c) $\frac{M}{V}$ (d) $\frac{PM}{RT}$

19. An ideal gas will have maximum density when **[CPMT 2000]**

(a) $P = 0.5\, atm, T = 600\ K$

(b) $P = 2\, atm, T = 150\ K$

(c) $P = 1\, atm, T = 300\ K$

(d) $P = 1.0\, atm, T = 500\ K$

20. If the inversion temperature of a gas is -80^oC, then it will produce cooling under Joule-Thomson effect at

(a) 298 K (b) 273 K

(c) 193 K (d) 173 K

21. Ratio of C_p and C_v of a gas 'X' is 1.4. The number of atoms of the gas 'X' present in 11.2 *litres* of it at N.T.P. is **[CBSE PMT 1989]**

(a) 6.02×10^{23} (b) 1.2×10^{24}

(c) 3.01×10^{23} (d) 2.01×10^{23}

22. The density of air is 0.00130 g/ml. The vapour density of air will be **[DCE 2000]**

(a) 0.00065 (b) 0.65

(c) 14.4816 (d) 14.56

23. At 100^oC and 1 atm, if the density of liquid water is 1.0 $g\ cm^{-3}$ and that of water vapour is 0.0006 $g\ m^{-3}$, then the volume occupied by water molecules in 1 *litre* of steam at that temperature is **[IIT 2000]**

(a) 6 cm^3 (b) 60 cm^3

(c) 0.6 cm^3 (d) 0.06 cm^3

24. The ratio γ for inert gases is **[AFMC 1990]**

(a) 1.33 (b) 1.66

(c) 2.13 (d) 1.99

25. The density of neon will be highest at **[CBSE PMT 1990]**

(a) S.T.P. (b) $0^oC, 2\, atm$

(c) $273^oC, 1\, atm$ (d) $273^oC, 2\, atm$

26. Absolute zero is defined as the temperature **[CBSE PMT 1990]**

(a) At which all molecular motion ceases

(b) At which liquid helium boils

(c) At which ether boils

(d) All of the above

27. Consider the following statements :

(1) Joule-Thomson experiment is isoenthalpic as well as adiabatic.

(2) A negative value of μ_{JT} (Joule Thomson coefficient corresponds to warming of a gas on expansion.

(3) The temperature at which neither cooling nor heating effect is observed is known as inversion temperature.

Which of the above statements are correct

(a) 1 and 2 (b) 1 and 3

(c) 2 and 3 (d) 1, 2 and 3

28. Vibrational energy is **[Pb. CET 1985]**

(a) Partially potential and partially kinetic

(b) Only potential

(c) Only kinetic

(d) None of the above

29. At the same temperature and pressure, which of the following gases will have the highest kinetic energy per mole **[MNR 1991]**

(a) Hydrogen (b) Oxygen

(c) Methane (d) All the same

30. Dimensions of pressure are the same as that of **[CBSE PMT 1995]**

(a) Energy (b) Force

(c) Energy per unit volume (d)Force per unit volume

31. The density of a gas An is three times that of a gas B. if the molecular mass of A is M, the molecular mass of B is **[CPMT 1987]**

(a) 3 M (b) $\sqrt{3}\ M$

(c) $M/3$ (d) $M/\sqrt{3}$

Molecular speeds

1. The ratio of root mean square velocity to average velocity of gas molecules at a particular temperature is **[IIT 1981]**

(a) 1.086 : 1 (b) 1 : 1.086

(c) 2 : 1.086 (d) 1.086 : 2

2. Which is not true in case of an ideal gas **[CBSE PMT 1991]**

(a) It cannot be converted into a liquid

(b) There is no interaction between the molecules

(c) All molecules of the gas move with same speed

(d) At a given temperature, PV is proportional to the amount of the gas

3. The ratio among most probable velocity, mean velocity and root mean square velocity is given by **[CBSE P**

(a) 1 : 2 : 3 (b) $1:\sqrt{2}:\sqrt{3}$

(c) $\sqrt{2}:\sqrt{3}:\sqrt{8/\pi}$ (d) $\sqrt{2}:\sqrt{8/\pi}:\sqrt{3}$

4. Which of the following has maximum root mean square velocity at the same temperature **[Manipal PMT 200**

(a) SO_2 (b) CO_2

(c) O_2 (d) H_2

5. The temperature at which RMS velocity of SO_2 molecules is half that of He molecules at 300 K is **[NTSE 1991]**

(a) 150 K (b) 600 K
(c) 900 K (d) 1200 K

6. At 27^oC, the ratio of rms velocities of ozone to oxygen is **[EAMCET 1992]**

(a) $\sqrt{3/5}$ (b) $\sqrt{4/3}$
(c) $\sqrt{2/3}$ (d) 0.25

7. The average kinetic energy of an ideal gas per molecule in SI units at 25^oC will be **[CBSE PMT 1996]**

(a) $6.17\times10^{-21}\,kJ$ (b) $6.17\times10^{-21}\,J$
(c) $6.17\times10^{-20}\,J$ (d) $7.16\times10^{-20}\,J$

8. At what temperature the RMS velocity of SO_2 be same as that of O_2 at 303 K **[KCET 2001]**

(a) 273 K (b) 606 K
(c) 303 K (d) 403 K

9. Among the following gases which one has the lowest root mean square velocity at 25^oC **[EAMCET 1983]**

(a) SO_2 (b) N_2
(c) O_2 (d) Cl_2

10. The root mean square velocity of an ideal gas in a closed container of fixed volume is increased from $5\times10^4\,cm\,s^{-1}$ to $10\times10^4\,cm\,s^{-1}$. Which of the following statement correctly explains how the change is accomplished **[Pb. CET 1986]**

(a) By heating the gas, the temperature is doubled
(b) By heating the gas, the pressure is quadrupled (*i.e.* made four times)
(c) By heating the gas, the temperature is quadrupled
(d) By heating the gas, the pressure is doubled

11. The rms velocity at NTP of the species can be calculated from the expression **[EAMCET 1990]**

(a) $\sqrt{\frac{3P}{d}}$ (b) $\sqrt{\frac{3\,PV}{M}}$
(c) $\sqrt{\frac{3\,RT}{M}}$ (d) All the above

12. Root mean square velocity of a gas molecule is proportional to **[CBSE PMT 1990]**

(a) $m^{1/2}$ (b) m^0
(c) $m^{-1/2}$ (d) m

13. At constant volume, for a fixed number of moles of a gas, the pressure of the gas increases with increase in temperature due to **[IIT 1992]**

(a) Increase in the average molecular speed
(b) Increased rate of collision amongst molecules
(c) Increase in molecular attraction
(d) Decrease in mean free path

14. Molecular velocities of the two gases at the same temperature are u_1 and u_2. Their masses are m_1 and m_2 respectively. Which of the following expressions is correct **[BHU 1994]**

(a) $\frac{m_1}{u_1^2}=\frac{m_2}{u_2^2}$ (b) $m_1u_1=m_2u_2$
(c) $\frac{m_1}{u_1}=\frac{m_2}{u_2}$ (d) $m_1u_1^2=m_2u_2^2$

15. The temperature of the gas is raised from 27^oC to 927^oC, the root mean square velocity is **[CBSE PMT 1994**

(a) $\sqrt{927/27}$ times the earlier value
(b) Same as before
(c) Halved
(d) Doubled

16. The ratio between the root mean square velocity of H_2 at 50 K and that of O_2 at 800 K is **[IIT 1996]**

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

17. The root mean square velocity of an ideal gas at constant pressure varies density (d) as **[IIT 2000]**

(a) d^2 (b) d
(c) $\sqrt{d}$ (d) $1/\sqrt{d}$

18. Consider a mixture of SO_2 and O_2 kept at room temperature. Compared to the oxygen molecule, the SO_2 molecule will hit the wall with

(a) Smaller average speed (b)Greater average speed
(c) Greater kinetic energy (d) Greater mass

19. The rms speed of N_2 molecules in a gas is u. If the temperature is doubled and the nitrogen molecules dissociate into nitrogen atoms, the rms speed becomes

(a) $u/2$ (b) $2u$
(c) $4u$ (d) $14u$

20. Choose the correct arrangement, where the symbols have their usual meanings

(a) $\bar{u}>u_p>u_{rms}$ (b) $u_{rms}>\bar{u}>u_p$
(c) $u_p>\bar{u}>u_{rms}$ (d) $u_p>u_{rms}>\bar{u}$

21. The ratio of most probable velocity to that of average velocity is **[JEE Orissa 2004]**

(a) $\pi/2$ (b) $2/\pi$
(c) $\sqrt{\pi}/2$ (d) $2/\sqrt{\pi}$

22. The r.m.s. velocity of a certain gas is v at $300\,K$. The temperature, at which the r.m.s. velocity becomes double **[Pb. CET 2002]**

(a) $1200\,K$ (b) $900\,K$
(c) $600\,K$ (d) $150\,K$

23. The r.m.s. velocity of a gas depends upon **[DCE 2002]**

(a) Temperature only
(b) Molecular mass only
(c) Temperature and molecular mass of gas
(d) None of these

24. What is the pressure of 2 mole of NH_3 at 27^oC when its volume is 5 litre in vander Waal's equation (a = 4.17, b = 0.03711) **[JEE Orissa 2004]**
(a) 10.33 atm (b) 9.33 atm
(c) 9.74 atm (d) 9.2 atm

25. The root mean square velocity of one mole of a monoatomic having molar mass M is U_{rms}. The relation between the average kinetic energy (E) of the U_{rms} is

[IIT-JEE Screening 2004]

(a) $U_{rms}=\sqrt{\frac{3E}{2M}}$ (b) $U_{rms}=\sqrt{\frac{2E}{3M}}$
(c) $U_{rms}=\sqrt{\frac{2E}{M}}$ (d) $U_{rms}=\sqrt{\frac{E}{3M}}$

26. Ratio of average to most probable velocity is
[Orissa JEE 2005]
(a) 1.128 (b) 1.224
(c) 1.0 (d) 1.112

27. If the v_{rms} is $30R^{1/2}$ at 27^oC then calculate the molar mass of gas in kilogram. **[DPMT 2005]**
(a) 1 (b) 2
(c) 4 (d) 0.001

Real gases and Vander waal's equation

1. The Vander Waal's equation explains the behaviour of

[DPMT 1981]

(a) Ideal gases (b) Real gases
(c) Vapour (d) Non-real gases

2. Gases deviate from the ideal gas behaviour because their molecules **[NCERT 1981]**
(a) Possess negligible volume
(b) Have forces of attraction between them
(c) Are polyatomic
(d) Are not attracted to one another

3. The compressibility factor of a gas is defined as $Z = PV/RT$. The compressibility factor of ideal gas is

[Pb. CET 1986]

(a) 0 (b) Infinity
(c) 1 (d) –1

4. In Vander Waal's equation of state for a non-ideal gas, the term that accounts for intermolecular forces is

[CBSE PMT 1990; IIT 1988]

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

5. Vander Waal's equation of state is obeyed by real gases. For n moles of a real gas, the expression will be

[IIT 1992; Pb. CET 1986; DPMT 1986]

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

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

(c) $\left(P+\frac{na}{V^2}\right)(nV-b)=nRT$

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

6. Any gas shows maximum deviation from ideal gas at

[CPMT 1991]

(a) 0^oC and 1 atmospheric pressure
(b) 100^oC and 2 atmospheric pressure
(c) -100^oC and 5 atmospheric pressure
(d) 500^oC and 1 atmospheric pressure

7. The temperature at which the second virial coefficient of real gas is zero is called **[AFMC 1993]**
(a) Critical temperature
(b) Eutetic point
(c) Boiling point
(d) Boyle's temperature

8. When is deviation more in the behaviour of a gas from the ideal gas equation $PV = nRT$
[DPMT 1981; NCERT 1982; CBSE PMT 1993]
(a) At high temperature and low pressure
(b) At low temperature and high pressure
(c) At high temperature and high pressure
(d) At low temperature and low high pressure

9. Vander Waal's constants 'a' and 'b' are related with..... respectively **[RPMT 1994]**
(a) Attractive force and bond energy of molecules
(b) Volume and repulsive force of molecules
(c) Shape and repulsive forces of molecules
(d) Attractive force and volume of the molecules

10. Gas deviates from ideal gas nature because molecules

[CPMT 1996]

(a) Are colourless
(b) Attract each other
(c) Contain covalent bond
(d) Show Brownian movement

11. The Vander Waal's equation reduces itself to the ideal gas equation at**[Kerala MEE 2001; CBSE PMT 2002]**
(a) High pressure and low temperature
(b) Low pressure and low temperature
(c) Low pressure and high temperature
(d) High pressure and high temperature

12. The compressibility factor for an ideal gas is**[IIT 1997]**
(a) 1.5 (b) 1.0
(c) 2.0 (d) ∞

13. When an ideal gas undergoes unrestrained expansion, no cooling occurs because the molecules **[IIT 1984, 89]**
(a) Are above the inversion temperature

(b) Exert no attractive force on each other

(c) Do work equal to loss in kinetic energy

(d) Collide without loss of energy

14. A gas is said to behave like an ideal gas when the relation $PV/T = \text{constant}$. When do you expect a real gas to behave like an ideal gas **[IIT 1999; CBSE PMT 1990; CPMT 1991]**

(a) When the temperature is low

(b) When both the temperature and pressure are low

(c) When both the temperature and pressure are high

(d) When the temperature is high and pressure is low

15. A real gas most closely approaches the behaviour of an ideal gas at **[KCET 1992]**

(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

16. The temperature at which real gases obey the ideal gas laws over a wide range of pressure is called **[AFMC 1993; IIT 1981, 94]**

(a) Critical temperature

(b) Boyle temperature

(c) Inversion temperature

(d) Reduced temperature

17. At low pressure, the Vander Waal's equation is reduced to

(a) $Z = \frac{pV_m}{RT} = 1 - \frac{ap}{RT}$ (b) $Z = \frac{pV_m}{RT} = 1 + \frac{b}{RT}p$

(c) $pV_m = RT$ (d) $Z = \frac{pV_m}{RT} = 1 - \frac{a}{RT}$

18. At high temperature and low pressure, the Vander Waal's equation is reduced to

(a) $\left(p + \frac{a}{V_m^2}\right)(V_m) = RT$

(b) $pV_m = RT$

(c) $p(V_m - b) = RT$

(d) $\left(p + \frac{a}{V_m^2}\right)(V_m - b) = RT$

19. When helium is allowed to expand into vacuum, heating effect is observed. Its reason is that **[CPMT 1987]**

(a) Helium is an ideal gas

(b) Helium is an inert gas

(c) The inversion temperature of helium is very low

(d) The boiling point of helium is the lowest among the elements

20. In van der Waal's equation of state of the gas law, the constant 'b' is a measure of **[AIEEE 2004]**

(a) Volume occupied by the molecules

(b) Intermolecular attraction

(c) Intermolecular repulsions

(d) Intermolecular collisions per unit volume

21. In which molecule the vander Waal's force is likely to be the most important in determining the *m.pt.* and *b.pt.* **[DPMT 2000]**

(a) H_2S (b) Br_2

(c) HCl (d) CO

22. Pressure exerted by 1 mole of methane in a 0.25 litre container at 300K using vander Waal's equation (given $1 = 2.253\, atm\, l^2\, mol^{-2}, b = 0.0428\, lit\, mol^{-1}$) is **[Orissa JEE 2005]**

(a) 82.82 atm (b) 152.51 atm

(c) 190.52 atm (d) 70.52 atm

Critical state and Liquefaction of gases

1. Which set of conditions represents easiest way to liquefy a gas **[NCERT 1983]**

(a) Low temperature and high pressure

(b) High temperature and low pressure

(c) Low temperature and low pressure

(d) High temperature and high pressure

2. Adiabatic demagnetisation is a technique used for **[BHU 1984]**

(a) Adiabatic expansion of a gas

(b) Production of low temperature

(c) Production of high temperature

(d) None

3. An ideal gas can't be liquefied because **[CBSE PMT 1992]**

(a) Its critical temperature is always above 0^oC

(b) Its molecules are relatively smaller in size

(c) It solidifies before becoming a liquid

(d) Forces operative between its molecules are negligible

4. However great the pressure, a gas cannot be liquefied above its

(a) Boyle temperature

(b) Inversion temperature

(c) Critical temperature

(d) Room temperature

5. An ideal gas obeying kinetic theory of gases can be liquefied if **[CBSE PMT 1995]**

(a) Its temperature is more than critical temperature T_c

(b) Its pressure is more than critical pressure P_c

(c) Its pressure is more than P_c at a temperature less than T_c

(d) It cannot be liquefied at any value of P and T

6. The Vander Waal's parameters for gases W, X, Y and Z are

Gas	a (atm L^2 mol^{-2})	b (L mol^{-1})
W	4.0	0.027
X	8.0	0.030
Y	6.0	0.032
Z	12.0	0.027

Which one of these gases has the highest critical temperature

(a) W (b) X
(c) Y (d) Z

7. The Vander Waal's constant 'a' for the gases O_2, N_2, NH_3 and CH_4 are 1.3, 1.390, 4.170 and 2.253 $L^2\, atm\, mol^{-2}$ respectively. The gas which can be most easily liquefied is **[IIT 1989]**

(a) O_2 (b) N_2
(c) NH_3 (d) CH_4

8. A gas can be liquefied **[AFMC 2005]**

(a) Above its critical temperature
(b) At its critical temperature
(c) Below its critical temperature
(d) At any temperature

9. Which of the following is correct for critical temperature

(a) It is the highest temperature at which liquid and vapour can coexist
(b) Beyond the critical temperature, there is no distinction between the two phases and a gas cannot be liquefied by compression
(c) At critical temperature (T_c) the surface tension of the system is zero
(d) At critical temperature the gas and the liquid phases have different critical densities

10. A gas has a density of 2.68 g/L at stp. Identify the gas

(a) NO_2 (b) Kr
(c) COS (d) SO_2

11. Weight of 112 *ml* of oxygen at NTP on liquefaction would be **[DPMT 1984]**

(a) 0.32 *g* (b) 0.64 *g*
(c) 0.16 *g* (d) 0.96 *g*

Critical Thinking

Objective Questions

1. As the temperature is raised from 20^oC to 40^oC the average kinetic energy of neon atoms changes by a factor of which of the following **[AIEEE 2004]**

(a) 313/293 (b) $\sqrt{(313/293)}$
(c) 1/2 (d) 2

2. A gas is found to have a formula$[CO]_x$. If its vapour density is 70, the value of x is **[DCE 2004]**

(a) 2.5 (b) 3.0
(c) 5.0 (d) 6.0

3. Which of the given sets of temperature and pressure will cause a gas to exhibit the greatest deviation from ideal gas behavior **[DCE 2003]**

(a) 100^oC and 4 atm (b) 100^oC and 2 atm
(c) -100^oC and 4 atm (d) 0^oC and 2 atm

4. The molecular weight of O_2 and SO_2 are 32 and 64 respectively. If one *litre* of O_2 at 15^oC and 750 *mm* pressure contains 'N' molecules, the number of molecules in two *litres* of SO_2 under the same conditions of temperature and pressure will be **[CBSE 1990; MNR 1991]**

(a) $N/2$ (b) N
(c) 2N (d) 4N

5. What is the relationship between the average velocity (v), root mean square velocity (u) and most probable velocity (a) **[AFMC 1994]**

(a) $\alpha : v : u :: 1 : 1.128 : 1.224$
(b) $\alpha : v : u :: 1.128 : 1 : 1.224$
(c) $\alpha : v : u :: 1.128 : 1.224 : 1$
(d) $\alpha : v : u :: 1.124 : 1.228 : 1$

6. Consider the following statements : For diatomic gases, the ratio C_p/C_v is equal to

(1) 1.40 (lower temperature)
(2) 1.66 (moderate temperature)
(3) 1.29 (higher temperature)

which of the above statements are correct

(a) 1, 2 and 3 (b) 1 and 2
(c) 2 and 3 (d) 1 and 3

7. The compressibility factor for an ideal gas is**[MP PET 2004**

(a) 1.5 (b) 1.0
(c) 2.0 (d) ∞

8. The compressibility factor of a gas is less than 1 at STP. Its molar volume V_m will be **[MP PET 2004]**

(a) $V_m > 22.42$ (b) $V_m < 22.42$
(c) $V_m = 22.42$ (d) None

9. If some *moles* of O_2 diffuse in 18 *sec* and same *moles* of other gas diffuse in 45 *sec* then what is the molecular weight of the unknown gas**[CPMT 1988]**

(a) $\frac{45^2}{18^2}\times 32$ (b) $\frac{18^2}{45^2}\times 32$

(c) $\frac{18^2}{45^2\times 32}$ (d) $\frac{45^2}{18^2\times 32}$

10. The ratio of rates of diffusion of SO_2, O_2 and CH_4 is **[BHU 1992]**

(a) $1:\sqrt{2}:2$ (b) 1 : 2 : 4

(c) $2:\sqrt{2}:1$ (d) $1:2:\sqrt{2}$

11. If C_1, C_2, C_3...... represent the speeds of n_1, n_2, n_3..... molecules, then the root mean square speed is **[IIT 1993]**

(a) $\left(\frac{n_1C_1^2+n_2C_2^2+n_3C_3^2+.....}{n_1+n_2+n_3+.....}\right)^{1/2}$

(b) $\frac{(n_1C_1^2+n_2C_2^2+n_3C_3^2+.....)^{1/2}}{n_1+n_2+n_3+.....}$

(c) $\frac{(n_1C_1^2)^{1/2}}{n_1}+\frac{(n_2C_2^2)^{1/2}}{n_2}+\frac{(n_3C_3^2)^{1/2}}{n_3}+......$

(d) $\left[\frac{(n_1C_1+n_2C_2+n_3C_3+....)^2}{(n_1+n_2+n_3+....)}\right]^{1/2}$

12. 50 *ml* of hydrogen diffuses out through a small hole from a vessel in 20 *minutes*. The time needed for 40 *ml* of oxygen to diffuse out is **[CBSE PMT 1994]**

(a) 12 *min* (b) 64 *min*

(c) 8 *min* (d) 32 *min*

13. At what temperature will the average speed of CH_4 molecules have the same value as O_2 has at 300 *K* **[CBSE PMT 1989]**

(a) 1200 *K* (b) 150 *K*

(c) 600 *K* (d) 300 *K*

14. A sample of O_2 gas is collected over water at 23^oC at a barometric pressure of 751 *mm Hg* (vapour pressure of water at 23^oC is 21 *mm Hg*). The partial pressure of O_2 gas in the sample collected is **[CBSE PMT 1993]**

(a) 21 *mm Hg* (b) 751 *mm Hg*

(c) 0.96 *atm* (d) 1.02 *atm*

15. In an experiment during the analysis of a carbon compound, 145 *l* of H_2 was collected at 760 *mm Hg* pressure and 27^oC temperature. The mass of H_2 is nearly **[MNR 1987]**

(a) 10 *g* (b) 12 *g*

(c) 24 *g* (d) 6 *g*

16. The volume of 1 *g* each of methane (CH_4), ethane (C_2H_6), propane (C_3H_8) and butane (C_4H_{10}) was measured at 350 *K* and 1 *atm*. What is the volume of butane **[NCERT 1981]**

(a) 495 cm^3 (b) 600 cm^3

(c) 900 cm^3 (d) 1700 cm^3

17. The ratio of the rate of diffusion of helium and methane under identical condition of pressure and temperature will be **[IIT 2005]**

(a) 4 (b) 2

(c) 1 (d) 0.5

18. At what temperature in the celsius scale, *V* (volume) of a certain mass of gas at 27^oC will be doubled keeping the pressure constant **[Orissa 1993]**

(a) 54^oC (b) 327^oC

(c) 427^oC (d) 527^oC

19. Pressure of a mixture of 4 *g* of O_2 and 2 *g* of H_2 confined in a bulb of 1 *litre* at 0^oC is **[AIIMS 2000]**

(a) 25.215 *atm* (b) 31.205 *atm*

(c) 45.215 *atm* (d) 15.210 *atm*

20. If pressure becomes double at the same absolute temperature on 2 $L\,CO_2$, then the volume of CO_2 becomes **[AIIMS 1992]**

(a) 2 *L* (b) 4 *L*

(c) 25 *L* (d) 1 *L*

21. Volume of the air that will be expelled from a vessel of 300 cm^3 when it is heated from 27^oC to 37^oC at the same pressure will be

(a) 310 cm^3 (b) 290 cm^3

(c) 10 cm^3 (d) 37 cm^3

22. 300 *ml* of a gas at 27^oC is cooled to -3^oC at constant pressure, the final volume is **[NCERT 1981, MP PMT 1992]**

(a) 540 *ml* (b) 135 *ml*

(c) 270 *ml* (d) 350 *ml*

Assertion & Reason

For AIIMS Aspirants

Read the assertion and reason carefully to mark the correct option out of the options given below :

(a) If both assertion and reason are true and the reason is the correct explanation of the assertion.

(b) If both assertion and reason are true but reason is not the correct explanation of the assertion.

(c) If assertion is true but reason is false.
(d) If the assertion and reason both are false.
(e) If assertion is false but reason is true.

1. Assertion : Plot of P Vs. $1/V$ (volume) is a straight line.

Reason : Pressure is directly proportional to volume.

2. Assertion : Jet aeroplane flying at high altitude need pressurization of the cabin.

Reason : Oxygen is not present at higher altitude.

3. Assertion : 1 mol of H_2 and O_2 each occupy 22.4 L of volume at 0^oC and 1 bar pressure.

Reason : Molar volume for all gases at the same temperautre and pressure has the same volume.

4. Assertion : Pressure exerted by a mixture of reacting gases is equal to the sum of their partial pressures.

Reason : Reacting gases react to form a new gas having pressure equal to the sum of both.

5. Assertion : Greater the value of Vander Waal's constant 'a' greater is the liquefaction of gas.

Reason : 'a' indirectly measures the magnitude of attractive forces between the molecules.

6. Assertion : Carbondioxide has greater value of root mean square velocity μ_{rms} than carbon monoxide.

Reason : μ_{rms} is directly proportional to molar mass.

7. Assertion : 4.58 mm and 0.0098^oC is known to be triple point of water.

Reason : At this pressure and temperature all the three states i.e., water, ice and vapour exist simultaneously.

8. Assertion : 1/4th of the gas is expelled if air present in an open vessel is heated from 27^oC to 127^oC.

Reason : Rate of diffusion of a gas is inversely proportional to the square root of its molecular mass.

9. Assertion : Compressibility factor for hydrogen varies with pressure with positive slope at all pressures.

Reason : Even at low pressures, repulsive forces dominate hydrogen gas. **[AIIMS 20**

10. Assertion : vander Waal's equation is applicable only to non-ideal gases.

Reason : Ideal gases obey the equation $PV = nRT$.

11. Assertion : Pressure exerted by gas in a container with increasing temperature of the gas.

Reason : With the rise in temperature, the average speed of gas molecules increases.

[AIIMS 1995]

12. Assertion : Gases do not settle to the bottom of container.

Reason : Gases have high kinetic energy.

[AIIMS 1997]

13. Assertion : A mixture of He and O_2 is used for respiration for deep sea divers.

Reason : He is soluble in blood. **[AIIMS 1998]**

14. Assertion : Wet air is heavier than dry air.

Reason : The density of dry air is more than density of water. **[AIIMS 1999]**

15. Assertion : All molecules in a gas have some speed.

Reason : Gas contains molecules of different size and shape. **[AIIMS 2001]**

16. Assertion : Effusion rate of oxygen is smaller than nitrogen.

Reason : Molecular size of nitrogen is smaller than oxygen. **[AIIMS 2004]**

Answers

Characteristics and Measurable properties of gases

1	c	2	c	3	b	4	a	5	c
6	d	7	a	8	b	9	c	10	a
11	a	12	b	13	a	14	d	15	c
16	e								

Ideal gas equation and Related gas laws

1	c	2	d	3	a	4	a	5	a
6	bc	7	a	8	a	9	c	10	d
11	a	12	a	13	a	14	a	15	c
16	c	17	b	18	c	19	a	20	b
21	c	22	c	23	d	24	a	25	c
26	c	27	b	28	c	29	c	30	c
31	a	32	a	33	b	34	b	35	b

36	c	37	d	38	c	39	c	40	a
41	b	42	c	43	c	44	d	45	a
46	a	47	b	48	d	49	b	50	c
51	d	52	a	53	a	54	c	55	a
56	a	57	b	58	a	59	b	60	a
61	b	62	d	63	c	64	b	65	d
66	b	67	d	68	a	69	a	70	b
71	c	72	a	73	d	74	c	75	a
76	c	77	d	78	b	79	b	80	b
81	d	82	d	83	c	84	c	85	a
86	d	87	a	88	a	89	c	90	c
91	a	92	c	93	a	94	b	95	a
96	a	97	d	98	b	99	e	100	a
101	c	102	c	103	c	104	c	105	a
106	a								

Kinetic molecular theory of gases and Molecular collisions

1	d	2	c	3	d	4	d	5	b
6	a	7	a	8	c	9	b	10	b
11	a	12	d	13	a	14	a	15	acd
16	a	17	b	18	d	19	b	20	d
21	a	22	d	23	c	24	b	25	b
26	a	27	d	28	a	29	d	30	c
31	c								

Molecular speeds

1	a	2	c	3	d	4	d	5	d
6	c	7	b	8	b	9	d	10	b
11	d	12	c	13	a	14	d	15	d
16	c	17	d	18	d	19	b	20	b
21	c	22	a	23	c	24	b	25	c
26	a	27	d						

Real gases and Vander waal’s equation

1	b	2	b	3	c	4	c	5	d
6	c	7	d	8	b	9	d	10	b
11	c	12	b	13	b	14	d	15	c
16	b	17	a	18	b	19	c	20	a
21	b	22	a						

Critical state and Liquefaction of gases

1	a	2	b	3	d	4	c	5	d
6	d	7	c	8	c	9	abc	10	c
11	c								

Critical Thinking Questions

1	a	2	c	3	c	4	c	5	a
6	d	7	b	8	b	9	a	10	a
11	a	12	b	13	b	14	c	15	b
16	a	17	b	18	b	19	a	20	d
21	c	22	c						

Assertion & Reason

1	c	2	c	3	a	4	d	5	a
6	d	7	a	8	b	9	a	10	b
11	a	12	a	13	c	14	e	15	d
16	c								

AS Answers and Solutions

Characteristics and Measurable properties of gases

1. (c) Gases do not have definite shape and volume. Their volume is equal to the volume of the container.

2. (c) All the three phases of water can coexist at 0^oC & 4.7 mm pressure.

3. (b) It is characteristic of gases *i.e.* Thermal energy >> molecular attraction

4. (a) In gases, molecular attraction is very less and intermolecular spaces are large hence kinetic energy of gases is highest.

5. (c) Gases and liquids, both can flow and posses viscosity.

7. (a) Newton is unit of force.

8. (b) $\frac{C^o}{5} = \frac{F^o - 32}{9}$

9. (c) $1\,L = 10^{-3}m^3 = 10^3 cm^3 = 1 dm^3 = 10^3 ml$.

10. (a) 1 atm = 10^6 dynes cm^{-2}

12. (b) Barometer is used to measure atmospheric pressure of mixture of gases. Staglometer is used to measure surface tension. Only manometer is used to measure pressure of pure gas in a vessel.

13. (a) 0^oC is equivalent to 273^oK *i.e.* conditions are same so volume will be V *ml.*

14. (d) The mass of gas can be determined by weighing the container, filled with gas and again weighing this container after removing the gas. The difference between the two weights gives the mass of the gas.

15. (c) Nobel gases has no intermolecular forces due to inertness.

16. (e) Total volume of two flasks = 1+ 3 = 4

If P_1 the pressure of gas N_2 in the mixture of N_2 and O_2 then

P = 100 kPa , $P_1 = ?$, V = 1 $litre$, $V_1 = 4 litre$

applying Boyle's law $PV = P_1V_1$

100 × 1 = $P_1 \times 4$; $P_1 = 25$

If P_2 is the pressure of O_2 gas in the mixture of O_2 and N_2 then, 320 × 3 = $P_2 \times 4$; $P_2 = 240$

Hence, Total pressure $P = P_1 + P_2 = 25 + 240$

$= 265\ kPa$

Ideal gas equation and Related gas laws

1. (c) Boyle's law is $V \propto \frac{1}{P}$ at constant T

2. (d) According to Boyle's law $V \propto \frac{1}{P}$

$V = \frac{\text{Constant}}{P}$; VP = Constant.

3. (a) At sea level, because of compression by air above the proximal layer of air, pressure increases hence volume decreases *i.e.* density increases. It is Boyle's law.

4. (a) At constant $T, P_1V_1 = P_2V_2$

$1 \times 20 = P_2 \times 50$; $P_2 = \frac{20}{50} \times 1$

5. (a) $P.V$ = constant at constant temperature. As temperature changes, the value of constant also changes.

6. (b,c)According to Boyle's Law PV = constant, at constant temperature either P increases or V increases both (b) & (c) may be correct.

7. (a) $\frac{d_1}{d_2} = \frac{1}{2}, \frac{T_1}{T_2} = \frac{2}{1} \because \frac{P_1}{P_2} = \frac{V_2}{V_1} \times \frac{T_1}{T_2} = \frac{T_1.d_1}{T_2.d_2}$

$\frac{P_1}{P_2} = \frac{2}{1}.\frac{1}{2} = \frac{1}{1}$

8. (a) Absolute temperature is temperature measured in $^o Kelvin$, expressed by T

11. (a) $T_1 = 273^oC = 273 + 273^oK = 546^oK$

$T_2 = 0^oC = 273 + 0^oC = 273^oK$

$P_1 = 1$; $P_2 = ?$

According to Gay-Lussac's law

$\frac{P_1}{T_1} = \frac{P_2}{T_2} \therefore P_2 = \frac{P_1T_2}{T_1} = \frac{1 \times 273^oK}{546^oK}$ atm; $\frac{1}{2}$ atm.

12. (a) $V_t = V_o(1 + \alpha_v t)$

$\because (V_2 - V_1) = \Delta V = V_o\alpha(t_2 - t_1)$

if $t_2 - t_1 = 1^o$ then $\Delta V = \alpha V_o$

For every $1^o C$ increase in temperature, the volume of a given mass of an ideal gas increases by a definite fraction $\frac{1}{273.15}$ of V_o.

Here V_o is volume at $0^o C$ temperature.

13. (a) $\frac{V_1}{V_2} = \frac{T_1}{T_2} \because V_2 = \frac{T_2}{T_1} V_1 = \frac{546^o K}{273^o K} \times 0.2L = 0.4L.$

14. (a) $V_2 = \frac{T_2}{T_1} . V_1 = \frac{270^o K}{300^o K} . 400\, cm^3 = 360\, cm^3$

contraction $= V_1 - V_2 = 400 - 360 = 40 cm^3$

15. (c) At constant volumes $P \propto T$

P = constant T; $PV = nRT \therefore P = \frac{nR}{V} T$

slope = $m = \frac{nR}{V} \because V_2 < V_1$

$\frac{m_1}{m_2} = \frac{V_2}{V_1} \therefore m_1 < m_2$ is curve for V_2 has a greater slope than for V_1

16. (c) $\frac{P_1}{T_1} + \frac{P_1}{T_1} = \frac{P}{T_1} + \frac{P}{T_2}$

$\frac{2P_1}{T_1} = P\left(\frac{T_1 + T_2}{T_1 T_2}\right)$; $\therefore P = \frac{2P_1(T_1 T_2)}{T_1(T_1 + T_2)} = \frac{2P_1 T_2}{T_1 + T_2}$

18. (c) At constant V of a definite mass

$\frac{P_1}{T_1} = \frac{P_2}{T_2} \therefore \frac{P_1}{P_2} = \frac{300}{600} = \frac{1}{2}$ *i.e.* pressure increases and on increasing temperature energy of molecules increases so the rate of collisions also increases and number of moles remains constant because there is neither addition nor removal of gas in the occurring.

19. (a) Avogadro number = 6.0224×10^{23}

20. (b) Compressibility = $\frac{PV}{nRT} < 1$ at STP (as given)

$nRT > PV$
$N \times 0.0821 \times 273 > 1V_m$
$22.41\ litres > V_m$

R = 0.821
T = 273°K
P = 1
n = 1

21. (c) The value of universal gas constant can be expressed in different units and its value would depend only on the units of the measurement.

22. (c) $PV = nRT$

$R = \frac{PV}{nT}$ = litre . atm. K^{-1} $mole^{-1}$

23. (d) (atm. K^{-1} mol^{-1}) is not a unit of R

24. (a) 8.31 $J.K^{-1}$ mol^{-1}

1 cal = 4.2 J.

$\therefore \frac{8.31}{4.2} cal.K^{-1} mol^{-1}$ = 1.987 cal K^{-1} mol^{-1}

30. (c) $PV = nRT \therefore \frac{n}{V} = \frac{P}{RT}$

31. (a) $P = \frac{nRT}{V} = \frac{2 \times 0.0821 \times 546}{44.8\, l} = 2$ atm.

32. (a) $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \therefore n_2 = \frac{P_2 V_2 T_1}{P_1 V_1 T_2} n_1$

at STP n_1 = one mole
P_1 = 1 atm.
V_1 = 22.4 lt
T_1 = 273°K

at $T = 273^o + 30^o = 303^o K$
P_2 = 1 atm
$V_2 = 22.4$

$n_2 = \frac{1}{1} \times \frac{22.4}{22.4} \times \frac{273}{303} \times 1 = 0.9$ moles

33. (b) $V = \frac{nRT}{P} = \frac{0.5 \times 0.082 \times 273^o K}{1} = 11.2 lit$

34. (b) $V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} \Rightarrow$ $P_1 = P$; $T_1 = 273^o K$

$P_2 = \frac{3}{2} P$; $T_2 = T_1 + \frac{T_1}{3} = \frac{4}{3} \times 273^o K$

$V_2 = \frac{2P}{3P} \times \frac{4}{3} \times \frac{273}{273} \times 100\, cc = \frac{800}{9} cc = 88.888\, cc$

= 88.9 cc

35. (b) $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T} \therefore \frac{P_1 V_1}{P_2 V_2} = \frac{T_1}{T_2}$

36. (c) $d_a = 2d_b$; $2M_a = M_b$

$PV = nRT = \frac{m}{M} RT$; $P = \frac{m}{V} . \frac{RT}{M} = \frac{dRT}{M}$

$\frac{P_a}{P_b} = \frac{d_a}{d_b} \frac{M_b}{M_a} = \frac{2d_b}{d_b} \times \frac{2M_a}{M_a} = 4$

37. (d) n of $O_2 = \frac{16}{32} = \frac{1}{2}$

n of $H_2 = \frac{3}{2}$

Total no. of moles = $\frac{3}{2} + \frac{1}{2} = 2$

$V = \frac{nRT}{P} = \frac{2 \times .082 \times 273}{1} = 44.8 lit$ = 44800 ml

38. (c) $n = \frac{PV}{RT} = \frac{m}{M}$

$m = \frac{MPV}{RT} = \frac{34 \times 2 \times 100}{0.082 \times 293} = 282.4 gm$

39. (c) $V_2 = \frac{P_1 V_1}{T_1} \frac{T_1}{P_2} = \frac{760}{600} \times \frac{546}{273} \times 273 = 691.6 ml.$

40. (a) $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \therefore T_2 = \frac{P_2}{P_1} \frac{V_2}{V_1} T_1 \frac{m_1}{m_2}$

$= \frac{0.75}{1} \times \frac{1}{1} \times \frac{2}{1} \times 300^o K = 450^o K$

41. (b) $V_2 = \frac{P_1}{P_2} \frac{T_2}{T_1} . V_1 = \frac{1}{0.5} \times \frac{250}{300} \times 12000\ lit. = 20000\ lit.$

42. (c) At constant pressure

$V \propto nT \propto \frac{m}{M} T$

$\frac{V_1}{V_2} = \frac{m_1 T_1}{m_2 T_2} \therefore \frac{T_1}{T_2} = \frac{V_1}{m_1} \times \frac{m_2}{V_2} = \frac{d_2}{d_1} \Rightarrow \frac{300^o K}{T_2} = \frac{0.75 d}{d}$

$T_2 = \frac{300}{0.75} = 400^o K$

43. (c) $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \therefore T_2 = \frac{P_2V_2}{P_1V_1} \times T_1 = \frac{740}{740} \times \frac{80}{100} \times 300$

$= 240^o K = -33^o C$

47. (b) Because H_2 & Cl_2 gases may react with each other to produce HCl gas hence Dalton's law is not applicable.

48. (d) Because HCl & NH_3 gases may react to produce NH_4Cl gas. Dalton's Law is applicable for non reacting gas mixtures.

49. (b) NH_3 and HCl & HBr is a reacting gas mixture to produce $NH_4Cl \,\&\, NH_4Br$ so Dalton's law is not applicable.

50. (c) No. of moles of lighter gas $= \frac{m}{4}$

No. of moles of heavier gas $= \frac{m}{40}$

Total no. of moles $= \frac{m}{4} + \frac{m}{40} = \frac{11m}{40}$

Mole fraction of lighter gas $= \frac{\frac{m}{4}}{\frac{11m}{40}} = \frac{10}{11}$

Partial pressure due to lighter gas $= P_o \times \frac{10}{11}$

$= 1.1 \times \frac{10}{11} = 1 atm.$

52. (a) m. wt. of $NH_3 = 17$; m.wt. of $N_2 = 28$

m.wt. of $CO_2 = 44$; m.wt. of $O_2 = 32$

beacuse NH_3 is lightest gas out of these gases

$$\left[r \propto \frac{1}{\sqrt{\text{Molecular Weight}}} \right]$$

55. (a) $\frac{r_g}{r_{He}} = \sqrt{\frac{M_{He}}{M_g}} \therefore M_g = M_{He} . \frac{r^2{}_{He}}{r^2{}_g} = \frac{4}{(1.4)^2} = \frac{4}{1.96} = 2$

$\left[Note : 1.4 = \sqrt{2} \right]$

56. (a) $r_g = \frac{1}{5} . r_{H_2}$

$\frac{M_g}{M_{H_2}} = \left[\frac{r_{H_2}}{r_g} \right]^2 = (5)^2 = 25$; $M_g = 2 \times 25 = 50$

57. (b) $r_g = \frac{1}{6} r_{H_2}$; $M_g = M_{H_2} . \left[\frac{r_{H_2}}{r_g} \right]^2 = 2 \times 6^2 = 2 \times 36 = 72$

58. (a) $M_1 = 64$; $r_2 = 2r_1$

$M_2 = M_1 \left[\frac{r_1}{r_2} \right]^2 = 64 \times \frac{1}{4} = 16$

59. (b) $r_O = r_H \sqrt{\frac{d_H}{d_O}} = 1 \sqrt{\frac{0.09}{1.44}} = \sqrt{\frac{1}{16}} = \frac{1}{4}$

60. (a) $r_a = 5r_b$; $\frac{d_a}{d_b} = \left[\frac{r_b}{r_a} \right]^2 = \left(\frac{1}{5} \right)^2 = \frac{1}{25}$

61. (b) $\frac{d_1}{d_2} = \frac{1}{16}$; $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{16} = \frac{4}{1}$

62. (d) $\frac{D_A}{D_B} = \sqrt{\frac{\rho_B}{\rho_A}} = \left[\frac{\rho_B}{\rho_A} \right]^{\frac{1}{2}}$; $\therefore D_A = D_B \left(\frac{\rho_B}{\rho_A} \right)^{\frac{1}{2}}$

63. (c) Gases may be separated by this process because of different rates of diffusion due to difference in their densities.

64. (b) NH_4Cl ring will first formed near the HCl bottle because rate of diffusion of NH_3 is more than that of HCl because $M_{NH_3} : M_{HCl} = 17 : 36.5)$. SO NH_3 will reach first to the HCl bottle & will react there with HCl to form NH_4Cl ring

65. (d) Because both NO and C_2H_6 have same molecular weights $\lfloor M_{NO} = M_{C_2H_6} = 30 \rfloor$ and rate of diffusion $\propto$ molecular weight.

67. (d) $\frac{M_A}{M_B} = \left(\frac{r_B}{r_A} \right)^2$ $\because r_A = 2r_B$ $\therefore \frac{r_B}{r_A} = \frac{1}{2} = \frac{1}{(2)^2} = \frac{1}{4} = .25$

68. (a) $r_H = \frac{2gm}{10 \min}$ if $r_O = \frac{xgm}{10 \min}$

$r_O = r_H \sqrt{\frac{M_{H_2}}{M_{O_2}}} = \frac{2}{10} \sqrt{\frac{2}{32}}$

$\frac{x}{10} = \frac{2}{10 \times 4} = \frac{1}{2} gm. = .5 gm$

69. (a) $r_{CH_4} = 2r_g$

$M_g = M_{CH_4} \left(\frac{r_{CH_4}}{r_g} \right)^2 = 16 \times 2^2 = 64$

70. (b) $r \propto \frac{1}{\sqrt{M}}$ $\because r = \frac{Volume \;\; effused}{time \;\; taken} = \frac{V}{t}$

$\frac{V}{t} \propto \frac{1}{\sqrt{M}}$ $\therefore$ for same volumes (V constant)

$t \propto \sqrt{M} \therefore \frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}}$

$t_{He} = t_{H_2} \sqrt{\frac{M_{He}}{M_{H_2}}} = 5\sqrt{\frac{4}{2}} = 5\sqrt{2}s.$

$t_{O_2} = t = 5\sqrt{\frac{32}{2}} = 20 s$

$t_{CO} = 5\sqrt{\frac{28}{2}} = 5\sqrt{14} s$; $t_{CO_2} = 5\sqrt{\frac{44}{2}} = 5\sqrt{22} s$

71. (c) $\frac{r_{N_2}}{r_{SO_2}} = \frac{V_{rms} N_2}{V_{rms} SO_2} = \sqrt{\frac{T_{N_2}}{T_{SO_2}} . \frac{M_{So_2}}{M_{N_2}}} = \sqrt{\frac{T_{N_2}}{323} \times \frac{64}{28}}$

$1.625 = \sqrt{\frac{T_{N_2}}{323} . \frac{16}{7}}$

$$T_{N_2} = \frac{(1.625)^2 \times 323 \times 7}{16} = 373^o K$$

72. (a) $C + H_2O \rightarrow CO_{(g)} + H_{2(g)}$

$12 gm \rightarrow 1 mol + 1 mol$

12 *gm C* produces 2*mole* of gases (1*mole CO* & 1 *mole* of H_2)

$\therefore$48 *gm C* may produce $\frac{48}{12} \times 2 = 4 \times 2 = 8 mole$

$= 22.4 \times 8\ L$ gases $= 179.2\ L$ gas.

73. (d) Molecular weight = $\frac{mRT}{PV} = \frac{4.4 \times .082 \times 273}{1 \times 2.24}$

= 44

So the gas should be CO_2

74. (c) $PV = nRT$

$P = \frac{n}{V} RT \quad \because \frac{n}{V} = C \Rightarrow \quad P = CRT$

$T = \frac{P}{CR} = \frac{1}{1 \times .821} = 12^o K$

75. (a) 6.02×10^{22} molecules of each N_2, O_2 and H_2

$= \frac{6.02 \times 10^{22}}{6.02 \times 10^{23}}$ *moles* of each

Weight of mixture = weight of 0.1 *mole* N_2 + weight of 0.1 *mole* H_2 + weight of 0.1 mole of O_2

$= (28 \times 0.1) + (2 \times 0.1) + (32 \times 0.1) \ = 6.2 gm$

76. (c) *M.wt* of CO_2 = 12+16+16 = 44

Volume of 44 *gm* of CO_2 at NTP = 22.4 *litre*

1 *gm* of CO_2 at NTP = $\frac{22.4}{44}$

$4.4 gm$ of CO_2 at N.T.P

$\Rightarrow \frac{22.4}{44} \times 4.4$ *litre* $= 2.24$ *litre*

78. (b) No. of moles of CO_2 present in 200 *ml* solution

= molarity × Volume (in *lt.*) $= 0.1 \times \frac{200}{1000} = .02$

Volume of 0.02 mole of $CO_2 = 22.4 \times .02 lt. = 0.448 lit.$

79. (b) Molecular weight = *V.d.* × 2 = 11.2 × 2 = 22.4

Volume of 22.4 *gm* Substance of NTP = 22.4 *litre*

1 *gm* substance at NTP = $\frac{22.4}{22.4}$ *litre*

11.2 *gm* substance of NTP = 11.2 *litre*

80. (b) $\frac{M.wt. of\ O_2}{M.wt. of\ SO_2} \Rightarrow \frac{M_1}{M_2} \Rightarrow \frac{32}{64} = \frac{1}{2}$

The weight of oxygen will be $\frac{1}{2}$ that of SO_2

81. (b) For HI has the least volume because of greater molecular weight $V \propto \frac{1}{M}$

83. (c) Since no. of molecules is halved so pressure should also be halved.

84. (c) H_2 will be filled first because of lower molecular weight

85. (a) Mixture of SO_2 and Cl_2 are reacted chemically and forms SO_2Cl_2. That is why mixture of these gases is not applicable for Dalton's law.

86. (d) According to Boyle's law

$P_1V_1 = P_2V_2 \Rightarrow P_1 \times 60 = 720 \times 100$

$P_1 = \frac{720 \times 100}{60} = 1200\ mm$

87. (a) Rate of diffusion $\propto \frac{1}{\sqrt{\text{Molecular Mass}}}$

that is why H_2 gas diffuse first

88. (a) Solution level will rise, due to absorption of CO_2 by sodium hydroxide.

$2NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$

89. (c) $\underset{(40+12+16\times3)=100\ gm}{CaCo_3} \xrightarrow{\hat{}} CaO + \underset{22.4\ litre}{CO_2} \uparrow$

$\because$ At S.T.P. $100\ g\ CaCO_3$ produce= $22.4 litre$ of CO_2

$\therefore$ At S.T.P. $1 g$ $CaCO_3$ produce = $\frac{22.4}{100} = .224\ litre$ of CO_2

90. (c) The density of gas $= \frac{Molecular\ wt. Of\ Metal}{Volume}$

$= \frac{45}{22.4}$

= $2 gm litre^{-1}$

91. (a) $M_1 = 32g$ for O_2, $M_2 = 2g$ for H_2

$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$; $\quad \frac{r_1}{r_2} = \sqrt{\frac{2}{32}} = \sqrt{\frac{1}{16}} = \frac{1}{4}$

92. (c) In $22.4 l$ of H_2 maximum number of molecules $= 6.023 \times 10^{23}$

In $1 l$ of H_2 maximum number of molecules $= \frac{6.023 \times 10^{23}}{22.4}$

In $15 l$ of H_2 maximum number of molecules $= \frac{6.023 \times 10^{23}}{22.4} \times 15 \ = 4.03 \times 10^{23}$ molecules.

93. (a) $22.4 l\ O_2$ at S.T.P. $= 32 gm$ of O_2

$1 l\ O_2$ at S.T.P. $= \frac{32}{22.4} = 1.43 gm$ of O_2

95 (a) We know that molecular mass of hydrogen $M_1 = 2$ and that of helium $M_2 = 4$, we also know that Graham's law of diffusion

$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{4}{2}} = \sqrt{2} = 1.4$; $r_1 = 1.4m$

96. (a) $\frac{r_A}{r_H} = \sqrt{\frac{M_H}{M_A}} = \frac{r}{6r} = \sqrt{\frac{2}{M_A}}$

$M_A = 6 \times 6 \times 2 = 72g$

97. (d) Given that:

$V_1 = 100ml, P_1 = 720mm, V_2 = 84ml, P_2 = ?$

By using $P_1V_1 = P_2V_2$ [According to the Boyle's law]

$$P_2 = \frac{P_1V_1}{V_2} = \frac{720 \times 100}{84} = 857.142$$

Hence, $P_2 = 857.14mm$

98. (b) According to gas law

$PV = nRT$, $n = \frac{PV}{RT}$

$$\frac{n_A}{n_B} = \frac{\frac{P_1V_1}{RT_1}}{\frac{P_2V_2}{RT_2}} \; ; \; \frac{n_A}{n_B} = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2V_2}$$

$$\frac{n_A}{n_B} = \frac{2P \times 2V}{2T} \times \frac{T}{PV} \; ; \; \frac{n_A}{n_B} = \frac{2}{1}$$

99. (e) No. of molecules $= 2 \times V.d$

$2 \times 38.3 = 76.3$

wt. of $NO_2 = x$

So that *wt.* of $N_2O_4 = 100 - x$

Hence, $\frac{x}{46} + \frac{100 - x}{92} = \frac{100}{76.6} = \frac{2x + 100 - x}{92} = \frac{100}{76.6}$

x = 20.10, no. of mole. of $NO_2 = \frac{20.10}{46} = 0.437$

100. (a) Given that

$P_1 = 76cm$ of Hg (Initial pressure at N.T.P.)

$P_2 = ?$, $V_1 = 5litre$, $V_2 = 30 + 5 = 35litres$

According to Boyle's law

$P_1V_1 = P_2V_2$; $76 \times 5 = P_2 \times 35$

$P_2 = \frac{76 \times 5}{35} \Rightarrow P_2 = 10.8cm$ of Hg

101. (c) Given initial volume $(V_1) = 300cc$, initial temperature $(T_1) = 27^oC = 300K$, initial pressure $(P_1) = 620mm$, final temperature $T_2 = 47^oC = 320K$ and final pressure $(P_2) = 640mm$ .We know from the general gas equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{620 \times 300}{300} = \frac{640 \times V_2}{320} \Rightarrow V_2 = 310cc$$

102. (c) $NH_3 + HCl \rightarrow NH_4Cl$

$4litre$ $1.5litre$

HCl is a limiting compound. That's why $1.5litre$ of HCl reacts with $1.5litre$ of NH_3 and forms NH_4Cl . Thus (4 - 1.5) $2.5litre$ NH_3 remains after the reaction.

103. (c) $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$; $\frac{P_1 \times 4}{T_1} = \frac{2P_1 \times V_2}{2T_1}$

$8 = 2 \times V_2$ so $V_2 = 4dm^3$

104. (c) $P_1 = P, V_1 = V, T_1 = T$

$P_2 \frac{P}{2}, V_2 = ?, T_2 = T$

According to gas equation

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ or $\frac{PV}{T} = \frac{PV_2}{2T}$

$\therefore \quad V_2 = 4V$

105. (a) $P_1 = P, V_1 = V, T_1 = 273 + 75 = 348K$

$P_2 = 2P, V_2 = \frac{85}{100}, T_2 = ?$

$\frac{P \times V}{398} = \frac{2P \times 85V}{T_2 \times 100} \Rightarrow T_2 = \frac{348 \times 2 \times 85}{100}$

$\therefore \; T_2 = 591.6K = 318.6^oC$

106. (a) Boyle's law $-PV$ = constant

On differentiating the equation,

$d(PV) = d(C) \Rightarrow PdV + VdP = 0$

$\Rightarrow VdP = -PdV \Rightarrow \frac{dP}{P} = -\frac{dV}{V}$.

Kinetic molecular theory of gases and Molecular collisions

4. (d) Kinetic energy $= \frac{3}{2}RT$

5. (b) $K.E$ = $\frac{3}{2}RT$ it means that $K.E$ depends upon T (absolute temperature) only.

7. (a) $KE = \frac{3}{2}RT = \frac{3}{2}PV$

$\therefore P = \frac{2}{3}\frac{E}{V}$ for unit volume (V = 1) $\Rightarrow P = \frac{2}{3}E$

8. (c) Tr. $K.E. = \frac{3RT}{2}$ it means that the Translational Kinetic energy of Ideal gas depends upon temperature only.

9. (b) $\frac{E_{He}}{E_{H_2}} = \sqrt{\frac{T_{He}}{T_{H_2}}}$ so energies will be same for He & H_2 at same temperature.

11. (a) $K.E. = \frac{3}{2}.RT = \frac{3}{2}.2.T \qquad \because R \approx 2calK^{-1}mol^{-1}$

$K.E. = 3T$

12. (d) All molecules of an ideal gas show random motion. They collide with each other and walls of container during which they lose or gain energy so they may not have same kinetic energy always.

13. (a) For same temperature kinetic energies of H_2 & He molecules will be same because kinetic energy depends only on temperature.

14. (a) For same temp. kinetic energies would be equal for all molecules, what ever their molecular weights will be, it doesn't matter.

15. (a,c,d)Kinetic energies per molecule will be same because it is proportional to absolute temperature only.

$\frac{d_{N_2}}{d_{CO_2}} = \frac{M_{N_2}}{M_{CO_2}} = \frac{28}{44}$ *i.e.* $dN_2 < dCO_2$

Total translational kinetic energy will also be same because at same temperature & pressure number of molecules present in same volume would be same (according to Avogadro's Law)

16. (a) On increasing pressure, the volume decreases and density increases. So molecules get closer to each other hence mean free path also decreases.

17. (b) Most probable velocity increase and fraction of molecule possessing most probable velocity decrease.

18. (d) $PV = nRT = \frac{m}{M}RT$

$\therefore \frac{m}{V} = \frac{PM}{RT} =$ density

19. (b) $d \propto \frac{P}{T}$ the value of $\frac{P}{T}$ is maximum for (*b*)

20. (d) If inversion temperature is $80^oC = 193^oK$ then the temperature, at which it will produce cooling under Joule Thomson's effect, would be below inversion temperature except 173^oK all other values given as

21. (a) Since $\frac{C_P}{C_V} = 1.4$, the gas should be diatomic.

If volume is 11.2 *lt* then, no. of moles = $\frac{1}{2}$

$\therefore$ no. of molecules = $\frac{1}{2} \times$ Avagadro's No.

no. of atoms = 2 × no. of molecules

2 × $\frac{1}{2} \times$ Avagadro's No.

$= 6.0223 \times 10^{23}$

22. (d) Density = $\frac{M}{V}$

$d = \frac{v.d \times 2}{V}$ $\quad (M = V.d \times 2)$

$V.d = \frac{d \times V}{2}$

$V.d = \frac{0.00130 \times 22400}{2} = 14.56\, gm^{-1}$

23. (c) Volume of steam = 1*lt* = $10^3 cm^3$

$\because m = d.V$

$\therefore$mass of $10^3 cm^3$ steam = density × Volume

= $\frac{0.0006\, gm}{cm^3} \times 10^3 cm^3 = 0.6 gm$

Actual volume occupied by H_2O molecules is equal to volume of water of same mass

$\therefore$ Actual volume of H_2O molecules in $6gm$ steam

= mass of steam/density of water

= $0.6\, gm$ /1 *gm/cm*³ $\Rightarrow 0.6\, cm^3$

24. (b) $r = \frac{C_P}{C_V} = \frac{5}{3} = 1.66$ (For Monoatomic as He, Ne, Ar)

25. (b) The density of neon will be highest at 0^oC 2 *atm* according to $d \propto \frac{P}{T}$

29. (d) $K.E.$ per mole $= \frac{3}{2}RT$

so all will have same $K.E.$ at same temperature.

30. (c) $\because W = P.dV = E$

$\therefore$Energy per unit volume = P

31. (c) $d \propto M \Rightarrow \frac{d_1}{d_2} = \frac{M_1}{M_2}$; $\frac{3d}{d} = \frac{M}{M_2}$; $M_2 = \frac{M}{3}$.

Molecular speeds

1. (a) $V_{rms} = \sqrt{\frac{3RT}{M}}, V_{av} = \sqrt{\frac{8RT}{\pi M}}$; $\frac{V_{rms}}{V_{av}} = \sqrt{\frac{3\pi}{8}}$

$= \sqrt{\frac{66}{56}} \Rightarrow \frac{1.086}{1}$

3. (d) most probable velocity : mean velocity : V_{rms}

$= \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}} = \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3}$

4. (d) $V_{rms} = \sqrt{\frac{3RT}{M}} \therefore V_{rms} \propto \frac{1}{\sqrt{M}}$ at same T

because H_2 has least molecular weight so its r.m.s. velocity should be maximum.

5. (d) $\frac{U_{SO_2}}{U_{He}} = \frac{1}{2} = \sqrt{\frac{M_{He}}{M_{SO_2}} \frac{T_{SO_2}}{T_{He}}} = \sqrt{\frac{4}{64} . \frac{T_{SO_2}}{300}}$

$= \frac{4}{64} \frac{T_{SO_2}}{300} = \frac{1}{4}$; $T_{SO_2} = 1200^oK$

6. (c) $\frac{U_{O_3}}{U_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{O_3}}} = \sqrt{\frac{32}{48}} = \sqrt{\frac{2}{3}}$

7. (b) Average kinetic energy per molecule

$= \frac{3}{2}KT = \frac{3}{2} \times 1.38 \times 10^{-23} \times 300\, J = 6.17 \times 10^{-21} J$

8. (b) $\frac{U_{SO_2}}{U_{O_2}} = \sqrt{\frac{M_{O_2} T_{SO_2}}{M_{SO_2} T_{O_2}}} = \sqrt{\frac{32 \times T_{SO_2}}{64 \times 303}} = 1$

$1 = \frac{32 \times T_{SO_2}}{64 \quad 303} \Rightarrow T_{SO_2} = 606$

9. (d) Among these Cl_2 has the highest molecular weight so it will posses lowest root mean square velocity.

10. (b) $\frac{U_1}{U_2} = \sqrt{\frac{T_1}{T_2}}$ $\quad \therefore \frac{T_1}{T_2} = \left(\frac{5 \times 10^4}{10 \times 10^4}\right)^2 = \frac{1}{4}$

12. (c) $V_{rms} = \sqrt{\frac{3KT}{Molecular\ weight}}$ *i.e.* $V_{rms} \propto \frac{1}{\sqrt{m}} \propto (m)^{-\frac{1}{2}}$

13. (a) When average speed of molecule is increased due to increase in temperature then the change in momentum during collision between wall of container and molecules of gas also increases.

14. (d) $\frac{U_1}{U_2} = \sqrt{\frac{m_2}{m_1} \cdot \frac{T_1}{T_2}}$ $\quad \because T_1 = T_2$

$\frac{U_1^2}{U_2^2} = \frac{m_2}{m_1}$ $\quad \therefore m_1 U_1^2 = m_2 U_2^2$

15. (d) $U_2 = U_1 \sqrt{\frac{T_2}{T_1}} = U_1 \sqrt{\frac{1200}{300}} = U_1 \times 2$

r.m.s. velocity will be doubled.

16. (c) $\frac{U_{H_2}}{U_{O_2}} = \sqrt{\frac{T_{H_2}}{M_{H_2}} \cdot \frac{M_{O_2}}{T_{H_2}}} = \sqrt{\frac{50}{2} \cdot \frac{32}{800}} = 1$

17. (d) $U = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3P}{d}} \therefore U \propto \frac{1}{\sqrt{d}}$

19. (b) $\frac{U_1}{U_2} = \sqrt{\frac{n_1 T_1}{n_2 T_2}} = \sqrt{\frac{n \times T}{2n \times 2T}} = \sqrt{\frac{1}{4}} = \frac{1}{2}$

$U_2 = 2U_1 = 2U$

21. (c) $\frac{V_{mp}}{V_{av}} = \frac{\sqrt{\frac{2RT}{M}}}{\sqrt{\frac{8RT}{\pi M}}} = \frac{\sqrt{\pi}}{2}$

22. (a) $V_{rms} = \sqrt{\frac{3RT}{M}} \Rightarrow V_{rms} = \sqrt{T}$

Given $V_1 = V$, $T_1 = 300\,K$, $V_2 = 2V$, $T_2 = ?$

$\frac{V_1}{V_2} = \sqrt{\frac{T_1}{T_2}}$; $\left(\frac{V}{2V}\right)^2 = \frac{300}{T_2} \Rightarrow T_2 = 300 \times 4 = 1200\,K$

24. (b) $\left(P - \frac{n^2 a}{V^2}\right)(V - nb) = nRT$

$\left(P - \frac{(2)^2 \times 4.17}{(5)^2}\right)(5 - 2 \times .03711) = 2 \times .0821 \times 300$

$P = \frac{2 \times .0821 \times 300}{5 - 2 \times .03711} - \frac{4.7 \times 2^2}{5^2} \Rightarrow 10 - 0.66 = 9.33\,atm$

26. (a) Average speed : most probable speed

$\sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{2RT}{M}} \Rightarrow \sqrt{\frac{8}{\pi}} : \sqrt{2} \Rightarrow 1.128 : 1.$

27. (d) $v_{rms} = \sqrt{\frac{3RT}{M}}$

$\sqrt{30^2 R} = \sqrt{\frac{3RT}{M}} \Rightarrow 30 \times 30 R = \frac{3R \times 300}{M}$

$\Rightarrow M = \frac{3 \times 300}{30 \times 30} 1\,gm = 0.001\,kg$.

Real gases and Vander waal's equation

2. (b) Because molecules of real gases have intermolecular forces of attraction so the effective impact on the wall of container is diminished. Pressure of real gas is reduced by $\frac{a}{v^2}$ factor hence behaviour of real gas deviate from ideal behaviour.

3. (c) $Z = \frac{PV}{RT}$ $\quad \because$ for ideal gas $PV = RT$ so $Z = 1$

12. (b) $Z = \frac{PV}{RT}$; for ideal gas $PV = RT$; so $Z = 1$

13. (b) Ideal gas has no attractive force between the particles

14. (d) $PV = nRT$ is a ideal gas equation it is allowed when the temperature is high and pressure is low.

16. (b) At Boyle temperature real gas is changed into ideal gas

17. (a) When pressure is low

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

or $PV = RT + Pb - \frac{a}{V} + \frac{ab}{V^2}$ or $\frac{PV}{RT} = 1 - \frac{a}{VRT}$

$Z = -\frac{a}{VRT}$ $\left(\because \frac{PV}{RT} = Z\right)$

18. (b) At high temperature and low pressure, Vander Waal's equation is reduced to ideal gas equation.

$PV = nRT$

$PV = RT$ (For 1 mole of gas)

20. (a) Vander waal's constant for volume correction b is the measure of the effective volume occupied by the gas molecule.

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

$\left(P + \frac{2.253}{0.25 \times 0.25}\right)(0.25 - 0.0428) = 0.0821 \times 300$

or $(P + 36.048)(0.2072) = 24.63$

$\Rightarrow P + 36.048 = 118.87 \Rightarrow P = 82.82$ atm.

Critical state and Liquefaction of gases

2. (b) A diabatic demagnetisation is a technique of liquefaction of gases in which temperature is reduced.

3. (d) An ideal gas can't be liquefied because molecules of ideal gas have not force of attraction between them.

4. (c) At above critical temperature, substances are existing in gaseous state, since gas cannot be liquefied above it.

5. (d) Absence of inter molecular attraction ideal gas cannot be liquefied at any volume of P and T.

6. (d) For Z gas of given gases, critical temperature is highest

$$T_c = \frac{8a}{27Rb} \Rightarrow T_c = \frac{8 \times 12}{27 \times .0821 \times .027} = 1603.98K$$

7. (c) Value of constant a is greater than other for NH_3 that's why NH_3 can be most easily liquefied.

8. (c) The temperature below which the gas can be liquefied by the application of pressure alone is called critical temperature.

10. (c) $d = \frac{M}{V} \Rightarrow M = d \times V$

$M = 2.68 \times 22.4$ at N.T.P. $(\because V = 22.41)$

$M = 60.03\, gm$

m. wt of $COS = 12 + 16 + 32 = 60$

11. (c) 22400 ml is the volume of O_2 at N.T.P =32gm of O_2

1ml is the volume of O_2 at NTP = $\frac{32}{22400}$

112 ml is the volume of O_2 at NTP = $\frac{32}{22400} \times 112$

$= 0.16\, gm$ of O_2

Critical Thinking Questions

1. (a) Average kinetic energy $\propto$ (T Kelvin)

(Factor) $\frac{K.E_2}{K.E_1} = \frac{T_2}{T_1} = \frac{40 + 273}{20 + 273} = \frac{313}{293}$

2. (c) M. wt. = $V.d.$ × 2

= 70 × 2 = 140 ⇒ $x = \frac{m.wt.}{wt.of[CO]} = \frac{140}{[12 + 16]} = 5$

3. (c) Gas deviate from ideal gas behaviour to real gas (according to Vander Waal's at low temperature and high pressure)

4. (c) At same temperature and pressure, equal volumes have equal number of molecules. If 1lit. of oxygen consists N molecules then at same temperature and pressure 1 lit of SO_2 will consists N molecules. So 2 lit. of SO_2 will contain $2N$ molecules.

5. (a) $V_{av} : V_{rms} : V_{\text{most probable}} = V : U : \alpha$

$$\sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}} : \sqrt{\frac{2RT}{M}}$$

$$\alpha : V : U = \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3} = 1 : 1.128 : 1.224$$

6. (d) $\frac{C_p}{C_v}$ ratio for diatomic gases is 1.40 at lower temperature & 1.29 at higher temperature so the answer is 'd'.

7. (b) $PV = nRT$ (For ideal gas)

$Z = \frac{PV}{nRT} = 1$ (For ideal gas)

8. (b) If $Z < 1$ then molar volume is less than 22.4 L

9. (a) $r_{O_2} = \frac{x}{18} mole / \sec \Rightarrow r_g = \frac{x}{45} mol / \sec$

$$M_g = M_{O2}\left(\frac{r_{O_2}}{r_g}\right)^2 = 32\left(\frac{x}{18} \times \frac{45}{x}\right)^2 = 32 \times \frac{45^2}{18^2}$$

10. (a) $r_{SO_2} : r_{O_2} : r_{CH_4} = \frac{1}{\sqrt{M_{SO_2}}} : \frac{1}{\sqrt{O_2}} : \frac{1}{\sqrt{CH_4}}$

$$= \frac{1}{\sqrt{64 : 32 : 16}} = \frac{1}{\sqrt{4 : 2 : 1}}$$

$$\frac{1}{2} : \frac{1}{\sqrt{2}} : \frac{1}{1}; \quad \frac{2}{2} : \frac{2}{\sqrt{2}} : \frac{2}{1}; \quad 1 : \sqrt{2} : 2$$

11. (a) Root mean square speed $= \left[\frac{n_1c_1^2 + n_2c_2^2 + n_3c_3^2 + ...}{n_1 + n_2 + n_3 + ...}\right]^{1/2}$.

12. (b) If 40 ml O_2 will diffuse in t min. then.

$$r_{O_2} = \frac{40}{t}$$

$$r_H = \frac{50}{20} \Rightarrow r_O = r_H 2\sqrt{\frac{M_{H_2}}{M_{O_2}}} = \frac{50}{20}\sqrt{\frac{2}{32}} = \frac{50}{20}.\frac{1}{4}$$

$$\frac{40}{t} = \frac{50}{80} \therefore t = \frac{40 \times 80}{50} = 60 \ min.$$

13. (b) $\frac{V_{av}CH_4}{V_{ab}O_2} = \sqrt{\frac{T_{CH_4}}{T_{O_2}}.\frac{M_{O_2}}{M_{CH_4}}} = 1$

$$\frac{T_{CH_4}}{300}.\frac{32}{16} = 1; \ T_{CH_4} = 150^o K$$

14. (c) Pressure of O_2 (dry) = 751-21 = 730 $mm\ Hg$

$$= \frac{730}{760} = 0.96\, atm$$

15. (b) $PV = nRT$, $n = \frac{PV}{RT} = \frac{1 \times 145}{0.082 \times 3} = 5.8 \approx 6\, mole$.

16. (a) $V = \frac{nRT}{P} = \frac{m}{M}.\frac{RT}{P} = \frac{1}{58} \times \frac{0.82 \times 350}{1} = 0.495\, lit.$

= 495 cm^3

17. (b) $\frac{r_{He}}{r_{CH_4}} = \sqrt{\frac{M_{CH_4}}{M_{He}}} = \sqrt{\frac{16}{4}} = 2$

18. (b) $\frac{V_1}{V_2} = \frac{T_1}{T_2} \because T_2 = \frac{T_1V_2}{V_1} = 300^o K, \frac{2V}{V} = 600^o K$

$$T_2 = 600^o K = (600 - 273)^o C = 327^o C$$

19. (a) no. of moles of $O_2 = \frac{4}{32} = 0.125$

no. of moles of $H_2 = \frac{2}{2} = 1$

total no. of moles = $1 + 0.125 = 1.125$

$$P = \frac{nRT}{V} = \frac{1.125 \times 0.082 \times 273}{1} = 25.184\, atm.$$

20. (d) $\frac{P_1}{P_2} = \frac{1}{2}, \because \frac{V_1}{V_2} = \frac{P_2}{P_1} = \frac{2}{1}$

$\frac{2L}{V_2} = \frac{2}{1}$; $V_2 = 1L$

21. (c) $\frac{V_2}{V_1} = \frac{T_2}{T_1}$

$\because V_2 = \frac{T_2}{T_1} V_1 = \frac{310^o K}{300^o K} \times 300\,cm^3 = 310\,cm^3$

22. (c) $V_2 = \frac{T_2}{T_1}.V_1 = \frac{270^o K}{300^o K} \times 300\,ml = 270\,ml$

Assertion & Reason

1. (c) Pressure is inversly proportional to volume (Boyle's law). $p\alpha\frac{1}{V}$ (n, T constant).

2. (c) The air pressure decreases with increase in altitude. So the partial pressure of Oxygen is not sufficient for breathing at higher altitude and thus pressurization is needed.

3. (a) At a given temperature and pressure the volume of a gas is directly proportional to the amount of gas $V\alpha n$ (P and T constant).

4. (d) According to Dalton's law of partial pressure, the pressure exerted by a mixture of non – interacting gases is equal to the sum of their partial pressures (pressure exerted by individual gases in mixture) $P_{Total} = P_1 + P_2 + P_3 \ldots$ (T and V constant). Both the gases if non-interacting would spread uniformly to occupy the whole volume of the vessel.

5. (a) Considering the attractive force pressure in ideal gas equation (PV = nRT) is correct by introducing a factor of $\frac{an^2}{V^2}$ where a is a vander waal's constant.

6. (d) $\mu_{rms} = \sqrt{\frac{3RT}{M}}$ is inversly related to molecular mass. Therefore, $\mu_{rms}(CO) > \mu_{rms}(CO_2)$.

8. (b) $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (Initial fraction $\frac{V_1}{V_2} = 1$ when temperature is $27^o C$. At $127^o C$ the new fraction is $\frac{V_1}{V_2} = \frac{300}{400} = \frac{3}{4}$ $\therefore$ air expelled $\Rightarrow 1 - \frac{3}{4} = \frac{1}{4}$

9. (a) In case of H_2, compressibility factor increases with the pressure. At 273 K, Z > 1 which shows that it is difficult to compress the gas as compared to ideal gas. In this case repulsive forces dominate.

10. (b) In real gases, the intermolecular forces of attraction and the volume occupied by the gas molecules cannot be neglected.

11. (a) When the temperature increase, the average speed of gas molecules increases and by this increase the pressure of gas is also increases.

12. (a) It is correct that gases do not settle to the bottom of container and the reason for this is that due to higher kinetic energy of gaseous molecules they diffuse.

13. (c) The assertion, that a mixture of helium and oxygen is used for deep sea divers, is correct. The He is not soluble in blood. Therefore, this mixture is used.

14. (e) Dry air is heavier than wet air because the density of dry air is more than water.

15. (d) All molecule of a gas have different speed. Therefore, they move by its own speed.

16. (c) Assertion is true but reason is false because of effusion rate $\propto \frac{1}{\sqrt{M}}$ (Molecular weight) but it does not depend on molecular size.

Gaseous State

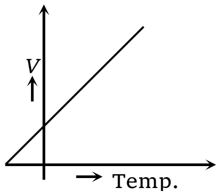
Self Evaluation Test - 6

1. Same mass of CH_4 and H_2 is taken in container. The partial pressure caused by H_2 is **[IIT 1989; CPMT 1996]**

(a) 8 / 9 (b) 1 / 9

(c) 1 / 2 (d) 1

2. The following graph illustrates **[JIPMER 2000]**



(a) Dalton's law (b) Charle's law

(c) Boyle's law (d) Gay-Lussac's law

3. If the pressure and absolute temperature of 2 *litres* of CO_2 are doubled, the volume of CO_2 would become

[CBSE PMT 1991]

(a) 2 *litres* (b) 4 *litres*

(c) 5 *litres* (d) 7 *litres*

4. What is kinetic energy of $1g$ of O_2 at 47^oC

[Orissa JEE 2004]

(a) 1.24×10^2J (b) 2.24×10^2J

(c) 1.24×10^3J (d) 3.24×10^2J

5. The root mean square speeds at STP for the gases H_2, N_2, O_2 and HBr are in the order

[Pb. CET 1994; CBSE PMT 1991]

(a) $H_2<N_2<O_2<HBr$

(b) $HBr<O_2<N_2<H_2$

(c) $H_2<N_2=O_2<HBr$

(d) $HBr<O_2<H_2<N_2$

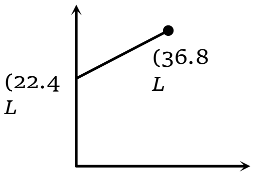
6. By what ratio the average velocity of the molecule in gas change when the temperature is raised from 50 to 200^oC

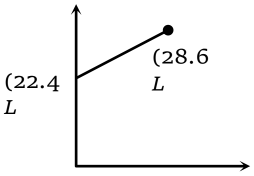
[DCE 2003]

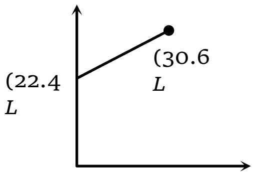
(a) 1.21 / 1 (b) 1.46 / 1

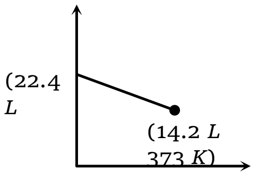
(c) 1.14 / 1 (d) 4 / 1

7. Which of the following volume (V) – temperature (T) plots represents the behaviour of one mole of an ideal gas at one atmospheric pressure **[IIT Screening 20**

(a)


(b)


(c)


(d)


8. If the average velocity of N_2 molecule is $0.3m/s$ at 27^oC, then the velocity will be $0.6m/s$ at **[Pb. CET 2001]**

(a) $1200K$ (b) $600K$

(c) $400K$ (d) $1800K$

9. Equal volumes of two gases which do not react together are enclosed in separate vessels. Their pressures at 100 *mm* and 400 *mm* respectively. If the two vessels are joined together, then what will be the pressure of the resulting mixture (temperature remaining constant) **[CBSE PMT 1981]**

(a) 125 *mm* (b) 500 *mm*

(c) 1000 *mm* (d) 250 *mm*

10. A gas of volume 100 *cc* is kept in a vessel at pressure 10^4 *Pa* maintained at temperature 24^oC. If now the pressure is increased to 10^5 *Pa*, keeping the temperature constant, then the volume of the gas becomes **[AFMC 1992]**

(a) 10 *cc* (b) 100 *cc*

(c) 1 *cc* (d) 1000 *cc*

11. If a gas is expanded at constant temperature **[IIT 1986]**

(a) The pressure increases

(b) The kinetic energy of the molecules remains the same

(c) The kinetic energy of the molecules decreases

(d) The number of molecules of the gas increases

12. The rate of diffusion of SO_2 and O_2 are in the ratio

[Assam JET 1991; EAMCET 1980]

(a) $1:\sqrt{2}$ (b) 1 : 32

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

AS Answers and Solutions

(SET -6)

1. (a) N_{CH4} = number of moles of $CH_4 = \frac{m}{16}$

N_{H2} = number of moles of $H_2 = \frac{m}{2}$

fraction partial pressure of H_2 is

$$H_2 = \frac{n_{H_2}}{n_{H_2}+n_{CH_4}} = \frac{\frac{m}{2}}{\frac{m}{2}+\frac{m}{16}} = \frac{\frac{m}{2}}{\frac{9m}{16}} = \frac{8}{9}$$

2. (b) According to Charle's Law $V \propto T$

$V_t = V_o + V_o \alpha t$

compare it with $Y = C + mx$

3. (a) $V2 = \frac{P_1V_1}{T_1} . \frac{T_2}{P_2} = \frac{P}{2P} \times 2lt \times \frac{2T}{T} = 2lt$

4. (a) K.E. $= \frac{3}{2}nRT = \frac{3}{2} \times \frac{1}{32} \times 8.314 \times 320\, J$.

$= 1.24 \times 10^2\, J$

5. (b) $V_{rms} \propto \frac{1}{\sqrt{m}}$

$U_{H_2} : U_{N_2} : U_{O_2} : U_{HBr} = \frac{1}{\sqrt{2}} : \frac{1}{\sqrt{20}} : \frac{1}{\sqrt{32}} : \frac{1}{\sqrt{81}}$ is

$U_{HB_r} < U_{O_2} < U_{N_2} < U_{H_2}$

6. (c) $T_1 = 150 + 273 = 423K$; $T_2 = 50 + 273 = 323K$

Hence, $\frac{(V_{av})_1}{(V_{av})_2} = \sqrt{\frac{T_1}{T_2}} = \sqrt{\frac{423}{323}} = \frac{1.14}{1}$

7. (c) $\frac{V_1}{V_2} = \frac{T_1}{T_2} \therefore V_2 = V_1 . \frac{T_2}{T_1} = \frac{22.4 \times 373}{273} = 30.6L$

8. (a) $V_{rms} = \sqrt{\frac{3RT}{M}}$; $V_{rms} = \sqrt{T}$

Given, $V_1 = V$, $T_1 = 300\,K$

$V_2 = 2V$, $T_2 = ?$

$= \frac{V_1}{V_2} = \sqrt{\frac{T_1}{T_2}} = \left(\frac{V}{2V}\right)^2 = \frac{300}{T_2}$

$T_2 = 300 \times 4 = 1200\,K$

9. (d) When two vessels are joined together, the volume will be doubled hence effective pressure will be halved

$P = \frac{P_1+P_2}{2} = \frac{100+400}{2} = 250\,mm$

10. (a) $P_1V_1 = P_2V_2$ at constant T

$10^4 .100 = 10^5 \times V_2$

$V_2 = 10\,cc$

11. (b) Kinetic energy will also remain constant if Temperature is constant.

12. (a) $\frac{r_{SO_2}}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{SO_2}}} = \sqrt{\frac{32}{64}} = \frac{1}{\sqrt{2}}$

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