



ICT 1107: Physics

Dr M Mahbubur Rahman

Associate Professor

Department of Physics

Jahangirnagar University, Savar, Dhaka 1342

Email: M.Rahman@juniv.edu



ICT 1107: Physics

2. Heat and Thermodynamics

Principle of temperature measurements: platinum resistance thermometer, thermoelectric thermometer, pyrometer; Kinetic theory of gases: Maxwell's distribution of molecular speeds, mean free path, equipartition of energy, Brownian motion, Van der Waal's equation of state, review of the First Law of thermodynamics and its application, reversible and irreversible processes, Second Law of thermodynamics, Carnot cycle; Efficiency of heat engines, Carnot's Theorem, entropy and disorder, thermodynamic functions, Maxwell relations, Clausius-Clapeyron Equation, Gibbs Phase Rule, Third Law of thermodynamics.

8 Lectures



ICT 1107: Physics

Chapter 2

Heat and Thermodynamics



Chapter 2: Heat & Thermodynamics

Kinetic Theory of Gases



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Kinetic Theory of Matter

According to Kinetic theory of matter, every substance (in the form of solid, liquid or gas) consists of a very large number of very small particles called the molecules. The molecules are the smallest particles of a substance that can exist in free state. The molecules possess the characteristic properties of the parent substance. The molecules are in a state of continuous motion with all possible velocities. The velocity of the molecules increases with rise in temperature.



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Kinetic Theory of Matter

The energy possessed by the molecules can be of two forms, kinetic or potential. When the stop-cock of an evacuated flask is opened, air rushes in quickly to fill in the space. This shows that the molecules possess rapid motion and hence they possess kinetic energy. When a solid expands on heating, the molecules are pulled apart against the forces of intermolecular expansion. The amount of work done in separating the molecules to larger distances manifests itself as the potential energy of the molecules. Thus the amount of heat given to a solid substance increases the energy of the molecules and this increase in energy is partly kinetic and partly potential.

Thus, the kinetic theory of matter is based on the following three points : (i) matter is made up of molecules, (ii) molecules are in rapid motion, and (iii) molecules experience forces of attraction between one another.



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Kinetic Theory of Gases

The continuous collision of the molecules of the gas with the walls of the containing vessel and their reflection from the walls results in the change of momentum of the molecules. According to Newton's second law of motion, the rate of change of momentum per unit area of the wall surface corresponds to the force exerted by the gas per unit area. The force per unit area measures the pressure of the gas.



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Basic Postulates of the Kinetic Theory of Gases

1. The molecules in a gas are small and very far apart. Most of the volume which a gas occupies is empty space.
2. A gas consists of molecules in constant random motion.
3. Gas molecules influence each other only by collision; they exert no other forces on each other.
4. All collisions between gas molecules are perfectly elastic; all kinetic energy is conserved. (no loss in KE).



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5. Molecules can collide with each other and with the walls of the container. Collisions with the walls account for the pressure of the gas.
6. The molecules exert no attractive or repulsive forces on one another except during the process of collision. Between collisions, they move in straight lines.
7. The gas is composed of small indivisible particles called molecules. The properties of the individual molecules are the same as that of the gas as a whole.



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8. The distance between the molecules is large as compared to that of a solid or liquid and hence the forces of inter-molecular attraction are negligible.
9. The size of the molecules is infinitesimally small as compared to the average distance traversed by a molecule between any two consecutive collisions. The distance between any two consecutive collisions is called free path and the average distance is called the mean free path. The mean free path is dependent on the pressure of the gas. If the pressure is high the mean free path is less and if the pressure is low the mean free path is more.
10. The molecules are perfectly hard elastic spheres, and the whole of their energy is kinetic.



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11. The time of impact is negligible in comparison with the taken to traverse the mean free path.



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Sample Question

Q. Applying the kinetic theory of gases, the following expression for the pressure of a gas in terms of the properties of its molecules can be derived:

$$P = \frac{MC^2}{3V}$$
$$P = \frac{1}{3} \rho C^2$$



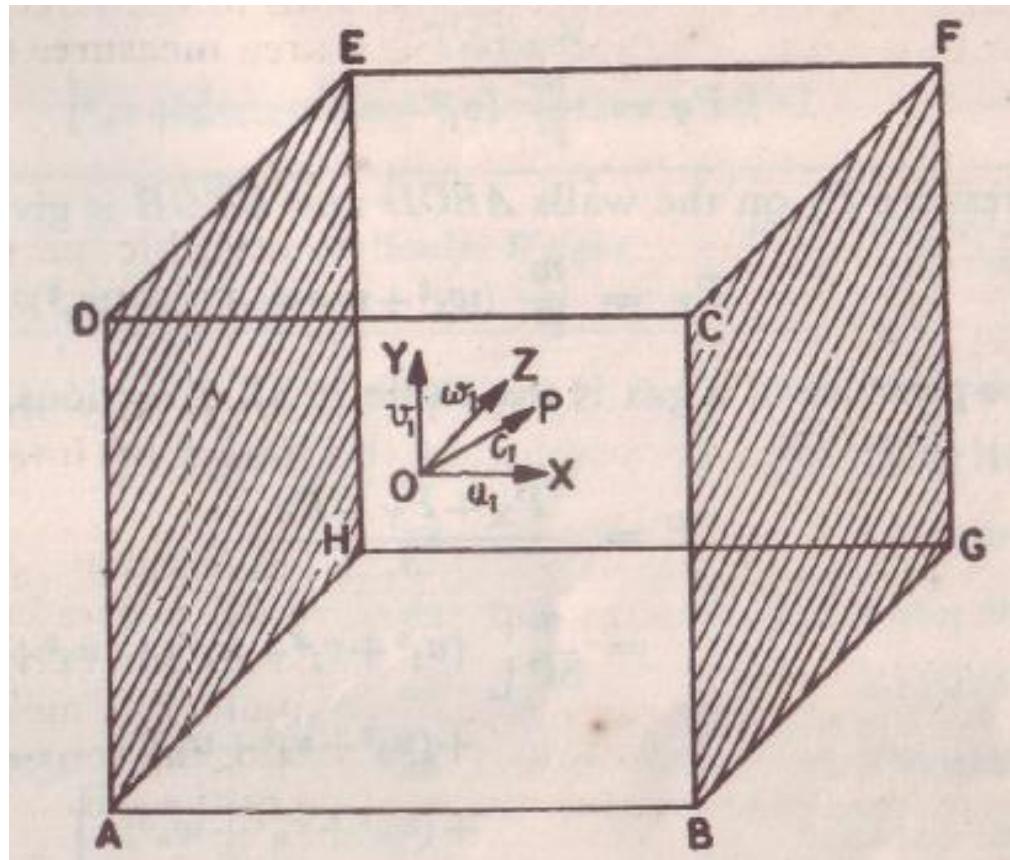
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Answer



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Expression for the Pressure of a Gas





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Consider a cubical vessel $ABCDEFGH$ of side l cm containing the gas. The volume of the vessel and hence that of the gas is l^3 cc. Let n and m represent the very large number of molecules present in the vessel and the mass of each molecule respectively.

Consider a molecule P moving in a random direction with a velocity C_1 . The velocity can be resolved into three perpendicular components u_1 , v_1 and w_1 along the X , Y and Z axes respectively. Therefore,

$$C_1^2 = u_1^2 + v_1^2 + w_1^2$$



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The component of the velocity with which the molecule P will strike the opposite face $BCFG$ is u_1 and the momentum of the molecule is mu_1 . This molecule is reflected back with the same momentum mu_1 in an opposite direction and after traversing a distance l will strike the opposite face $ADEH$.

The change in momentum produced due to the impact is

$$mu_1 - (-mu_1) = 2mu_1$$

As the velocity of the molecule is u_1 , the time interval between two successive impacts on the wall $BCFG$ is

$$\frac{2l}{u_1} \text{ seconds}$$



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∴ No. of impacts per second

$$= \frac{1}{2l} \\ = \frac{u_1}{2l}$$

Change in momentum produced in one second due to the impact of this molecule is

$$2mu_1 \times \frac{u_1}{2l} = \frac{mu_1^2}{l}$$

The force F_x due to the impact of all the n molecules in one second

$$= \frac{m}{l} [u_1^2 + u_2^2 + \dots + u_n^2]$$



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Force per unit area on the wall $BCFG$ or $ADEH$ is equal to the pressure P_X

$$P_X = \frac{m}{l \times l^2} (u_1^2 + u_2^2 + u_3^2 + \dots + u_n^2)$$

Similarly the pressure P_Y on the walls $CDEF$ and $ABGH$ is given by

$$P_Y = \frac{m}{l^3} (v_1^2 + v_2^2 + \dots + v_n^2)$$

and the pressure P_Z on the walls $ABCD$ and $EFGH$ is given by

$$P_Z = \frac{m}{l^3} (w_1^2 + w_2^2 + \dots + w_n^2)$$



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As the pressure of a gas is the same in all directions, the mean pressure P is given by

$$\begin{aligned} P &= \frac{P_x + P_y + P_z}{3} \\ &= \frac{m}{3l^3} \left[(u_1^2 + v_1^2 + w_1^2) + (u_2^2 + v_2^2 + w_2^2) \right. \\ &\quad + (u_3^2 + v_3^2 + w_3^2) + \dots \\ &\quad \left. + (u_n^2 + v_n^2 + w_n^2) \right] \\ &= \frac{m}{3l^3} \left[C_1^2 + C_2^2 + C_3^2 + \dots + C_n^2 \right] \quad \dots(i) \end{aligned}$$

But volume, $V = l^3$. Let C be the root-mean-square velocity of the molecules (R.M.S. velocity).

Then

$$C^2 = \frac{C_1^2 + C_2^2 + C_3^2 + \dots + C_n^2}{n}$$

or

$$nC^2 = C_1^2 + C_2^2 + C_3^2 + \dots + C_n^2$$

Substituting this value in equation (i), we get

$$P = \frac{m \cdot n C^2}{3V} \quad \dots(ii)$$



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But $M = mn$ where M is the mass of the gas of volume V , m is the mass of each molecule and n is the number of molecules in a volume V .

∴

$$P = \frac{MC^2}{3V}$$

or

$$P = \frac{1}{3} \rho C^2 \quad \dots(iii)$$

∴

$$\frac{M}{V} = \rho \text{ the density of the gas.}$$

From equation (iii)

$$C^2 = \frac{3P}{\rho}$$

$$C = \sqrt{\frac{3P}{\rho}}$$

...(iv)



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Self Study

Kinetic Interpretation of Temperature



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Self Study

Derivation of Gas Laws from Kinetic Theory of Gases



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Sample Questions

1. Give the kinetic interpretation of temperature.
2. Applying the kinetic theory of gases derive the ideal gas equations



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Brownian Motion (Robert Brown in 1827)

You might well have studied Brownian Motion before. It is unlikely that you fully appreciated or understood it!!!

One of the first pieces of evidence for how gases move was discovered by Robert Brown a botanist who noticed that pollen grains in water moved about randomly.

Robert Brown noticed that small grains of pollen (clarkia) suspended in water, moved in a haphazard or random fashion.



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Brownian Motion

It was only visible under a microscope (he used fairly low magnification).

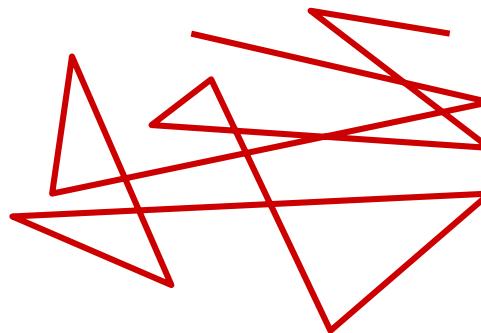
Being a botanist he logically thought that the grains of pollen were alive.

He then tried various other small particles such as carbon (in ink) and realized that the motion was not truly due to life.



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Brownian Motion



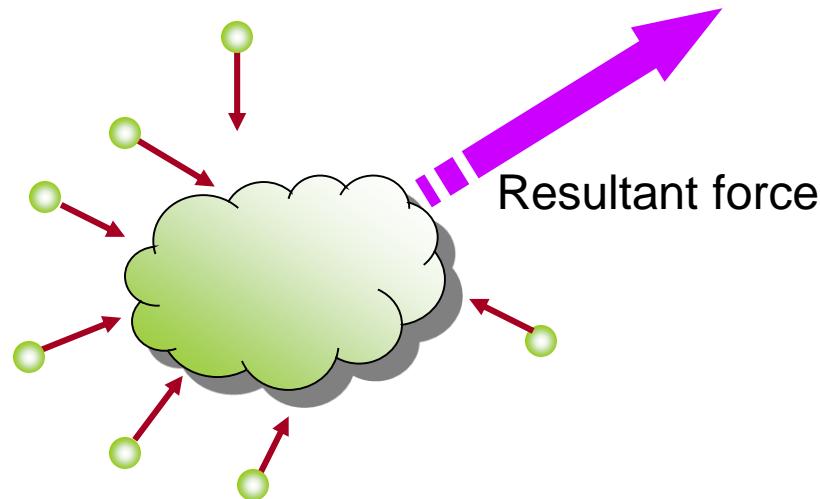
Remember

- There is no appreciable displacement of the particle
- The particle travels more or less in straight lines
- The motion is RANDOM



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Brownian Motion



At any point of time the forces on the particles are uneven causing a net force and therefore acceleration in that direction.

Their light mass means they change direction quickly and easily.



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Brownian Motion

The non-stop, random, perpetual, spontaneous, and haphazard motion of a colloidal solution is called the Brownian motion.



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Essential Features of Brownian Motion

1. The motion of each particle is completely irregular, and random. No two particles are found to execute the same motion.
2. The motion is continuous and takes place for ever.
3. The smaller particles appear to be more agitated than the longer ones.
4. The motion is independent of the nature of the suspended articles. Two particles of the same weight and size move equally fast at the same temperature.



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5. The motion becomes more vigorous when the temperature is increased.
6. The motion is more conspicuous in a liquid of lower viscosity.
7. The motion is not modified due to the shaking of the containing vessel.



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Basis of Brownian Motion

Brownian motion is due to the bombardment of the dispersed particles by molecules of the medium which are themselves in a state of continuous haphazard motion. When the particles are sufficiently large, the forces due to molecular impact is nearly balanced. This is why Brownian motion cannot be observed with particles of large size. But when the size of the particles is very small each particle will be acted upon by a resultant unbalanced force and will consequently move in a haphazard manner in response to the magnitude and direction of this force.



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In other words, the particles move in response to the molecules of the liquid in a completely haphazard manner. The motion of the molecules of a gas is similar to Brownian movement if suspended particles in colloidal solution. Thus, the laws of kinetic theory of gases are applicable to Brownian motion.



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Question

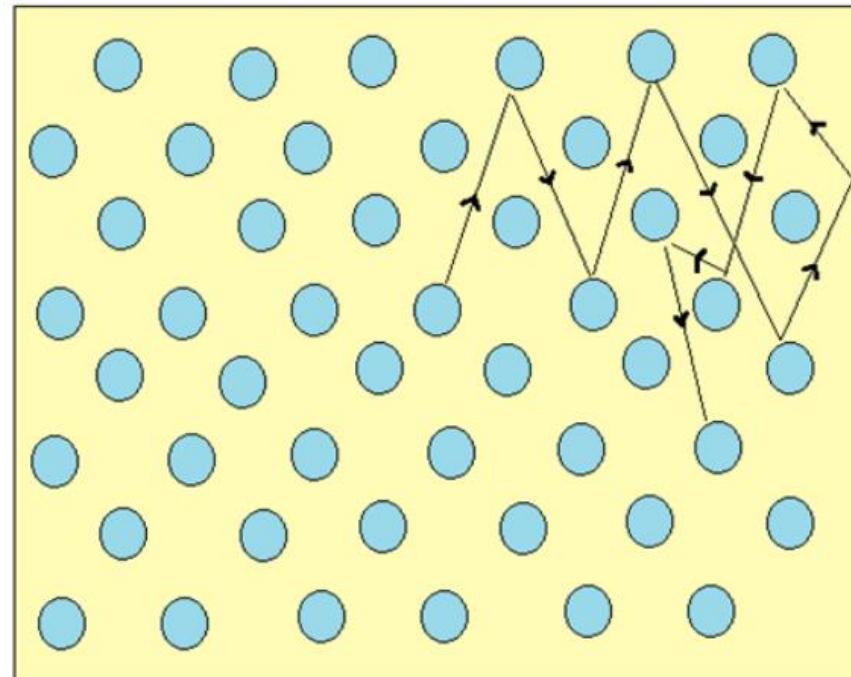
Q. Define Brownian motion with its essential features. Explain the basis of Brownian motion.



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Mean Free Path of Gases

The mean free path λ of a gas molecule is its average path length between collisions.



The random path followed by a gas molecule.



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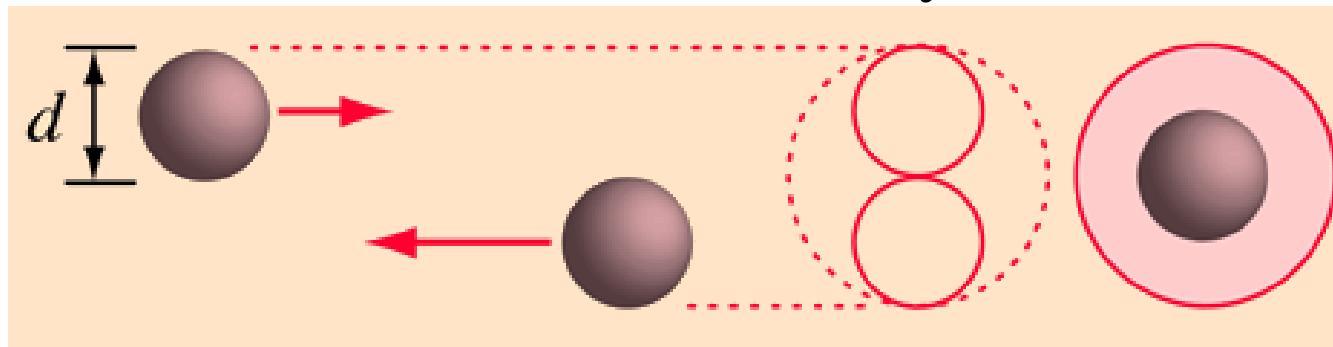
- On the basis of kinetic theory of gases, it is assumed that the molecules of a gas are continuously colliding against each other.
- Molecules move in straight line with constant speeds between two successive collisions.
- Thus path of a single molecule is a series of zig-zag paths of different lengths as shown in figure.

These paths of different lengths are called free paths of the molecule. Mean free path is the average distance traversed by molecule between two successive collisions.



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If the molecules have diameter d , then the effective cross-section for collision can be defined by,

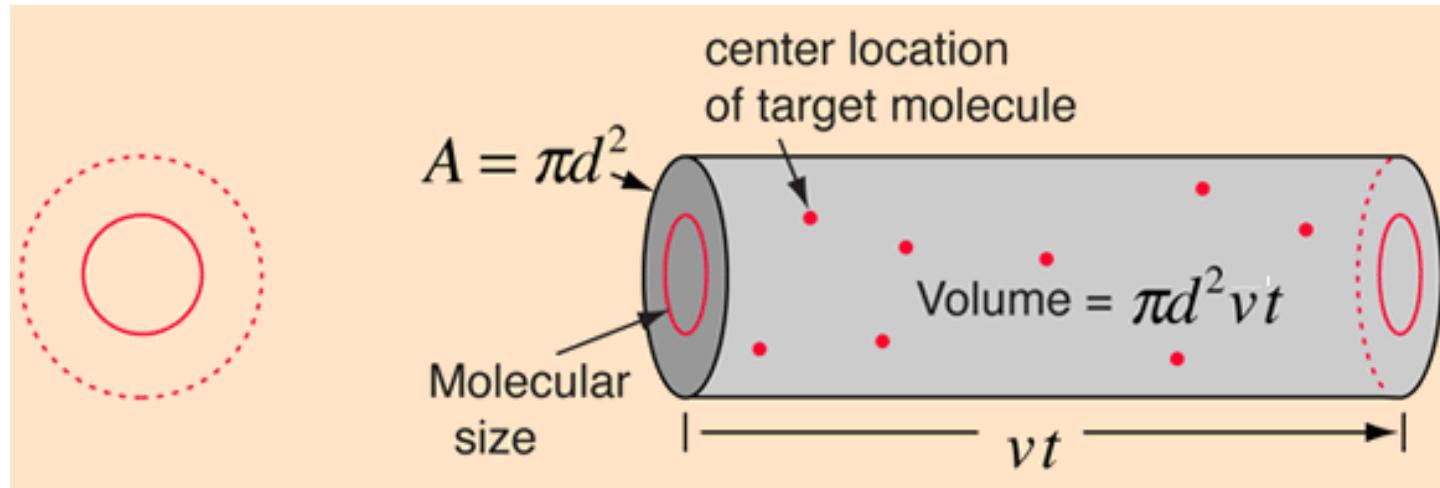


The effective cross-section for collision, $A = \pi d^2$.

Using a circle of diameter $2d$ to represent a molecule's effective collision area while treating the 'target' molecules as point masses. In time t , the circle would sweep out the volume shown and the number of collisions can be estimated from the number of gas molecules available in that volume.



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The mean free path could then be taken as the length of the path divided by the number of collisions. Mathematically,

$$\lambda = \frac{\text{Distance traveled in time } t}{\text{No of collisions suffered in time } t}$$



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The mean free path,

$$\begin{aligned}\lambda &= \frac{vt}{\pi d^2 v t \cdot \frac{N}{V}} \\ &= \frac{1}{\pi d^2 n}\end{aligned}$$

This is the expression for the mean free path of a gas molecule.

This equation was deduced by Clausius.

$$\therefore \lambda \propto \frac{1}{d^2}$$

The mean free path is inversely proportional to the square of the diameter of the molecules.



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Here, we have assumed that all the particles are stationary with respect to the particle we are studying, but in reality all the molecules are moving relative to each other, we have cancelled two velocities in the equation but actually the v in the numerator is the average velocity and v in the denominator is relative velocity hence they both differ from each other with a factor $\sqrt{2}$ therefore the final expression of the mean free path would be,

$$\lambda = \frac{1}{\sqrt{2}\pi d^2 n}$$

This is known as the Maxwell's mean free path based on the law of equipartition of energy.



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Let m be the mass of each molecule.

Then, $m \times n = \rho$

$$\lambda = \frac{m}{\pi d^3 \rho}$$

The mean free path is inversely proportional to the density of the gas.

The expression for the mean path according to Boltzmann is

$$\lambda = \frac{3}{4\pi d^2 n}$$

He assumed that all molecules have the same average speed.



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Law of Equipartition of Energy

A molecule in a gas can move along any of the three co-ordinate axes. It has three degrees of freedom. Degrees of freedom mean the number of independent variables that must be known to describe the state or the position of the body completely. A monoatomic gas molecule has three degrees of freedom. A diatomic gas molecule has three degrees of freedom of translation and two degrees of freedom of rotation. It has in all five degrees of freedom.

According to kinetic theory of gases, the mean kinetic energy of a molecule at a temperature T is given by

$$\frac{1}{2}mC^2 = \frac{3}{2} kT \quad \dots(i)$$

But

$$C^2 = u^2 + v^2 + w^2$$



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As x , y and z are all equivalent, mean square velocities along the three axes are equal

$$\therefore u^2 = v^2 = w^2$$

or

$$\frac{1}{2} m(u^2) = \frac{1}{2} m(v^2) = \frac{1}{2} m(w^2)$$

$$\begin{aligned}\therefore \frac{1}{2} mC^2 &= 3 [\frac{1}{2} m(u^2)] = 3 [\frac{1}{2} m(v^2)] \\ &= 3 [\frac{1}{2} m(w^2)]\end{aligned}$$

$$= \frac{3}{2} kT$$

$$\therefore \frac{1}{2} mu^2 = \frac{1}{2} kT \quad \dots(iii)$$

$$\frac{1}{2} mv^2 = \frac{1}{2} kT \quad \dots(iv)$$

$$\frac{1}{2} mw^2 = \frac{1}{2} kT \quad \dots(iv)$$

Therefore, the average kinetic energy associated with each degree of freedom
= $\frac{1}{2} kT$



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Thus the energy associated with each degree of freedom (whether translatory or rotatory) is $\frac{1}{2} kT$.

This represents the theorem of equipartition of energy.



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(i) *Mono-atomic gas.* A mono-atomic gas molecule has one atom. Each molecule has three degrees of freedom due to translatory motion only.

$$\text{Energy associated with each degree of freedom} = \frac{1}{2} kT$$

Energy associated with three degrees of freedom

$$= \frac{3}{2} kT$$

Consider one gram molecule of a gas.

Energy associated with one gram molecule of a gas

$$= N \times \frac{3}{2} kT$$

$$= \frac{3}{2} (N \times k) T$$



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[But, $N \times k = R$]

$$\therefore U = \frac{3}{2} RT$$

This energy of the gas is due to the energy of its molecules. It is called internal energy U . For an ideal gas, it depends upon temperature only.

$$\therefore C_v = \frac{dU}{dT} = \frac{3}{2} R$$

$\left(\frac{dU}{dT} \right)$ is the increase in internal energy per unit degree rise of temperature)

But

$$C_p - C_v = R$$

$$C_p = C_v + R$$

$$= \frac{3}{2} R + k = \frac{5}{2} R$$



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For a mono-atomic gas

$$\begin{aligned}\gamma &= \frac{C_p}{C_v} \\ &= \frac{\frac{5}{2}R}{\frac{3}{2}R} = 1.67\end{aligned}$$

The value of γ is found to be true experimentally for mono-atomic gases like argon and helium.



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(ii) *Diatom gas.* A diatomic gas molecule has two atoms. Such a molecule has three degrees of freedom of translation and two degrees of freedom of rotation.

- Energy associated with each degree of freedom

$$= \frac{1}{2} kT$$

Energy associated with 5 degrees of freedom = $\frac{5}{2} kT$

Consider one gram molecule of gas.

Energy associated with 1 gram molecule of a diatomic gas

$$= N \times \frac{5}{2} kT = \frac{5}{2} RT$$

$$U = \frac{5}{2} RT$$

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

$$= \frac{5}{2} R$$



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But

$$C_p - C_v = R$$

$$C_p = C_v + R$$

$$= \frac{5}{2}R + R = 7/2R$$

$$\gamma = \frac{C_p}{C_v}$$

$$= \frac{\frac{7}{2}R}{\frac{5}{2}R} = 1.40$$

The value of $\gamma = 1.40$ has been found to be true experimentally for diatomic gases like hydrogen, oxygen, nitrogen etc.



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(iii) *Triatomic gas.* (a) A triatomic gas having 6 degrees of freedom has an energy associated with 1 gram molecule

$$= N \times \frac{6}{2} kT = 3RT$$

$$U = 3RT$$

$$C_V = \frac{dU}{dT} = 3R$$

But

$$C_P - C_V = R$$

$$C_P = C_V + R$$

$$= 3R + R = 4R$$

$$\begin{aligned}\gamma &= \frac{C_P}{C_V} \\ &= \frac{4R}{3R} = 1.33\end{aligned}$$



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(b) A triatomic gas having 7 degrees of freedom has an energy associated with 1 gram molecule = $N \times \frac{7}{2} kT = \frac{7}{2} R$

$$U = \frac{7}{2} RT$$

$$C_V = \frac{dU}{dT} = \frac{7}{2} R$$

But

$$C_P - C_V = R$$

$$C_P = C_V + R$$

$$= \frac{7}{2} R + R = \frac{9}{2} R$$

$$\gamma = \frac{C_P}{C_V}$$

$$= \frac{9/2 R}{7/2 R} = 1.28$$

Thus the value of γ , C_P and C_V can be calculated depending upon the degrees of freedom of a gas molecule.



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Ideal Gas Equation

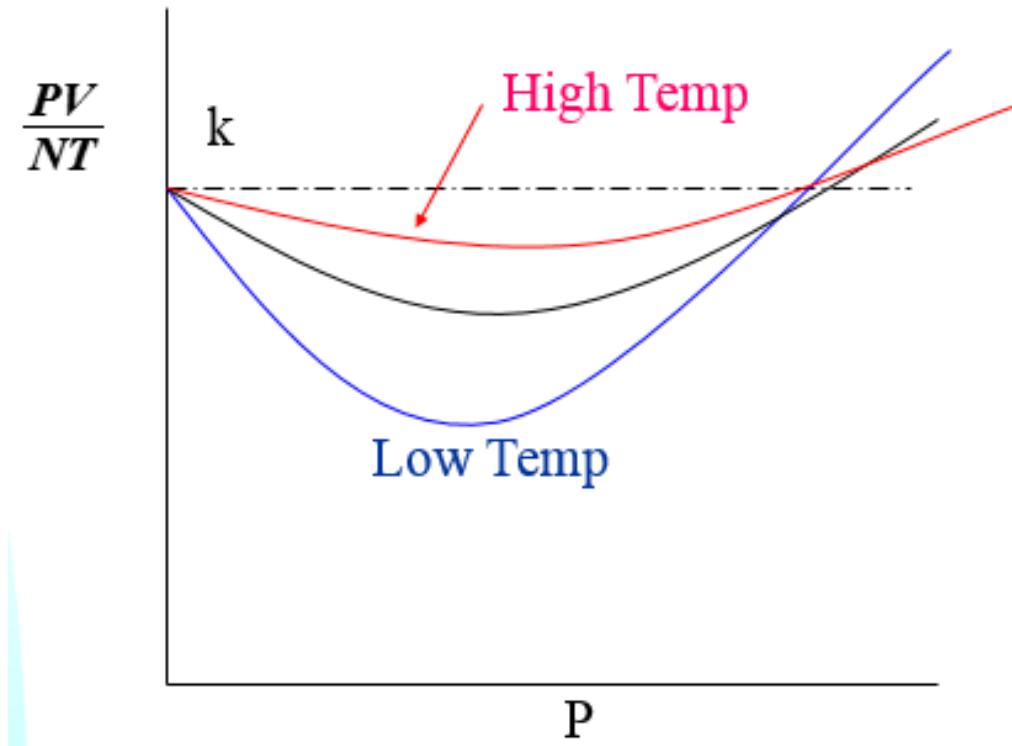
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Van der Waal's Modifications of the Ideal Gas Equation



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If P , V , and T measurements are made on a sample of a real gas, if the results are plotted graphically, it is found that (N =number of molecules),





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As the pressure decreases $\frac{PV}{NT}$ approaches a fixed value. Furthermore this fixed value is found to be the same for all gases. This experimental number is called Boltzmann's constant and the measured value is:

$$k = 1.3806504 \times 10^{-23} \frac{J}{K}$$

Thus, at low pressure, $\frac{PV}{NT} \approx k$ $PV \approx NkT$

By definition, an ideal gas, obeys this equation of state exactly.

$$PV = NkT$$

and a real gas behaves like an ideal gas when its pressure is low.



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The universal gas constant R is defined as,

$$R \equiv N_A k = \left(6.02214179 \times 10^{23} \frac{\text{molecules}}{\text{mole}} \right) \left(1.3806504 \times 10^{-23} \frac{\text{J}}{\text{K}} \right)$$

$$R = 8.314472 \frac{\text{J}}{\text{moleK}} \quad n = \text{number of moles}$$

$$R = N_A k = \left(\frac{N}{n} \right) k, \quad nR = Nk$$

Thus for ideal gas, $PV = nRT$

This is the ideal gas equation.



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Van der Waal's Equation of State



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The ideal gas equation: $PV = nRT$, derived using the KTG assuming that

1. The size of the molecules of the gas is negligible and
2. The forces of intermolecular attraction are absent

However, at high pressure (at higher densities), the size of the molecules of the gas becomes significant and can't be neglected in comparison with the gas volume.



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In the same way, at high pressure, the gas molecules come closer and the forces of intermolecular attraction are appreciable. Thus, corrections are needed to the ideal gas equation.



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The ideal gas equation is modified based on:

- ✓ Attractive forces between gas particles/
Correction for pressure
(*attractive effect*)

- ✓ Non-zero volumes of gas particles/ Correction
for volume
(*repulsive effect*)



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Attractive effect (Correction for Pressure)

A molecule in the interior of a gas experiences forces of attraction in all directions and the resultant cohesive force is zero. A molecule near the walls of the container experiences a resultant force inwards. For this reason, the observed pressure of the gas is less than the actual pressure. The correction for p depends on:

- (i) The number of molecules striking unit area of the walls of the container/sec
- (ii) The number of the molecules present in a given volume.



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Both factors depend on the density of the gas. Hence the correction for pressure,

$$p \propto \rho^2 \propto \frac{1}{V^2}$$
$$\text{or, } p = \frac{a}{V^2}$$

Here, a is a constant and V is the volume of the gas. Thus, the correct pressure of the gas will be

$$P + p = P + \frac{a}{V^2}, \text{ where } P \text{ is the observed pressure.}$$



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Repulsive effect (correction for Volume)

- Gas molecules behave like small, impenetrable spheres
- Actual volume available for gas smaller than volume of container, V
- The molecules have the sphere of influence around them and due to this factor, the correction for volume b is approximately 4 times actual volume of the gas molecules.



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Thus, the corrected volume of gas = $V - b$.

Let the radius of one molecule be r . The volume of the molecule,

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

The volume of the sphere of influence of each molecule,

$$S = \frac{4}{3}\pi(2r)^3 = 8x$$



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Consider a container of volume V and the molecules are allowed to enter one by one.

The volume available for the first molecule = V

The volume available for the 2nd molecule = $V-S$

The volume available for the 3rd molecule = $V-2S$

.....

.....

Volume available for the n th molecule = $V-(n-1)S$



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Thus, the average space available for each molecule

$$= \frac{V + (V - S) + (V - 2S) + \dots + \{V + (n-1)S\}}{n}$$

$$= V - \frac{S}{n} \{1 + 2 + 3 + \dots + (n-1)\}$$

$$= V - \frac{S}{n} \cdot \frac{(n-1)n}{2}$$

$$= V - \frac{nS}{2} + \frac{S}{2}$$

As the number molecule is very high, we can ignore the last term $\frac{S}{2}$.

Hence, the space available for each molecule

$$= V - \frac{nS}{2} = V - \frac{n(8x)}{2} = V - 4(nx) = V - b, \text{ where } b = 4nx, 4 \text{ times the actual volume of the gas molecules}$$



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Combining both derivations, we get the Van der Waal's Equation of State

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

Where a and b are Van der Waal's constants. From Van der Waal's equation of state, we can write

$$P = \frac{RT}{V - b} - \frac{a}{V^2}$$



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$\frac{a}{V^2}$ comes from interactions among gas molecules.

b comes from volume of the gas molecules.

a and b are different for different gases.



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The Van der Waal's constant a

- ✓ Gives a rough indication of magnitude of intermolecular attraction. Usually, the stronger the attractive forces, the higher is the value of a



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The Van der Waal's constant b

Gives a rough indication of the size of gas molecules. Usually, the bigger the gas molecules, the higher is the value of b



A Brief Summary

- ✓ Ideal gas equation combines Avogadro's, Boyle's and Charles' Laws.

- ✓ Assumptions of ideal gas equation fail for real gases, causing deviations.

- ✓ Van der Waals gas equation accounts for attractive and repulsive effects ignored by ideal gas equation.



A Brief Summary

- ✓ Constants a and b represent the properties of a real gas
- ✓ A gas with higher a value usually has stronger IMF
- ✓ A gas with higher b value is usually bigger
- ✓ A gas cannot be condensed into liquid at temperatures higher than its critical temperature



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Van der Waal's Constants



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Self Study

Calculate the Van der Waal's
Constants (a , and b)



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Critical Constants



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Self Study

Calculate the Critical Constants

P_c , V_c and T_c



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Laws of Thermodynamics



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0th Law of Thermodynamics



0th Law of Thermodynamics

When two systems are put in contact with each other, there will be a net exchange of energy between them unless or until they are in thermal equilibrium, that is, they are at the same temperature

“If A and C are each in thermal equilibrium with B , A is also in thermal equilibrium with C .”

Single temperature and pressure can be attributed to the whole system



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1st Law of Thermodynamics



Chapter 2: Heat & Thermodynamics

✓ Energy is conserved

- **FIRST LAW OF THERMODYNAMICS**

- Examples:

- Engines (Heat -> Mechanical Energy)
 - Friction (Mechanical Energy -> Heat)



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Adding heat Q can:

- ✓ Change temperature
 - ✓ Change state of matter
 - ✓ Can also change ΔU by doing work on the gas
- Change of
Internal Energy ΔU



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The internal energy $\Delta U (=Q + W)$ of a gas can be changed, by

- Adding heat to gas: Q
- Doing work on gas: $W = -P\Delta V$

Note:

(Work done by the gas) = - (Work done on the gas)

$$W_{\text{by the gas}} = +P\Delta V$$

$$Q = \Delta U + W_{\text{by the gas}}$$

Add heat => Increase Int. Energy + Gas does work⁸¹



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What is INTERNAL ENERGY?

The internal energy of the system is the sum of the kinetic and potential energies of the atoms and molecules making up the system.



Another Definition of 1st Law of Thermodynamics

The increase in the internal energy of a system is equal to the amount of heat added to the system, plus the amount of work done **on** the system.

The increase in the internal energy of a system is equal to the amount of heat added to the system, minus the amount of work done **by** the system.

$$\Delta U = Q - W$$

Q = heat add to system

W = work done by system



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The 1st Law of thermodynamics is an empirical statement regarding the conservation of energy.



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Gas Equation in an Adiabatic Process



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Consider 1 gram of the working substance (ideal gas) perfectly insulated from the surroundings. Let the external work done by the gas be δW .

Applying the first law of thermodynamics

$$\delta H = dU + \delta W$$

But

$$\delta H = 0$$

and

$$\delta W = P.dV$$

where P is the pressure of the gas and dV is the change in volume.

$$\therefore 0 = dU + \frac{P.dV}{J} \quad \dots(i)$$

As the external work is done by the gas at the cost of its internal energy, there is fall in temperature by dT .

$$dU = 1 \times C_v \times dT$$

$$C_v.dT + \frac{P.dV}{J} = 0 \quad \dots(ii)$$



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For an ideal gas

$$PV = rT \quad \dots(iii)$$

Differentiating,

$$P.dV + V.dP = r.dT$$

Substituting the value of dT in equation (ii),

$$C_v \left[\frac{P.dV + V.dP}{r} \right] + \frac{P.dV}{J} = 0$$

$$C_v[P.dV + V.dP] + r \cdot \frac{P.dV}{J} = 0$$

But,

$$\frac{r}{J} = C_p - C_v$$



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$$\therefore C_v \cdot P \cdot dV + C_v \cdot V \cdot dP + C_p \cdot P dV - C_v \cdot P dV = 0$$
$$C_p \cdot P \cdot dV + C_v \cdot V \cdot dP = 0$$

Dividing by $C_v \cdot PV$,

$$\frac{C_p}{C_v} \cdot \frac{dV}{V} + \frac{dP}{P} = 0$$

But

$$\frac{C_p}{C_v} = \gamma$$

$$\therefore \frac{dP}{P} + \gamma \frac{dV}{V} = 0$$

Integrating, $\log P + \gamma \log V = \text{const.}$

$$\log PV^\gamma = \text{const.}$$

$$PV^\gamma = \text{const.}$$

...(iv)

~~This is the equation connecting pressure and volume during an adiabatic process.~~



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Taking

$$PV = rT$$

or

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

$$\left(\frac{rT}{V}\right) \cdot V^{\gamma} = \text{const.}$$

But r is const.

$$rTV^{\gamma-1} = \text{const.}$$

∴

$$TV^{\gamma-1} = \text{const.}$$

... (v)

Also

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

$$P\left[\frac{rT}{P}\right]^{\gamma} = \text{const.}$$

or

$$\frac{r^{\gamma} T^{\gamma}}{P^{\gamma-1}} = \text{const.}$$

or

$$\frac{P^{\gamma-1}}{T^{\gamma}} = \text{const.}$$

... (vi) 89



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Thus, during an adiabatic process

$$(i) \quad PV^\gamma = \text{const.}$$

$$(ii) \quad TV^{\gamma-1} = \text{const. and}$$

$$(iii) \quad \frac{P^{\gamma-1}}{T^\gamma} = \text{const.}$$



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Home Work

Solve a few problems related to the theory we have studied in this chapter



Thank You for Listening



Physics is hopefully simple but Physicists are not