Chapter 13
Kinetic Theory

1 Marks Questions

1.Given Samples of 1 cm^3 of Hydrogen and 1 cm^3 of oxygen, both at N. T. P. which sample has a larger number of molecules?

Ans. Acc. to Avogadro's hypothesis, equal volumes of all gases under similar conditions of temperature and pressure contain the same number of molecules. Hence both samples have equal number of molecules. Hence both samples have equal number of molecules.

2.Find out the ratio between most probable velocity, average velocity and root Mean square Velocity of gas molecules?

Ans. Since,

Most Probable velocity, $V_{mP} = \sqrt{\frac{2KT}{m}}$

Average velocity, $\overline{V} = \sqrt{\frac{8KT}{\pi m}}$

Root Mean Square velocity: Vr.m.s. $= \sqrt{\frac{3KT}{m}}$

So, $V_{mp} : \overline{V}\ V_{r.m.s} = \sqrt{\frac{2KT}{m}} : \sqrt{\frac{3KT}{\pi m}} : \sqrt{\frac{3KT}{m}}$

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

$V_{mp} : \overline{V}\ V_{r.m.s.} = 1 : 1.3 : 1.23$

3.What is Mean free path?

Ans. Mean free path is defined as the average distance a molecule travels between collisions. It is represented by $\lambda\ (\text{lambda})$. Units are meters (m).

4.What happens when an electric fan is switched on in a closed room?

Ans. When electric fan is switched on, first electrical energy is converted into mechanical energy and then mechanical energy is converted into heat. The heat energy will increase the Kinetic energy of air molecules; hence temperature of room will increase.

5.State the law of equi-partition of energy?

Ans. According to law of equi partition of energy, the average kinetic energy of a molecule in each degree of freedom is same and is equal to $\frac{1}{2}KT$.

6.On what factors, does the average kinetic energy of gas molecules depend?

Ans. Average kinetic energy depends only upon the absolute temperature and is directly proportional to it.

7.Why the temperature less than absolute zero is not possible?

Ans. Since, mean square velocity is directly proportional to temperature. If temperature is zero then mean square velocity is zero and since K. E. of molecules cannot be negative and hence temperature less than absolute zone is not possible.

8.What is the relation between pressure and kinetic energy of gas?

Ans. Let, Pressure = P

Kinetic energy = E

From, Kinetic theory of gases, $P = \frac{1}{3} Sc^2 \rightarrow 1)$

S = Density

C = r.m.s velocity of gas molecules

Mean Kinetic energy of translation per unit

Volume of the gas = $E = \frac{1}{2} Sc^2 \rightarrow 2)$

Dividing 1) by 2)

$$\frac{P}{E} = \frac{1Sc^2 2}{3 \times Sc^2} = \frac{2}{3}$$

$$\Rightarrow P = \frac{2}{3} E$$

9.What is an ideal perfect gas?

Ans.A gas which obeys the following laws or characteristics is called as ideal gas.

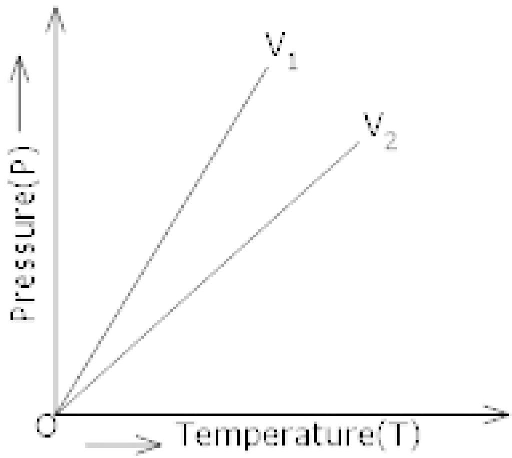
1) The size of the molecule of gas is zero

2) There is no force of attraction or repulsion amongst the molecules of gas.

2 Marks Questions

1.If a certain mass of gas is heated first in a small vessel of volume V_1 and then in a large vessel of volume V_2. Draw the P – T graph for two cases?

Ans. From Perfect gas equation; $P = \frac{RT}{V}$



For a given temperature, $P \alpha \frac{1}{V}$ therefore when the gas is heated in a small vessel (Volume V_1) , the pressure will increases more rapidly than when heated in a large vessel (Volume V_2). As a result, the slope of P – T graph will be more in case of small vessel than that of large vessel.

2.Derive the Boyle's law using kinetic theory of gases?

Ans. According to Boyle's law, temperature remaining constant, the volume v of a given mass of a gas is inversely proportional to the pressure P i.e. PV = constant.

Now, according to kinetic theory of gases, the pressure exerted by a gas is given by:-

P = Pressure

V = Volume

$\overline{V}$ = Average Velocity

m = Mass of 1 molecule

N = No. of molecules

M = mN (Mass of gas)

$$P = \frac{1mN\overline{V}^2}{3V}$$

$$P_V = \frac{1}{3}M\overline{V}^2$$

3.At what temperature is the root mean square speed of an atom in an argon gas cylinder equal to the r.m.s speed of a helium gas atom at- 20^0C? Given Atomic Mass of Ar = 39.9 and

of He = 4.0?

Ans. Suppose, Vr.m.s. and V^1r.m.s. are the root mean square speeds of Argon and helium atoms at temperature T and T^1 respectively.

R = Universal Gas constant

T = Temperature

M = Atomic Mass of Gas

Now, Vr.m.s. = $\sqrt{\frac{3RT}{M}}$

$$V^1r.m.s.=\sqrt{\frac{3RT^1}{M^1}}$$

Now, M = Mass of Argon = 39.9

M^1 = Mass of Helium = 4.0

T^1 = Temperature of helium = -20^0C

$T^1 = 273 + (-20) = 253$ K.

T = Temperature of Argon = ?

Since Vr.m.s. = V^1r.m.s

$$\sqrt{\frac{3RT}{M}} = \sqrt{\frac{3RT^1}{M^1}}$$

Squaring both side,

$$\frac{\cancel{3}\cancel{R}T}{M} = \frac{\cancel{3}\cancel{R}T^1}{M^1}$$

$$\frac{T}{M} = \frac{T^1}{M^1} \Rightarrow T = \frac{T^1 M}{M^1}$$

Putting the values of $T^1, M^1 \,\&\, M$

$$T = \frac{253 \times 39.9}{4.0} = 2523.7K$$

4.Show that constant – temperature bulk modulus K of an ideal gas is the pressure P of the gas?

Ans. When a substance is subjected to a Pressure increase ΔP will undergo a small fractional volume decrease $\frac{\Delta V}{V}$ That is related to bulk modulus K by :-

$$K = \frac{\Delta P}{\frac{-\Delta V}{V}} \quad \rightarrow 1)$$

Negative sign indicates decrease in volume. In case of an ideal gas at constant temperature before compression,

$$PV = \frac{m}{M} RT \quad \rightarrow 2)$$

M = Molecular Mass of gas

After compression at constant temperature,

$$(P + \Delta P)(V + \Delta V) = \frac{m}{M} RT$$

From equation 2)

$$PV = (P + \Delta P)(V + \Delta V)$$

$$\cancel{PV} = \cancel{PV} + P\Delta V + V\Delta P + \Delta P \Delta V$$

or $-P\Delta V = V\Delta P + \Delta P \Delta V$

$$-\frac{P\Delta V}{V} = \Delta P + \frac{\Delta P \Delta V}{V} \quad (\because \text{ Dividing by V on both sides})$$

$$-\frac{P\Delta V}{V} = \Delta P\left(1 + \frac{\Delta V}{V}\right)$$

$$-\frac{\Delta V}{V} = \frac{\Delta P}{P}\left(1 + \frac{\Delta V}{V}\right)$$

We are concerned with only a small fractional changes. Therefore, $\frac{\Delta V}{V}$ is much smaller than 1, As a result, it can be neglected as compared to 1.

$$\therefore -\frac{\Delta V}{V} = \frac{\Delta P}{P}$$

Substituting this value of $\frac{\Delta V}{V}$ in equation 1) we get

$$K = \frac{\Delta P}{\frac{\Delta P}{P}} = P$$

Hence, bulk modulus of an ideal gas is equal to the pressure of the gas in compression carried out at constant temperature.

5.The earth with out its atmosphere would be inhospitably cold. Explain Why?

Ans. The lower layers of earth's atmosphere reflect infrared radiations from earth back to the surface of earth. Thus the heat radiations received by the earth from the sun during the day are kept trapped by the atmosphere. If atmosphere of earth were not there, its surface would become too cold to live.

6.If a vessel contains 1 mole of O_2 gas (molar mass 32) at temperature T. The pressure of the gas is P. What is the pressure if an identical vessel contains 1 mole of He at a temperature 2 T?

Ans. By ideal gas equation :→

$$PV = nRT$$

$$P = \text{pressure}$$

V=volume

n=No. of molecule per unit volume

R=Universal Gas Constant

T=Temperature

Now, $\frac{PV}{T} = nR$ or $\frac{PV}{T} = \text{constant}$

Hence $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \rightarrow 1)$

Now, according to question:→

$$\begin{array}{l|l} P_1 = P & T_1 = T \\ V_1 = V & T_2 = 2T \end{array}$$

Using above equations in equation 1)

$$P_2 = \frac{P_1V_1}{T_1} X \frac{T_2}{V_2}$$

$$P_2 = \frac{PV}{T} X \frac{2T}{V} \qquad V_1 = V_2 = V \left(\because \textit{idantical} \text{ vessels}\right)$$

$P_2 = 2P$

Hence pressure gets doubled.

7.At very low pressure and high temperature, the real gas behaves like ideal gas. Why?

Ans. An ideal gas is one which has Zero volume of molecule and no intermolecular forces. Now:

1) At very low pressure, the volume of gas is large so that the volume of molecule is negligible compared to volume of gas.

2) At very high temperature, the kinetic energy of molecules is very large and effect of intermolecular forces can be neglected.

Hence real gases behave as an ideal gas at low pressure and high temperature.

8.Calculate the degree of freedom for monatomic, diatomic and triatomic gas?

Ans . The degrees of freedom of the system is given by:- f = 3 N – K

Where, f = degrees of freedom

N = Number of Particles in the system.

K = Independent relation among the particles.

1) For a monatomic gas; N = 1 and K = 0

f = 3 X 1 – 0 = 3

2) For a diatomic gas ; N = 2 and K = m1

f = 3 X 2 – 1 = 5

3) For a triatomic gas; N = 3 and K = 3

f = 3 X 3 – 3

f = 6

9.Determine the volume of 1 mole of any gas at s. T. P., assuming it behaves like an ideal gas?

Ans. From ideal gas equation: →

P = Pressure

V = Volume

n = No. of moles of gas

R = Universal Gas Constant

T = Temperature

PV = nRT

V = $\frac{nRT}{P}$

Here n = 1 mole; R = 8:31 J/mol/K ; T = 273K

P = 1.01×10^5 N|m^2

$$V=\frac{1\times(8.31)\times273}{1.01\times10^5}$$

$$V=22.4\times10^{-3}m^3$$

Since 1 litre

$$=1000cm^3$$

$$=1\times10^{-3}m^3$$

Hence V = 22.4 l

i.e. 1 mol of any gas has a volume of 22.4l at S. T. P. (Standard Temperature & Pressure).

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10. A tank of volume $0.3m^3$ contains 2 moles of Helium gas at 20^0C. Assuming the helium behave as an ideal gas;

1) Find the total internal energy of the system.

2) Determine the r. m. s. Speed of the atoms.

Ans .1) n = No. of moles = 2

T = Temperature = 273+20 = 293K

R = Universal Gas constant = 8.31 J/mole.

Total energy of the system = E = $\frac{3}{2}nRT$

$$E=\frac{3}{2}\times n\times8.31\times293 \quad E=7.30\times10^3 J$$

2) Molecular Mass of helium = 4 g | mol

$$= \frac{4\times10^{-3} Kg}{mol}$$

Root Mean speed = Vr.m.s $= \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3X8.31X293}{4X10^{-3}}}$

Vr.m.s. = $1.35X10^3$ m|s

11.State Graham's law of diffusion and derive it?

Ans. According to Graham's law of diffusion, the rates of diffusion of two gases are inversely proportional to the square roots of their densities.

Consider two gases A and B diffusing into each other at a Pressure P. Let S_A and S_B be their densities. The root Mean square velocities of the molecules of gases A and B will be:→

$$V^A r.m.s = \sqrt{\frac{3P}{S_A}} \rightarrow 1)$$

$$V^B r.m.s = \sqrt{\frac{3P}{S_B}} \rightarrow 2)$$

Dividing equation 1) by 2)

$$\frac{V^A r.m.s.}{V^B r.m.s.} = \sqrt{\frac{3P}{S_A}} X \sqrt{\frac{S_B}{3P}} \qquad = \sqrt{\frac{S_B}{S_A}} \rightarrow 1)$$

Now, the rate of diffusion of a gas is directly proportional to r.m.s. velocity of its molecules. If r_A and r_B are the rates of diffusion of gases A and B respectively then

$$\frac{r_A}{r_B} = \frac{V^A r.m.s.}{V^B r.m.s.} = \sqrt{\frac{S_B}{S_A}}$$

Or $\Rightarrow$ This is Graham's law.

$$\frac{r_A}{r_B} = \sqrt{\frac{S_B}{S_A}}$$

12.State Charles's law? If air is filled in a vessel at 60^0c. To what temperature should it be heated in order that $\frac{1^{rd}}{3}$ of air may escape out of vessel?

Ans. Acc. to Charles's law, for pressure remaining constant the volume of the given mass of a gas is directly proportional to its Kelvin temperature i.e.

$V \alpha T$ if pressure is constant; V = volume T = Temperature

Or $\frac{V}{T}$ = constant

Now, $T_1 = 60+273 = 333k$

$V_1 = V$;

$T_2 = ?$

$V_2 = V + \frac{V}{3} = \frac{4}{3}V$

So, from Charles's show;

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

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

$T_2 \frac{333 \times 4V}{3 \times V}$

$T_2 = 171^0c$ or 444k

13.Show that average kinetic energy of translation per molecule of gas is directly proportional to the absolute temperature of gas?

Ans. Acc. to kinetic theory of gases, the pressure p exerted by one mole of an ideal gas is

P = $\frac{1MC^2}{3V}$ M = Mass of gas

or PV = $\frac{1MC^2}{3}$ V = Volume of gas

Since PV = RT (for 1mole of gas)

or $\frac{1}{3}MC^2 = RT$ R=Universal gas constant

$C^2 = \frac{3RT}{M}$ T=Temperature

So, C $\alpha\sqrt{T}$

Also, $\frac{1}{3}MC^2 = RT$

Dividing by number of molecules of gas = N

$\frac{1M}{3N}C^2 = \frac{R}{N}T$ K=Boltzman constant

$\frac{1}{3}mc^2 = KT$ $\rightarrow$ Dividing

or $\frac{1}{2}mc^2 = \frac{3}{2}KT$

Since, $\frac{1}{2}mc^2 = Kinetic$ energy per molecule of gas

So, $\frac{1}{2}mc^2 \alpha T$

as $\frac{3}{2}k =$ constant

14.Air pressure in a car tyre increases during driving? Why?

Ans. During driving, the temperature of air inside the tyre increases due to motion. Acc. to Charles's law, pressure α Temperature, ∴ As temperature increases, Pressure inside the tyres also increases

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15.Four molecules of gas have speeds 2, 4, 6, 8, km/s. respectively.

Calculate 1) Average speed

2) Root Mean square speed?

Ans. Here, C_1 = km/s= velocity of first gas

C_2=4km/s \=velocity of second gas

C_3=6km/s=velocity of third gas

C_4=8km/s=velocity of fourth gas

1) ∴ Average speed = $\frac{C_1 + C_2 + C_3 + C_4}{4}$

Average Speed = $\frac{2+4+6+8}{4}$

Average Speed = $\frac{20}{4} = 5$ km/s

2) Root Mean Square Speed = $\sqrt{\frac{C_1^2 + C_2^2 + C_3^2 + C_4^2}{4}}$

R. m. s of gas = $\sqrt{\frac{2^2+4^2+6^2+8^2}{4}}$

R. m. s. of gas = $\sqrt{\frac{120}{4}}$

R. m. s of gas = 5.48km/s

3 Marks Questions

1.If Nine particles have speeds of 5, 8, 12, 12, 12, 14, 14, 17 and 20 m/s. find :→

1) the average speed

2) the Most Probable speed

3) the r.m.s. Speed of the particles?

Ans. 1) The average speed is the sum of speeds divided by the total number of particles.

Hence, Average speed, $\overline{V} = \frac{5+8+12+12+12+14+14+17+20}{9} = 12.7 m|s$

2) The average value of the square of speeds is given by:-

$$\overline{V}^2 = \frac{5^2+8^2+12^2+12^2+12^2+14^2+14^2+17^2+20^2}{9}$$

$$= \frac{25+64+144+144+144+196+196+289+400}{9} = \frac{1602}{9}$$

$$\overline{V}^2 = 178 m^2/s^2$$

$$\therefore R.M.S\, speed, Vr.m.s = \sqrt{\overline{V}^2} = \sqrt{178} = 13.3 m/s$$

3) Three of particles have a speed of 12m|s; two have a speed of 14m|s and the remaining have different speeds. Therefore, the most probable speed,

V_{mP} = 12 m /s.

2.Establish the relation between $Y\left(=\frac{C_P}{C_V}\right)$ and degrees of freedom (n)?

Ans. Now $y = \frac{C_P}{C_V}$

Where CP = specific heat at constant pressure

CV = Specific heat at constant volume.

and n = Degrees of freedom → is the total number of co-ordinates or independent quantities required to describe completely the position and configuration of the system.

Suppose, a polyatomic gas molecule has 'n' degrees of freedom.

∴ Total energy associated with a gram molecule of the gas i. e.

N = Total number of molecules

R = Universal Gas Constant

R = NK

K = Boltzmann Constant

E = $n \times \frac{1}{2}$ KT $\times$ N = $\frac{n}{2} RT$

As,

Specific heat at constant volume,

$$C_V = \frac{dE}{dT}$$

$$C_V = \frac{d}{dT}\left(\frac{n}{2} RT\right)$$

$$C_V = \frac{n}{2} R$$

Now Specific heat at constant Pressure, $C_P = C_V + R$

$$C_P = \frac{n}{2}R + R$$

$$C_P = \left(\frac{n}{2} + 1\right)R$$

As, $Y = \frac{C_P}{C_V}$

$$Y = \frac{\left(\frac{n}{2} + 1\right)R}{\frac{n}{2}R}$$

$$Y = \left(\frac{n}{2} + 1\right) \times \frac{2}{n}$$

$$Y = \frac{\cancel{2}}{\cancel{n}} \times \frac{\cancel{n}}{\cancel{2}} + 1 \times \frac{2}{n}$$

$$Y = \left(1 + \frac{2}{n}\right)$$

3.Two perfect gases at absolute temperature T_1 and T_2 are mixed. There is no loss of energy. Find the temperature of the mixture if the masses of molecules are m_1 and m_2 and number of molecules is n_1 and n_2?

Ans. In a perfect gas, there is no mutual interaction between the molecules.

Now, K.E of gas = $\frac{1}{2}mv^2$

By equi partition of energy:

$$\frac{1}{2}mv^2 = \frac{3}{2}KT.$$

K.E of one gas = $n_1 \times \left(\frac{3}{2}KT_1\right) \rightarrow 1)$

K.E. of other gas = $n_2 \times \left(\frac{3}{2}KT_2\right) \rightarrow 2)$

n_1, n_2 = Number of molecules in gases

K = Bolt zman' Constant

T_1, T_2 → Temperatures.

Total K.E. = $\frac{3}{2}K\left(n_1T_1 + n_2T_2\right)$ (adding equation 1)&2)

Let T be the absolute temperature of the mixture of gases

Then,

Total Kinetic energy = $n_1 \times \left(\frac{3}{2}KT\right) + n_2 \times \left(\frac{3}{2}KT\right)$

Total K.E = $\frac{3}{2}KT\left(n_1 + n_2\right) \rightarrow 4)$

Since there is no loss of energy, hence on equating eq^4 3) &4) for total K.E.:→

$$\frac{\cancel{3}}{\cancel{2}}\cancel{K}T\left(n_1 + n_2\right) = \frac{\cancel{3}}{\cancel{2}}\cancel{K}\left(n_1T_1 + n_2T_2\right)$$

$$T\left(n_1 + n_2\right) = \left(n_1T_1 + n_2T_2\right)$$

$$T = \frac{n_1T_1 + n_2T_2}{n_1 + n_2}$$

4.Derive Avogadro's law?

Ans. Avogadro's law states that equal volumes of all gases under identical conditions of temperature and pressure, contain the same number of molecules consider two gas having equal volumes 'V' at temperature 'T' and pressure 'P'.

Let M_1=Mass of first gas

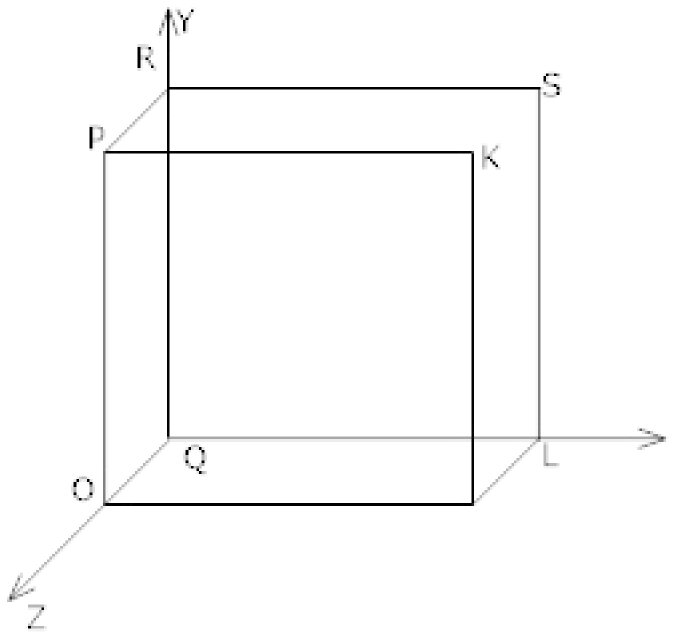
M_2=Mass of second gas

C_1=C_2=r.m.s velocity of gas molecules of 2 gases m1/ m2 = Mass of each molecule of gas

M_1, m_2=Number of molecules of gas

Now, M_1=m_1 n_1 and M_2=m_2 n_2

From kinetic theory of gas :-



For first gas $\Rightarrow P = \frac{1M_1}{3V} C_1^2 \rightarrow (1)$

For second gas $\Rightarrow P = \frac{1M_2}{3V} C_2{}^2 \rightarrow (2)$

Equating equation 1) &2) for pressure

$$\frac{1M_1}{3V} C_1^2 = \frac{1M_2}{3V} C_2{}^2$$

M_1 $C_1{}^2$ = $M_2{}^2$ $C_2{}^2$ →3)

$$\frac{Average K.E}{Molecule \text{ of first gas}} = \frac{Average \text{ K. E}}{Molecule of \sec ond \ gas} \Rightarrow for \ same \text{ temperatures}$$

$$\frac{1}{2} M_1 C_1^2 = \frac{1}{2} M_2 C_2{}^2$$

M_1 $C^2{}_1$ = M_2 $C_2{}^2$ →4)

Let C1, C2 ------ Cn = Random velocities of gases molecules

Let (x_1, y_1, z_1) ----- (xn, yn, zn) → Radom rectangular co-ordinates of η – molecules

So, $\left. \begin{array}{l} x_1{}^2 + y_1{}^2 + z_1{}^2 = c_1{}^2 \\ x_n{}^2 + y_n{}^2 + z_n{}^2 = c_n{}^2 \end{array} \right]$

A)

Initial Molomentum of A, = mx_1

on collision with wall, Momentum = -mx,

Change in Momentum = $-mx_1-mx_1$

= $-2mx_1$

The molecule in between the collisions of two walls OPKT and QRSL covers a distance = 2a time between 2 collisions = $\frac{2a}{x_1}$

Momentum transferred in 1 second = 2m x_1 x $\frac{x_1}{2a} = \frac{mx_1^2}{a}$

From Newton's second law $\Rightarrow f_1 = \frac{mx_1^2}{a}$

$$fn = \frac{mx_n^2}{a}$$

Total force in X-direction = f_1+f_2+ ------ fn

$$= \frac{mx_1^2}{a} + \frac{mx_2^2}{a} + ----- \frac{mx_n^2}{a}$$

Pressure exerted on wall QRSL

$$= \frac{Fx}{a^2} = \frac{m}{a^3}\left(x_1^2 + x_2^2 + ----- x_n^2\right)$$

Dividing equation 4) by 3)

$$\frac{M_1C_1^2}{m_1c_1^2} = \frac{M_2C_2^2}{m_2c_2^2} \quad \text{M=m x n}$$

$$\frac{m_1n_1}{m_1} = \frac{m_2n_2}{m_2}$$

$\Rightarrow$ Avogadro's law

$n_1 = n_2$

5.What are the assumptions of kinetic theory of gas?

Ans. The assumptions of kinetic theory of gases are:-

1) A gas consists of a very large number of molecules which should be elastic spheres and identical for a given gas.

2) The molecules of a gas are in a state of continuous rapid and random motion.

3) The size of gas molecules is very small as compared to the distance between them.

4) The molecules do not exert any force of attraction or repulsion on each other.

5) The collisions of molecules with one another and with walls of the vessel are perfectly elastic.

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6. Estimate the fraction of molecular volume to the actual volume occupied by oxygen gas at STP. Take the diameter of an oxygen molecule to be $3\ \mathring{A}$.

Ans. Diameter of an oxygen molecule, *d*= $3\ \mathring{A}$

Radius, $r = \frac{d}{2} = \frac{3}{2}$ = 1.5 $\mathring{A}$ = $1.5\times10^{-8} cm$

Actual volume occupied by 1 mole of oxygen gas at STP = 22400 cm^3

Molecular volume of oxygen gas, $V = \frac{4}{3}\pi r^3 \cdot N$

Where, *N* is Avogadro's number = 6.023 $\times$ 1023 molecules/mole

$\therefore V = \frac{4}{3}3.14\times\left(1.5\times10^{-8}\right)^3 \times 6.023\times10^{23} = 8.51 cm^3$

Ratio of the molecular volume to the actual volume of oxygen = $\frac{8.51}{22400}$

= 3.8×10^{-4}

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7. Estimate the total number of air molecules (inclusive of oxygen, nitrogen, water vapour and other constituents) in a room of capacity 25.0 m^3 at a temperature of 27 °C

and 1 atm pressure.

Ans. Volume of the room, *V*= 25.0 m^3

Temperature of the room, *T*= 27°C = 300 K

Pressure in the room, *P*= 1 atm = $1 \times 1.013 \times 10^5 \, Pa$

The ideal gas equation relating pressure (*P*), Volume (*V*), and absolute temperature (*T*) can be written as:

$PV = k_B NT$

Where,

K_B is Boltzmann constant = $1.38 \times 10^{-23} \, m^2 \, kg \, s^{-2} K^{-1}$

N is the number of air molecules in the room

$$\therefore N = \frac{PV}{k_B T}$$

$$= \frac{1.013 \times 10^5 \times 25}{1.38 \times 10^{-23} \times 300}$$

= 6.11×10^{26} molecules

Therefore, the total number of air molecules in the given room is 6.11×10^{26}.

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8. From a certain apparatus, the diffusion rate of hydrogen has an average value of 28.7 cm3 s^{-1}. The diffusion of another gas under the same conditions is measured to have an average rate of $7.2 \, cm^3 s^{-1}$. Identify the gas.

[Hint: Use Graham's law of diffusion: $R_1 / R_2 = (M_2 / M_1)^{1/2}$, where R_1, R_2 are diffusion rates of gases 1 and 2, and M_1 and M_2 their respective molecular masses. The law is a simple consequence of kinetic theory.]

Ans. Rate of diffusion of hydrogen, $R_1 = 28.7\ cm^3 s^{-1}$

Rate of diffusion of another gas, $R_2 = 7.2\ cm^3 s^{-1}$

According to Graham's Law of diffusion, we have:

$$\frac{R_1}{R_2} = \sqrt{\frac{\mathbf{M}_2}{\mathbf{M}_1}}$$

Where,

M_1 is the molecular mass of hydrogen = 2.020 g

M_2 is the molecular mass of the unknown gas

$$\therefore \mathbf{M}_2 = \mathbf{M}_1 \left(\frac{R_1}{R_2}\right)^2$$

$$= 2.02\left(\frac{28.7}{7.2}\right)^2 = 32.09 \text{ g}$$

32 g is the molecular mass of oxygen. Hence, the unknown gas is oxygen.

4 Marks Questions

1. Molar volume is the volume occupied by 1 mol of any (ideal) gas at standard temperature and pressure (STP: 1 atmospheric pressure, 0 °C). Show that it is 22.4 liters.

Ans: The ideal gas equation relating pressure (*P*), volume (*V*), and absolute temperature (*T*) is given as:

PV= net

Where,

R is the universal gas constant = $8.314\,J\,mol^{-1}K^{-1}$

n= Number of moles = 1

T= Standard temperature = 273 K

P= Standard pressure = 1 atm = $1.013\times10^{5}\,Nm^{-2}$

$$\therefore V = \frac{nRT}{P}$$

$$= \frac{1\times8.314\times273}{1.013\times10^{5}}$$

= 0.0224 m^3

= 22.4 liters

Hence, the molar volume of a gas at STP is 22.4 liters.

---

2. Three vessels of equal capacity have gases at the same temperature and pressure. The first vessel contains neon (monatomic), the second contains chlorine (diatomic), and the third contains uranium hexafluoride (polyatomic). Do the vessels contain equal number of respective molecules? Is the root mean square speed of molecules the same in the three cases? If not, in which case is the largest?

Ans. Yes. All contain the same number of the respective molecules.

No. The root mean square speed of neon is the largest.

Since the three vessels have the same capacity, they have the same volume.

Hence, each gas has the same pressure, volume, and temperature.

According to Avogadro's law, the three vessels will contain an equal number of the respective molecules. This number is equal to Avogadro's number, $N = 6.023 \times 10^{23}$.

The root mean square speed (v_{rms}) of a gas of mass *m*, and temperature *T*, is given by the relation:

$$v_{rms} = \sqrt{\frac{3kT}{m}}$$

Where, *k* is Boltzmann constant

For the given gases, *k* and *T* are constants.

Hence v_{rms} depends only on the mass of the atoms, i.e.,

$$v_{rms} \propto \sqrt{\frac{1}{m}}$$

Therefore, the root mean square speed of the molecules in the three cases is not the same. Among neon, chlorine, and uranium hexafluoride, the mass of neon is the smallest. Hence, neon has the largest root mean square speed among the given gases.

---

3. A metre long narrow bore held horizontally (and closed at one end) contains a 76 cm long mercury thread, which traps a 15 cm column of air. What happens if the tube is held vertically with the open end at the bottom?

Ans. Length of the narrow bore, $L = 1$ m = 100 cm

Length of the mercury thread, $l = 76$ cm

Length of the air column between mercury and the closed end, l_a = 15 cm

Since the bore is held vertically in air with the open end at the bottom, the mercury length that occupies the air space is: 100–(76 + 15) = 9 cm

Hence, the total length of the air column = 15 + 9 = 24 cm

Let *h* cm of mercury flow out as a result of atmospheric pressure.

∴Length of the air column in the bore= 24 + *h* cm

And, length of the mercury column = 76 – *h* cm

Initial pressure, P_1 = 76 cm of mercury

Initial volume, V_1 = 15 cm^3

Final pressure, P_2 = 76 –(76– *h*) = *h* cm of mercury

Final volume, V_2 = (24 + *h*) cm^3

Temperature remains constant throughout the process.

$$\therefore P_1V_1 = P_2V_2$$

76 $\times$ 15 = *h* (24 + *h*)

h_2 + 24*h* –1140 = 0

$$\therefore h = \frac{-24 \pm \sqrt{(24)^2 + 4 \times 1 \times 1140}}{2 \times 1}$$

= 23.8 cm or – 47.8 cm

Height cannot be negative. Hence, 23.8 cm of mercury will flow out from the bore and 52.2 cm of mercury will remain in it. The length of the air column will be 24 + 23.8 = 47.8 cm.

---

4. An air bubble of volume 1.0 cm3rises from the bottom of a lake 40 m deep at a temperature of 12 °C. To what volume does it grow when it reaches the surface, which is at a temperature of 35 °C?

Ans. Volume of the air bubble, V_1 = $1.0\ cm^3 = 1.0\times10^{-6}\ m^3$

Bubble rises to height, *d* = 40 m

Temperature at a depth of 40 m, T_1 = 12°C = 285 K

Temperature at the surface of the lake, T_2 = 35°C = 308 K

The pressure on the surface of the lake:

P_2 = 1 atm = $1\times1.013\times10^5\ Pa$

The pressure at the depth of 40 m:

P_1 = 1 atm + $d\rho g$

Where,

ρ is the density of water = $10^3\ kg/m^3$

g is the acceleration due to gravity = $9.8\ m/s^2$

$\therefore P_1$ = $1.013\times10^5 + 40\times10^3\times9.8$

= 493300 Pa

We have: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Where, V_2 is the volume of the air bubble when it reaches the surface

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

$= \frac{(493300)(1.0\times10^{-6})308}{285\times1.013\times10^{5}}$

= $5.263\times10^{-6}\ m^3\ or\ 5.263\ cm^3$

Therefore, when the air bubble reaches the surface, its volume becomes 5.263 cm^3.

---

5. Estimate the average thermal energy of a helium atom at (i) room temperature (27 °C), (ii) the temperature on the surface of the Sun (6000 K), (iii) the temperature of 10 million Kelvin (the typical core temperature in the case of a star).

Ans.(i) At room temperature, T= 27°C = 300 K

Average thermal energy $= \frac{3}{2}kT$

Where k is Boltzmann constant = $1.38\times10^{-23}\ m^2\ kg\ s^{-2}K^{-1}$

$\therefore \frac{3}{2}kT = \frac{3}{2}\times1.38\times10^{-38}\times300$

= $6.21\times10^{-21}J$

Hence, the average thermal energy of a helium atom at room temperature (27°C) is $6.21\times10^{-21}J$.

(ii) On the surface of the sun, T= 6000 K

Average thermal energy $= \frac{3}{2}kT$

$\frac{3}{2}\times1.38\times10^{-38}\times6000$

= $1.241\times10^{-19}J$

Hence, the average thermal energy of a helium atom on the surface of the sun is $1.241\times10^{-19}\,J$.

(iii) At temperature, *T*= 107K

Average thermal energy $=\frac{3}{2}kT$

$\frac{3}{2}\times1.38\times10^{-38}\times10^{7}$

= $2.07\times10^{-16}\,J$

Hence, the average thermal energy of a helium atom at the core of a star is $2.07\times10^{-16}\,J$.

---

6. At what temperature is the root mean square speed of an atom in an argon gas cylinder equal to the rms speed of a helium gas atom at - 20 °C? (atomic mass of Ar = 39.9 u, of He = 4.0 u).

Ans. Temperature of the helium atom, T_{He} = –20°C= 253 K

Atomic mass of argon, M_{Ar} = 39.9 u

Atomic mass of helium, M_{He} = 4.0 u

Let, $(v_{rms})_{Ar}$ be the rms speed of argon.

Let $(v_{rms})_{He}$ be the rms speed of helium.

The rms speed of argon is given by:

$$(v_{rms})_{Ar}=\sqrt{\frac{3RT_{Ar}}{M_{Ar}}}\ \dots(i)$$

Where,

R is the universal gas constant

T_{Ar} is temperature of argon gas

The rms speed of helium is given by:

$$(v_{rms})_{He} = \sqrt{\frac{3RT_{He}}{M_{He}}} \quad \ldots (ii)$$

It is given that:

$$(v_{rms})_{Ar} = (v_{rms})_{He}$$

$$\sqrt{\frac{3RT_{Ar}}{M_{Ar}}} = \sqrt{\frac{3RT_{He}}{M_{He}}}$$

$$\frac{T_{Ar}}{M_{Ar}} = \frac{T_{He}}{M_{He}}$$

$$T_{Ar} = \frac{T_{He}}{M_{He}} \times M_{Ar}$$

$$= \frac{253}{4} \times 39.9$$

= 2523.675 = $2.52 \times 10^3\, K$

Therefore, the temperature of the argon atom is $2.52 \times 10^3\, K$.

5 Marks Questions

1.Derive an expression for the pressure due to an ideal gas?

Ans. Consider an ideal gas contained in a cubical container OPQ RSTKL, each of side a, having volume V now, $V=a^3$ ($(Side)^3$ = volume of cube)

Let n = Molecule of gas

m = Mass of each molecule

M = m x n = Mass of gas

Similarly Py = $\frac{m}{a^3}\left(y_1^2+y_2^2+--y_n^2\right)$

$$Pz=\frac{m}{a^3}\left(z_1^2+z_2^2+--z_n^2\right)$$

P = Total pressure = $\frac{P_x+P_y+P_z}{3}$

$$=\frac{1}{3}\left[\frac{m}{a^3}\left(x_1^2+x_2^2+--x_n^2\right)+\frac{m}{a^3}\left(y_1^2+y_2^2+--y_n^2\right)+\frac{m}{a^3}\left(z_1^2+z_2^2+--z_n^2\right)\right]$$

P = $\frac{m}{3a^3}\left[\left(x_1^2+y_1^2+z_1^2\right)+---\left(x_n^2+y_n^2+z_n^2\right)\right]$ (∵ from equation A)

$$P=\frac{m}{3v}\left[C_1^2+C_2^2+--C_n^2\right]$$

→ Multiply & divide by n (no of molecules of gas)

$$P=\frac{1mxn}{3v}\left[\frac{C_1^2+C_2^2+--C_n^2}{n}\right]$$

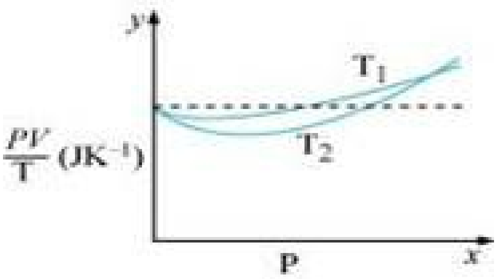
$$P = \frac{1M}{3v}C^2$$

$$c^2 = \frac{C_1^2 + C_2^2 + -- C_n^2}{n} \text{ or } C = \sqrt{\frac{C_1^2 + C_2^2 + -- C_n^2}{n}}$$

C = r. m s. velocity of gas.

---

2. Figure 13.8 shows plot of *PV*/*T* versus Poor 1.00×10^{-3} kg of oxygen gas at two different temperatures.



(a) What does the dotted plot signify?

(b) Which is true: $T_1 > T_2$ or $T_1 < T_2$?

(c) What is the value of *PV*/*T* where the curves meet on the *y*-axis?

(d) If we obtained similar plots for 1.00×10^{-3} kg of hydrogen, would we get the same value of *PV*/*T* at the point where the curves meet on the *y*-axis? If not, what mass of hydrogen yields the same value of *PV*/*T* (for low pressure high temperature region of the plot)? (Molecular mass of H_2 = 2.02 u, of O2= 32.0 u, *R* = $8.31 \text{ J mol}^{-1}\text{K}^{-1}$.)

Ans.(a) The dotted plot in the graph signifies the ideal behavior of the gas, i.e., the ratio $\frac{PV}{T}$ is equal. μR (μ Is the number of moles and R is the universal gas constant) is a constant quality. It is not dependent on the pressure of the gas.

(b) The dotted plot in the given graph represents an ideal gas. The curve of the gas at temperature T_1 is closer to the dotted plot than the curve of the gas at temperature T_2 . A real gas approaches the behaviour of an ideal gas when its temperature increases. Therefore, $T_1 > T_2$ is true for the given plot.

(c) The value of the ratio *PV*/*T*, where the two curves meet, is μR. This is because the ideal gas equation is given as:

$PV = \mu RT$

$$\frac{PV}{T} = \mu R$$

Where,

P is the pressure

T is the temperature

V is the volume

μ is the number of moles

R is the universal constant

Molecular mass of oxygen = 32.0 g

Mass of oxygen = 1×10^{-3} kg = 1 g

R = 8.314 J $mole^{-1}$ K^{-1}

$$\therefore \frac{PV}{T} = \frac{1}{32} \times 8.314$$

= 0.26 J K^{-1}

Therefore, the value of the ratio *PV*/*T*, where the curves meet on the *y*-axis, is

0.26 J K^{-1}.

(d) If we obtain similar plots for 1.00×10^{-3} kg of hydrogen, then we will not get the same value of *PV*/*T* at the point where the curves meet the *y*-axis. This is because the molecular mass of hydrogen (2.02 u) is different from that of oxygen (32.0 u).

We have:

$$\frac{PV}{T}=0.26J\,K^{-1}$$

R = 8.314 J $mole^{-1}\;K^{-1}$

Molecular mass (*M*) of H_2 = 2.02 u

$\frac{PV}{T}=\mu R$ at constant temperature

Where, $\mu=\frac{m}{M}$

m = Mass of H_2

$$\therefore m=\frac{PV}{T}\times\frac{M}{R}$$

$$=\frac{0.26\times2.02}{8.31}$$

= $6.3\times10^{-2}\;g=6.3\times10^{-5}\;kg$

Hence, 6.3×10^{-5} kg of H_2 will yield the same value of *PV*/*T*.

---

3. An oxygen cylinder of volume 30 litres has an initial gauge pressure of 15 atm and a temperature of 27 °C. After some oxygen is withdrawn from the cylinder, the gauge pressure drops to 11 atm and its temperature drops to 17 °C. Estimate the mass of oxygen taken out of the cylinder (*R* = 8.31 J $mol^{-1}\;K^{-1}$, molecular mass of O_2 = 32 u).

Ans. Volume of oxygen, V_1 = 30 litres = $30\times10^{-3}m^3$

Gauge pressure, P_1 = 15 atm = $15 \times 1.013 \times 10^5\ Pa$

Temperature, T_1 = 27°C = 300 K

Universal gas constant, R= 8.314 J $\text{mol}^{-1}\ \text{K}^{-1}$

Let the initial number of moles of oxygen gas in the cylinder be n_1.

The gas equation is given as:

$$P_1V_1 = n_1RT_1$$

$$\therefore n_1 = \frac{P_1V_1}{RT_1}$$

$$= \frac{15.195 \times 10^5 \times 30 \times 10^{-3}}{(8.314) \times 300} = 18.276$$

But, $n_1 = \frac{m_1}{M}$

Where,

m_1 = Initial mass of oxygen

M= Molecular mass of oxygen = 32 g

$\therefore m_1 = n_1 M$ = 18.276 $\times$ 32 = 584.84 g

After some oxygen is withdrawn from the cylinder, the pressure and temperature reduces.

Volume, V_2 = 30 litres = $30 \times 10^{-3} m^3$

Gauge pressure, P_2 = 11 atm = $11 \times 1.013 \times 10^5\ Pa$

Temperature, T_2 = 17°C = 290 K

Let n_2 be the number of moles of oxygen left in the cylinder.

The gas equation is given as:

$P_2V_2 = n_2RT_2$

$\therefore n_2 = \frac{P_2V_2}{RT_2}$

$= \frac{11.143\times10^5\times30\times10^{-3}}{8.314\times290}$ = 13.86

But, $n_2 = \frac{m_2}{M}$

Where,

m_2 is the mass of oxygen remaining in the cylinder

$\therefore m_2 = n_2$ M = 13.86 $\times$ 32 = 453.1 g

The mass of oxygen taken out of the cylinder is given by the relation:

Initial mass of oxygen in the cylinder – Final mass of oxygen in the cylinder

= $m_1 - m_2$

= 584.84 g– 453.1 g

= 131.74 g

= 0.131 kg

Therefore, 0.131 kg of oxygen is taken out of the cylinder.

---

4. Estimate the mean free path and collision frequency of a nitrogen molecule in a cylinder containing nitrogen at 2.0 atm and temperature 17 °C. Take the radius of a nitrogen molecule to be roughly 1.0 $\mathring{A}$. Compare the collision time with the time the molecule moves freely between two successive collisions (Molecular mass of $N_2 = 28.0$ u).

Ans. Mean free path = $1.11\times10^{-7}\,m$

Collision frequency = $4.58\times10^{9}s^{-1}$

Successive collision time $\approx$ 500 $\times$ (Collision time)

Pressure inside the cylinder containing nitrogen, *P*= 2.0 atm = $2.026\times10^{5}\,Pa$

Temperature inside the cylinder, *T*= $\overset{\circ}{A}$ 17°C =290 K

Radius of a nitrogen molecule, *r*= 1.0 = $1\times10^{10}\,m$

Diameter, *d* = $2\times1\times10^{10}=2\times10^{10}\,m$

Molecular mass of nitrogen, *M*= 28.0 g = $28\times10^{-3}\,kg$

The root mean square speed of nitrogen is given by the relation:

$$v_{rms}=\sqrt{\frac{3RT}{M}}$$

Where,

R is the universal gas constant = 8.314 J $mole^{-1}K^{-1}$

$$\therefore v_{rms}=\sqrt{\frac{3\times8.314\times290}{28\times10^{-3}}}$$ = 508.26 m/s

The mean free path (*l*) is given by the relation:

$$l=\frac{kT}{\sqrt{2}\times d^2\times P}$$

Where,

K is the Boltzmann constant = $1.38\times10^{-23}\,kg\ m^2s^{-2}K^{-1}$

$$\therefore l=\frac{1.38\times10^{-23}\times290}{\sqrt{2}\times3.14\times\left(2\times10^{-10}\right)^2\times2.026\times10^{5}}$$

$= 1.11 \times 10^{-7} m$

Collision frequency $= \frac{v_{rms}}{l}$

$$= \frac{508.26}{1.11 \times 10^{-7}} = 4.58 \times 109 s^{-1}$$

Collision time is given as:

$$T = \frac{d}{v_{rms}}$$

$$= \frac{2 \times 10^{-10}}{508.26} = 3.93 \times 10^{-13} s$$

Time taken between successive collisions:

$$T' = \frac{l}{v_{rms}}$$

$$= \frac{1.11 \times 10^{-7}}{508.26 m/s} = 2.18 \times 10^{-10} s$$

$$\therefore \frac{T'}{T} = \frac{2.18 \times 10^{-10}}{3.93 \times 10^{-13}} \simeq 500$$

Hence, the time taken between successive collisions is 500 times the time taken for a collision.

---

5. A gas in equilibrium has uniform density and pressure throughout its volume. This is strictly true only if there are no external influences. A gas column under gravity, for example, does not have uniform density (and pressure). As you might expect, its density decreases with height. The precise dependence is given by the so-called law of atmospheres

$$n_2 = n_1 \exp \left[-mg(h_2 - h_1)/k_B T \right]$$

Where $n/_{,} n_1$ refer to number density at heights h_2 and h_1 respectively. Use this relation to derive the equation for sedimentation equilibrium of a suspension in a liquid column:

$$n_2 = n_1 \mathbf{exp} \left[-mg\ N_A(\rho - P')\ (h_2 - h_1)/\ (\rho RT)\right]$$

Where ρ is the density of the suspended particle, and ρ' that of surrounding medium. [N_A is Avogadro's number, and *R* the universal gas constant.] [Hint: Use Archimedes principle to find the apparent

Ans. According to the law of atmospheres, we have:

$$n_2 = n_1 exp\left[-mg(h_2 - h_1)/k_B T\right] \text{... } (i)$$

Where,

n_1 is the number density at height $h_{1,} and\ n_2$ is the number density at height h_2

mg is the weight of the particle suspended in the gas column

Density of the medium = ρ '

Density of the suspended particle = ρ

Mass of one suspended particle = *m*'

Mass of the medium displaced = *m*

Volume of a suspended particle = *V*

According to Archimedes' principle for a particle suspended in a liquid column, the effective weight of the suspended particle is given as:

Weight of the medium displaced – Weight of the suspended particle

= *mg*– *m*'*g*

$$= mg - V\rho' g = mg - \left(\frac{m}{\rho}\rho' g\right)$$

$$mg\left(1-\frac{\rho'}{\rho}\right) \text{...........(ii)}$$

Gas constant, R = $k_B N$

$$k_B = \frac{R}{N} \text{ ... } (iii)$$

Substituting equation (*ii*) in place of *mg* in equation (*i*) and then using equation (*iii*), we get:

$$n_2 = n_1 exp\left[-mg(h_2 - h_1)/k_B T\right]$$

$$= n_1 \exp\left[-mg\left(1-\frac{\rho'}{\rho}\right)(h_2 - h_1)\frac{N}{RT}\right]$$

$$= n_1 \exp\left[-mg(\rho-\rho')(h_2 - h_1)\frac{N}{RT\rho}\right]$$

---

6. Given below are densities of some solids and liquids. Give rough estimates of the size of their atoms:

Substance	Atomic Mass (u)	Density (10^3 Kg m^{-3})
Carbon (diamond)	12.01	2.22
Gold	197.00	19.32
Nitrogen (liquid)	14.01	1.00
Lithium	6.94	0.53
Fluorine (liquid)	19.00	1.14

[Hint: Assume the atoms to be 'tightly packed' in a solid or liquid phase, and use the known value of Avogadro's number. You should, however, not take the actual numbers you obtain for various atomic sizes too literally. Because of the crudeness of the tight packing approximation, the results only indicate that atomic sizes are in the range of a few $\overset{\circ}{A}$].

Ans.

Substance	Radius ($\overset{\circ}{A}$)
Carbon (diamond)	1.29
Gold	1.59
Nitrogen (liquid)	1.77
Lithium	1.73
Fluorine (liquid)	1.88

Atomic mass of a substance = M

Density of the substance = ρ

Avogadro's number = N= 6.023×10^{23}

Volume of each atom $=\frac{4}{3}\pi r^3$

Volume of N number of molecules $=\frac{4}{3}\pi r^3 N$ ... (*i*)

Volume of one mole of a substance = $\frac{M}{\rho}$ ... (*ii*)

$$=\frac{4}{3}\pi r^3$$

$$N = \frac{M}{\rho}$$

$$\therefore r = \sqrt[3]{\frac{3M}{4\pi\rho N}}$$

For carbon:

$$M = 12.01\times10^{-3}\, kg,$$

$$\rho = 2.22\times10^{3}\, kg\ m^{-3}$$

$$\therefore r = \left(\frac{3\times12.1\times10^{-3}}{4\pi\times2.22\times10^{3}\times6.023\times10^{23}}\right)^{\frac{1}{3}} = 1.29\,\overset{\circ}{A}$$

Hence, the radius of a carbon atom is 1.29 $\overset{o}{A}$.

For gold:

$M = 197.00 \times 10^{-3}\ kg$

$\rho = 19.32 \times 10^{3}\ kg\ m^{-3}$

$$\therefore r = \left(\frac{3 \times 197 \times 10^{-3}}{4\pi \times 19.32 \times 10^{3} \times 6.023 \times 10^{23}} \right)^{\frac{1}{3}} = 1.59\ \overset{o}{A}$$

Hence, the radius of a gold atom is 1.59 $\overset{o}{A}$.

For liquid nitrogen:

$M = 14.01 \times 10^{-3}\ kg$

$\rho = 1.00 \times 10^{3}\ kg\ m^{-3}$

$$\therefore r = \left(\frac{3 \times 14.01 \times 10^{-3}}{4\pi \times 1.00 \times 10^{3} \times 6.23 \times 10^{23}} \right)^{\frac{1}{3}} = 1.77\ \overset{o}{A}$$

Hence, the radius of a liquid nitrogen atom is 1.77 $\overset{o}{A}$.

For lithium:

$M = 6.94 \times 10^{-3}$ kg

$p = 0.53 \times 10^{3}$ kg m^{-3}

$$\therefore r = \left(\frac{3 \times 6.94 \times 10^{-3}}{4\pi \times 0.53 \times 10^{3} \times 6.23 \times 10^{23}} \right)^{\frac{1}{3}} = 1.73\ \overset{o}{A}$$

Hence, the radius of a lithium atom is 1.73 $\overset{o}{A}$.

For liquid fluorine:

$M = 19.00 \times 10^{-3}\ kg$

$\rho = 1.14 \times 10^{3}\ kg$