**Generally real gases obey ideal gas law PV = nRT only at very high temperature and very low pressure**

**Boyle's Law**

`P\_{1}V\_{1} = P\_{2}V\_{2}`

`PV = nRT`

`V = (1 \* 8.314 \* 273)/(1.013 \* 10 ^ 5)`

**Thus we can state that one mole of all gases occupy a volume of 22.4 litre at standard pressure and temperature.**

**Charles and Gay-Lussac law**

`V\_{1}/T\_{1} = V\_{2}/T\_{2}`

`V/T = constant`

As with Boyle's law, the amount of gas also must be held constant for equation

**Ideal Gas Law**

`PV=nRT`

`(P\_{1}V\_{1})/T\_{1} = (P\_{2}V\_{2})/T\_{2}`

As number of moles in initial and final state are equal

**Avogadro's Number and Avogadro Hypothesis**

Avogadro stated that equal volumes of gas at the same pressure and temperature contain equal numbers of molecules.

`N\_{A} = 6.023 \* 10 ^ 23 molecules/mole`

Thus in n moles of a gas total number of molecules of the gas are

`N = nN\_{A}`

or number of moles of a gas can be given as `n = N/N\_{A} = (m' N)/(m' N\_{A}) = m/M `

Here m' is the mass of each molecule, m is the total mass of gas and M is the mass of 1 mole of molecules of gas i.e. its molecular mass.

From gas law, we have

`PV = nRT`

`PV = N/N\_{A} RT= m/M RT`

`PV= NkT`

Where `k= R /N\_{AV}` is the Boltzmann's constant and has the value `k = 1.38 \* 10 ^ - 23 \* J/K`

This equation is another form of ideal gas law which is in terms of total number of molecules of gas N instead of moles n.

**Dalton's Law of Partial Pressures**

This law states that the total pressure of a mixture of gases is equal to the sum of the partial pressures of the gases in the mixture. The total pressure of the mixture is given by

`P\_{T} = P\_{1} + P\_{2} ..... P\_{n} `

**Different Forms of Ideal Gas Law**

`PV = nRT`

Where in are the number of moles of gas in the enclosed volume V. If m mass of a gas (molecular mass = M) is taken in the container of volume V, we have number of moles of gas as

`n = m/M `

`PV = m/M RT

`P = m/(VM) RT `

`P = (rhoRT)/M `

It is a modified form of ideal gas law which relates gas pressure, its density and temperature in a physical state.

PV curve is almost horizontal and it happens only at this point. This point C is called **critical point** and the gas temperature is called **critical temperature** of the gas.

Critical temperature of a gas is defined as - "A gas will change to liquid phase if its temperature is less then critical temperature if sufficient pressure is applied on the gas". Above critical temperature, no matters whatever be the amount of pressure applied, the gas

can not be liquefied. As pressure is increased, gas becomes denser and denser and acquires the properties resembling to a liquid but does not condense.

A distinction is made between "gas" and "vapour", students should keep in mind that a substance below its critical temperature in gaseous state is called "vapour" (which can be liquefied on applying sufficient pressure on it) and a substance above its critical temperature in gaseous state is called "gas" (which can not be liquefied).

**Pressure of Air**

a Torricellian tube, the principle of today's barometer. If a glass tube, closed at one end, is completely filled with mercury and then inverted into a bowl of mercury, the column of mercury in the tube drops until it reaches a height of about 76 cm above the lower surface

This is how barometer works and the height of mercury column in the tube gives the atmospheric pressure of air.   
If cross- sectional area of tube is A then the weight of mercury column is

`W = hArhog `

Thus pressure at its bottom is given as

`p = W/A = hrhog`

In general the height of mercury column in open atmosphere in standard conditions is about 76 cm, thus atmospheric pressure of air is given as

`p = hrhog

= (0.76)(13.6 \* 10 ^ 3)(9.81) W / (m ^ 2)

P\_{air} =1.01\*10^ 5 Pa = 1atm`

This pressure is atmospheric pressure. An atmosphere (1atm) is a unit of pressure equal to the pressure of the earth's atmosphere at sea level.

Maxwell developed a relation to find atmospheric pressure as a function of height from ground level (more precisely sea level). The relation is named **Barometric Relation.** We first derive the relation mathematically (in the book) and than we'll discuss some uses and applications of this relation.

Final result

`P = P\_{0} e ^ (- (Mgh)/(RT)) `

Here `P\_{0}` is the atmospheric pressure on ground-level or sea- level above equation gives the pressure of atmospheric air

at a height h above the earth surface and is called as Barometric formula. Using this we can also find the atmospheric air density at a height h above the ground level.

`rho = rho\_{0}e ^ (- (Mgh)/(RT)) `

(derivation in book)

`rho\_{0} = (P\_{0}M)/(RT)` can be taken as atmosphere air density at the ground level. Another form of Barometric formula can be written in terms of number of molecules per unit volume or the molecular density in atmosphere and can be given as

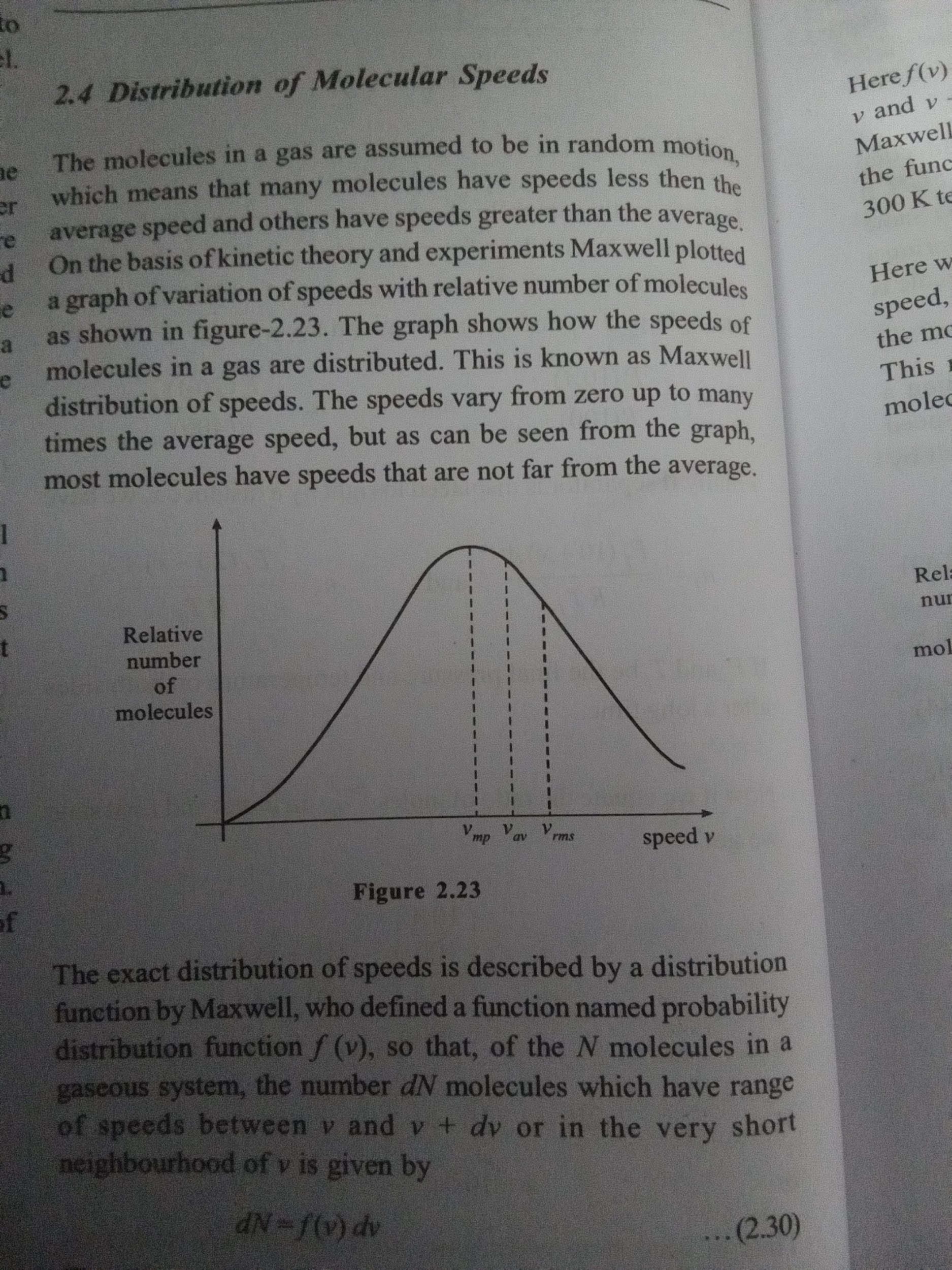
`n = n\_{0} e ^ (- (Mgh)/(RT)) `

Above equation are used with the assumption that the atmospheric temperature T and the acceleration due to gravity g remains constant within the limits of integration of expression in equation-(1). If in some region temperature gradient (dT/dh) exist then the above formula can be modified while integrating the expression in equation-(1).

**Distribution of Molecular Speeds**

- The molecules in a gas are assumed to be ramndom motion,.  
 The graph shows how the speeds of molecules in a gas are distributed. This is known as Maxwell distribution of speeds.

&image&



The exact distribution of speeds is described by a distribution function by Maxwell, who defined a function named probability distribution function f(v) so that, of the N molecules in a gaseous system, the number dN molecules which have range of speeds between v and v + dv or in the very short neighbourhood of v is given by

`dN = f(v) dv`

Thus f(v) is the number of molecules per unit range of speeds. As we have discussed that the speeds of molecules vary in a wide range from zero to infinity, we can write for total number of molecules as

`N = int\_0^infty f(v)dv`

The distribution function(v) was obtained by statical analysis and found out as

`f(v) = 4piNv^2[M/(2piRT)]^(3/2) e ^ (- (Mv^2)/(2RT))`

Where R is universal gas constant and M is the molar masswe can write

`f(v)dv = 4piNv^2[M/(2piRT)]^(3/2) e ^ (- (Mv^2)/(2RT))(dv)`

Here f(v) dv is the number of molecules that have speeds between v and v + dv and this equation is called as analytical Maxwell Boltzmann distribution function..

Here we can observe that the peak of each curve represents the speed, the maximum number of molecules have and it is called the most probable speed for that gas at particular temperature. This most probable speed is exhibited by large number of molecules of a gas at a given temperature.

**Different Speeds For Molecules of a Gas**

(i) Average velocity of Gas Molecules

Initially we've discussed that all molecules of a gas in a container are in Brownian motion, thus due to randomness the directions of motion of different molecules are random and continuously changing due to repeated collisions randomly. If `vec v\_{1}, vec v\_{2}..... vec v\_{N}` are the instantaneous velocity vectors of all N molecules of a gas, the average velocity vector of these N molecules can be written as

`< vec v > = (vec v \_{1} + vec v\_{2} +......+ vec v\_{N})/ N =0`

It is obvious that as all vectors are randomly scattered in space thus their sum i.e. the numerator of expression can be approximated to zero. Thus the average velocity vector of molecules in a gas is always taken as zero.

(**ii) Root-Mean-Square Velocity of Gas Molecules**

As the name implies this is the square root of mean of squares of velocities of all the molecules of a gas. Mean of squares can be simply written as

`< v ^ 2 > =(( | vec v 1 |^ 2 +| vec v 2 |^ 2 +.......| vec v\_{N} |^ 2 )/N )

v\_{rms} = sqrt (<v^ 2 >)

=sqrt (( | vec v\_{1} |^ 2 +| vec v\_{2} |^ 2 +.....| vec v\_{N} |^ 2 )/N )`

Mathematically its value can be calculated by using the distribution function and taking the average of the square of the speed. Since f (v) dv is the number of molecules with speed v in the range dv, the mean square speed can be given as

`< v ^ 2> = (int\_0^infty v ^ 2 \* f(v) dv )/N `

Integrating the expression in equation we finally get

`< v ^ 2 >= (3RT)/M`

So that rms speed can be given as

` v\_{rms} = sqrt ((3RT)/ M)`

In most of numeric calculation and analysis of kinetic theory we use rms velocity for gas molecules.

**(iii) Mean Speed of Gas Molecules**

We have discussed that due to randomness average velocity

vector of all the molecules of a gas comes out zero. But it is not same in case of mean or average speed. We obtain the mean speed by averaging the speed of the molecules. If can be simply defined as

`<v> = (| vec v\_{1} |+| vec v\_{2} |+.....| vec v\_{N} |)/ N `

Using distribution mean speed as

`<v> = (int\_0^infty v\*f(v) dv)/N`

Integrating expression in equation

`v\_{mean} = sqrt((8RT)/(piM))`

We can see on comparing equation

`v\_{mean} <v\_{rms}`

**(iv) Most Probable Speed of Gas Molecules**

As discussed this is the speed maximum number of molecules have, and this is the speed corresponding to which distribution function has maximum or peak value, It's value can be obtained by differentiating the distribution function f(v) and setting the derivative equal to zero. The final result obtained is

`v\_{mp} = sqrt ((2RT)/ M)`

Comparing equation we can see

` v\_{mp} <v\_{mean} <v\_{rms}`

**Pressure Exerted by a Gas**

When a gas is enclosed in a rigid container, due to motion of molecules the collide with wall and create pressure. Consider N molecules of a gas enclosed in a container. The length, width and height of container are l, w and h respectively along x, y and z direction. Out of several molecules, we consider one gas molecule of mass m' which is moving at an instant with a velocity v. This velocity has three components `v\_{x} v\_{y} and v\_{z}` along x, y and z directions respectively.

`P = 1/3 (m')/v \* v ^ 2`

`v ^ 2 = v\_{x} ^ 2 + v\_{y} ^ 2 + v\_{z} ^ 2` is the root mean square (rms) velocity Where of a gas already discussed. Equation gives the average pressure on container walls due to motion of only one molecule. There are total N molecules of gas enclosed in the container. Thus total average pressure exerted by a gas on its container walls is

`P = 1/3 (m')/v v ^ 2 \* N `

P = 1/3 m/v v ^ 2` [Where m = m' N is the total mass of gas]

`P = 1/3 rho v ^ 2 `

[Where `rho = m/v` is the density of gas]

Here we can also write

`P = 2/3 (1/2 rho v ^ 2) = 2/3 e`

Here e represents the average kinetic energy of gas per unit volume or average kinetic energy density of gas pressure of a gas in a container If total number of molecules in the container are N, then the molecular density `n\_{0}` can be given as

`n\_{0} = N/V`

Thus number of collisions with a container wall per second per square meter of its surface can be written as

`N\_{c} = n\_{0}/6 v `

this gives the number of collision per unit area of container wall in contact with a gas. As collisions are elastic, the momentum transferred to wall in each collision by a molecule is

`∆p = 2 m' v `

[If m' is the mass of each molecule]

Thus momentum transferred to a container wall per second per unit of its surface area or average pressure can be given as

`P = (n\_{0}v)/6\* 2 m' v = 1/3 rho v ^ 2 [As no m' = p density of gas]`

but in derivation of above equation we haven't taken any specific shape of container. Thus, this relation of average pressure always remains same irrespective of the shape of container.

Kinetic Energy of Gas Molecules

As all collisions are considered to be perfectly elastic we can assume that the total kinetic energy of all the gas molecules remains constant and all molecules are assumed to be always moving with their rms velocity. In this situation kinetic energy of a gas molecule of mass m' can be given as

`E\_{K} = 1/2 m' v\_{rms} ^ 2`

`M=m’ N\_{Av}` [`N\_{Av}` - Avogadro Number]comparing

`N\_{A}` = (molecular mass)/ (mass of one molecule)

`E\_{K} = 3 /2 kT `

`k= R /N\_{A} =1.38\*10^ -23 Joule/molecule`, k is a universal constant called **boltzmann** **constant**. This equation gives the kinetic energy of a moving gas molecule in a gas at absolute temperature T. If we find the total kinetic energy of all the molecules of a gas, it is given as

`E\_{T} = 3/2 nRT `

The total energy of a gas is the sum of kinetic energy of all of its molecules. Equation above thus gives the total energy of a gas in its molecular motion or the expression in equation gives the total translational energy of all the gas molecules at absolute temperature T.

**Microscopic Interpretation of Temperature**

the average molecular translational kinetic energy and absolute temperature are proportional.

`T = 2/3 E\_{T}/k`

Thus the average translational kinetic energy of gas molecules in an ideal gas depends only on the temperature, not on the pressure or type of gas.

**Degrees of Freedom and Equipartition of Energy**

Thus, the average kinetic energies of a gas molecule in the three directions respectively can be given as

`1/2 m v\_{x} ^ 2 = 1/2 m v\_{y} ^ 2 = 1/2 m v\_{z} ^ 2 = 1/2 kT `

The three kinetic energies in expression of equation are the three parts of total kinetic energy of a gas molecule used corresponding to the three velocity components in the three directions in space.

“the total energy of an gas molecule is distributed equally among all of its degrees of freedom and energy associated with each degree of freedom is `1/2kT`

This is called law of equipartition of energy

**Internal Energy of a Gas**

Thus if a gas molecule has f degrees of freedom or it can move in f number of ways, the total energy of that molecule can be given as

`E = f \* 1/2 kT = f/2 kT`

If an ideal gas having f degrees of freedom and have n moles in a container than total energy of this gas is given as

`U = f/2 nRT` ` [As kN\_{AV} =R] `

The expression in equation above gives the sum of total kinetic energies of all types of motion of all the molecules of a gas, called as total internal energy of a gas.

**Degrees of Freedom and Internal Energy For Different Types of Gases**

**Monoatomic Gases**

A gas with only one atom. eg: Helium, Argon. They have 3 degrees of freedom(translational)

`U=f/2nRT=3/2nRT`

**Diatomic Gases**

Thus total mechanical energy or internal energy of an ideal diatomic gas at a temperature T is given as

`U = f/2 nRT `

At low temperatures f = 5 thus (3 translational and 2 rotational)

`U = 5/2 nRT `

At very high temperatures f = 7 thus (extra 2 vibrational degrees of freedom)

`U = 7/2 nRT`

**Polyatomic Gases**

A polyatomic gas molecule should also have three translational degrees of freedom and the number of rotational degrees of freedom depending on the geometry of the molecule. If the molecule is a linear one like CO\_{2} then it has only two rotational degrees of freedom and if it is a non linear molecule like H\_{2}O CH\_{4} etc. then it has three rotational degrees of freedom as about all three co-ordinate axes.

Thus the total internal energy of n moles of a polyatomic gas at a temperature T is given as For a linear molecule

`U = f/2 \* nRT = ((5 + x)/2) \* nRT `

Where x are the number of vibrational degrees of freedoms

For a non linear molecule which has x number of vibrational degrees of freedom

`U = f/2 \* nRT = ((6 + x)/2) \* nRT`

**Equivalent Degrees of Freedom For a Gaseous Mixture**

We know if two substances at same temperature are connected or mixed, they do not exchange any thermal energy and the temperature of mixture remains same.

`F\_{eq} = (f\_{1} n\_{1} +f\_{2} n\_{2} +…. F\_{N} n\_{N})/( n\_{1} +n\_{2} +…. n\_{N})`

**Mixing of Gases at Constant Volume**

When some gases at different temperature are mixed at constant volume in a thermally insulated vessel, the total internal energy of all the gases remains constant.

`T\_{f} =( f\_{1} n\_{1}T\_{1} + f\_{2} n\_{2}T\_{2} +......+ f\_{N} n\_{N} T\_{N})/(f\_{1} n\_{1} +f\_{2} n\_{2} +… f\_{N} n\_{N})`

Equation can be used to find final temperture of the gaseous mixture

**Path of Molecules and Mean Free Path**

`lambda = 1/(sqrt(2) m d ^ 2 (N / V))`

Here m is the mass of molecules and d is the molecular diameter. From gas law we have

`PV = nRT` or `PV = NkT` as `R/N\_{AV} = k` if N are the total number of molecules in a gas, thus equation now become

`lambda = (kT)/(sqrt(2) m d ^ 2 P)`

**Vander Walls Gas Equation**

The ideal gas equation can be used for different ideal gases but not for real gases due to the above reason. This ideal gas law can be modified for real gases and by introducing some appropriate corrections, we can develop another equation of state for a real gas. This equation was developed by Dutch physicist J.D. Vander Walls. The equation is named vander walls gas equation after him and is given as

(P+ (an^2)/ V^2 ) (V-nb)=nRT

Here a and b are empirical constants, different for different gases

(detailed explanation in book)