CBSE Class 11 Physics Notes Chapter 13: CBSE Class 11 Physics Notes Chapter 13 is all about how gases behave. In this chapter you will learn about the idea that gases are made up of tiny particles, like atoms or molecules, that are always moving.

These notes explain things like how gas particles move and collide with each other and the walls of their container.

They also cover important topics such as different gas laws, which describe how gases behave under different conditions. These notes are helpful for students to understand the basic principles of how gases work in a simple and easy-to-understand way.

CBSE Class 11 Physics Notes Chapter 13 Kinetic Theory of Gases PDF

You can access the CBSE Class 11 Physics Notes Chapter 13 on the Kinetic Theory of Gases in PDF format through the link provided below. These notes cover essential topics such as the behavior of gas particles, gas laws, and the interpretation of temperature and pressure.

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CBSE Class 11 Physics Notes Chapter 13 Kinetic Theory of Gases

Kinetic Theory of Gas

The Kinetic Theory of Gas is a scientific theory that explains the behavior of gases in terms of the motion of their constituent particles, such as atoms or molecules. According to this theory, gases consist of a large number of tiny particles that are in constant, random motion. These particles move freely in all directions and collide with each other and with the walls of their container. The Kinetic Theory of Gas helps us understand various properties of gases, such as pressure, temperature, volume, and the relationship between them. It also provides a basis for explaining gas laws and the ideal gas behavior.

Let's consider a simple example to illustrate the Kinetic Theory of Gases:

Imagine a balloon filled with air. According to the Kinetic Theory of Gases, the air inside the balloon consists of countless tiny particles, such as nitrogen and oxygen molecules, moving rapidly in all directions.

When the balloon is squeezed or compressed, the volume of the gas decreases. This causes the gas particles to collide more frequently with each other and with the walls of the balloon. As a result, the pressure inside the balloon increases.

Conversely, if the balloon is allowed to expand, the volume of the gas increases. This leads to fewer collisions between gas particles and a decrease in pressure.

This example demonstrates how the Kinetic Theory of Gases helps us understand the relationship between pressure, volume, and the motion of gas particles. It explains why changes in volume or pressure of a gas affect its behavior, as well as how gas particles move and interact with their surroundings.

Assumptions of the Kinetic Theory of Gases

Here are the key assumptions of the Kinetic Theory of Gases:

1. All gas molecules constantly move in random directions.
2. The size of molecules is much smaller than the separation between them.
3. Gas molecules do not exert any force on the walls of the container during collisions.
4. Collisions between molecules and walls, as well as between molecules themselves, are elastic in nature.
5. The time interval of collision between two molecules and between a molecule and the wall is very small.
6. Newton's laws of motion can be observed in all the molecules within a gas sample.
7. Over time, a gas sample reaches a steady state, where the distribution and density of molecules do not depend on position, distance, or time.

Key Concepts

- Ideal Gas Law: Describes the relationship between pressure, volume, temperature, and the number of moles of a gas.
- Maxwell-Boltzmann Distribution: Describes the distribution of speeds among gas molecules in a sample.
- Brownian Motion: The random motion of particles suspended in a fluid, caused by collisions with gas molecules.
- Real Gas Behavior: Deviations from ideal gas behavior under conditions of high pressure or low temperature.

Solids, Liquids, Gases in Terms of Molecular Structure

Solids, liquids, and gases differ in their molecular structure, which influences their physical properties and behavior. Here's a comparison of their molecular structures:

Solids:

- In solids, molecules are tightly packed together in a regular, ordered arrangement.
- The intermolecular forces between molecules in solids are strong, leading to a fixed shape and volume.

- The molecules in solids have vibrational motion around their equilibrium positions but cannot move past each other easily.
- Examples of solids include ice, diamond, and metal.

Liquids:

- In liquids, molecules are still closely packed but are less ordered compared to solids.
- The intermolecular forces are weaker than in solids, allowing molecules to move past each other while maintaining contact.
- Liquids have a definite volume but take the shape of their container due to the ability of molecules to flow and slide past each other.
- The molecular motion in liquids includes both vibrational motion and translational motion.
- Examples of liquids include water, ethanol, and oil.

Gases:

- In gases, molecules are much farther apart compared to solids and liquids, with little to no order in their arrangement.
- The intermolecular forces are very weak or negligible, allowing molecules to move freely and independently of each other.
- Gases have neither a fixed shape nor a fixed volume and expand to fill the entire volume of their container.
- The molecular motion in gases is primarily translational, with molecules moving rapidly in all directions.
- Examples of gases include oxygen, nitrogen, and helium.

Justification of Avogadro’s Hypothesis from the Equation of Gas

Avogadro's hypothesis states that equal volumes of gases, under the same conditions of temperature and pressure, contain the same number of molecules. This hypothesis can be justified using the ideal gas equation, which relates the pressure, volume, temperature, and the number of moles of gas. The ideal gas equation is given as:

$PV=nRT$PV=nRT

Where:

- PP is the pressure of the gas,
- VV is the volume of the gas,
- nn is the number of moles of gas,
- RR is the ideal gas constant, and
- TT is the temperature of the gas in Kelvin.

To justify Avogadro's hypothesis, let's consider two gases, AA and BB, occupying volumes VAVA and VBVB, respectively, under the same conditions of temperature and pressure. According to the ideal gas equation:

$P \times VA = nA \times R \times T$P×VA=nA×R×T

$P \times VB = nB \times R \times T$P×VB=nB×R×T

Since the temperature, pressure, and gas constant are the same for both gases, we can write:

$VAnA = VBnB$nAVA=nBVB

This equation states that the ratio of the volume to the number of moles is the same for both gases. If we consider equal volumes of gases, then $VA = VB$VA=VB. Therefore, the ratio $nAVA$VAnA must equal $nBVB$VBnB, implying that the number of moles of gas nAnA is equal to the number of moles of gas nBnB.

In simpler terms, if two gases occupy the same volume under the same conditions, they must contain the same number of molecules. This justifies Avogadro's hypothesis and provides a theoretical basis for understanding the behavior of gases.

Perfect Gas Equation

The perfect gas equation, $PV = \mu RT$PV=μRT, describes the behavior of gases under specific conditions, where PP is pressure, VV is volume, TT is absolute temperature, μμ is the number of moles, and RR is the universal gas constant.

The universal gas constant, RR, is defined as $kB \cdot NA$kB·NA, where kBkB is the Boltzmann constant and NANA is Avogadro's number.

Another form of the perfect gas equation is $PV = NkBT$PV=NkBT, where NN is the number of molecules and kBkB is the Boltzmann constant.

Simplifying further, we get $P = NvkBT$P=NvkBT, where $n = NV$n=NV, and vv is the volume per molecule.

Alternatively, $\mu = MM\circ$μ=M○M, where MM is the mass of the sample and $M\circ$M○ is the molar mass of the sample.

So, the perfect gas equation can also be expressed as $PV = MM\circ RT$PV=M○MRT.

This can further simplify to $P = MVRTM\circ$P=RTM○MV.

And finally, $P = \varrho RT$P=ρRT, where ϱρ is the mass density of the gas, given by $\varrho = MV$ρ=VM.

=nmvx2

Expression of Mean Free Path

Consider each gas molecule to be a sphere with a diameter (dd). Let's denote the average speed of each molecule as vv.

Assume that a molecule will collide with any other molecule within a certain distance equal to its diameter (dd). Thus, any molecule that comes within this range will collide with it.

The volume in which a molecule suffers a collision can be expressed as: $v\Delta t\pi d2$vΔtπd2.

Now, let's assume the number of molecules per unit volume is represented by nn. Therefore, the total number of collisions will be: $v\Delta t\pi d2n$vΔtπd2n.

The rate of collision is given by: $v\Delta t\pi d2n\Delta t=v\pi d2n$ΔtvΔtπd2n=vπd2n.

The supposed time between collisions (ττ) will be: $\tau=1v\pi d2n$τ=vπd2n1.

And the average distance between collisions (τvτv) will be: $\tau v=1\pi d2n$τv=πd2n1.

This value is modified by adding a factor to it, resulting in the mean free path, represented as: $l=12-1\pi d2n$l=21−πd2n1.

Kinetic Interpretation of Temperature

The kinetic interpretation of temperature is based on the kinetic theory of gases, which states that temperature is a measure of the average kinetic energy of the particles in a substance. In simpler terms, it explains how the motion of particles relates to the concept of temperature.

According to this theory:

Temperature reflects the average kinetic energy: The higher the temperature of a substance, the greater the average kinetic energy of its particles. This is because temperature is directly proportional to the average kinetic energy of the particles in the substance.

Temperature and particle speed: As temperature increases, the particles in a substance move faster on average. Conversely, at lower temperatures, particle motion slows down. This relationship between temperature and particle speed helps us understand how changes in temperature affect the behavior of substances.

Absolute zero: According to the kinetic theory of gases, absolute zero is the temperature at which particle motion ceases entirely. At this temperature, particles have minimal kinetic energy, and their motion stops. Absolute zero serves as the zero point on the Kelvin temperature scale.

Temperature and pressure: The kinetic theory also explains how changes in temperature can affect the pressure exerted by a gas. When the temperature of a gas increases, the average kinetic energy of its particles increases, leading to more frequent and forceful collisions with the walls of the container. This results in an increase in pressure.

Kinetic Theory: Consistent With Ideal Gas Equation and Gas Laws

It is consistent with the ideal gas equation:

For the kinetic gas equation:

EN=32kBTEN=32kBT

E=32NkBTE=32NkBT

For an ideal gas, its internal energy is directly proportional to the temperature. This depicts that internal energy of an ideal gas is only dependent on its temperature, not on pressure or volume.

When Kinetic theory is consistent with Dalton’s law of partial pressure:

The equation for kinetic theory: P=13nmv2P=13nmv2

If a mixture of gases is present in the vessel, then:

P=13[n1m1v12+n2m2v22+...]P=13[n1m1v12+n2m2v22+...]

The average kinetic energy of molecules of different gases at equilibrium will be equal

12m1v12=12m2v22=32kBT12m1v12=12m2v22=32kBT

Then the total pressure will be:

P=13[32n1kBT+32n2kBT]P=13[32n1kBT+32n2kBT]

P=kBT[n1+n2+....]P=kBT[n1+n2+....]

P=RNA[n1+n2+....]P=RNA[n1+n2+....]

P=[μ1+μ2+....]RTP=[μ1+μ2+....]RT

P=P1+P2+.....P=P1+P2+.....

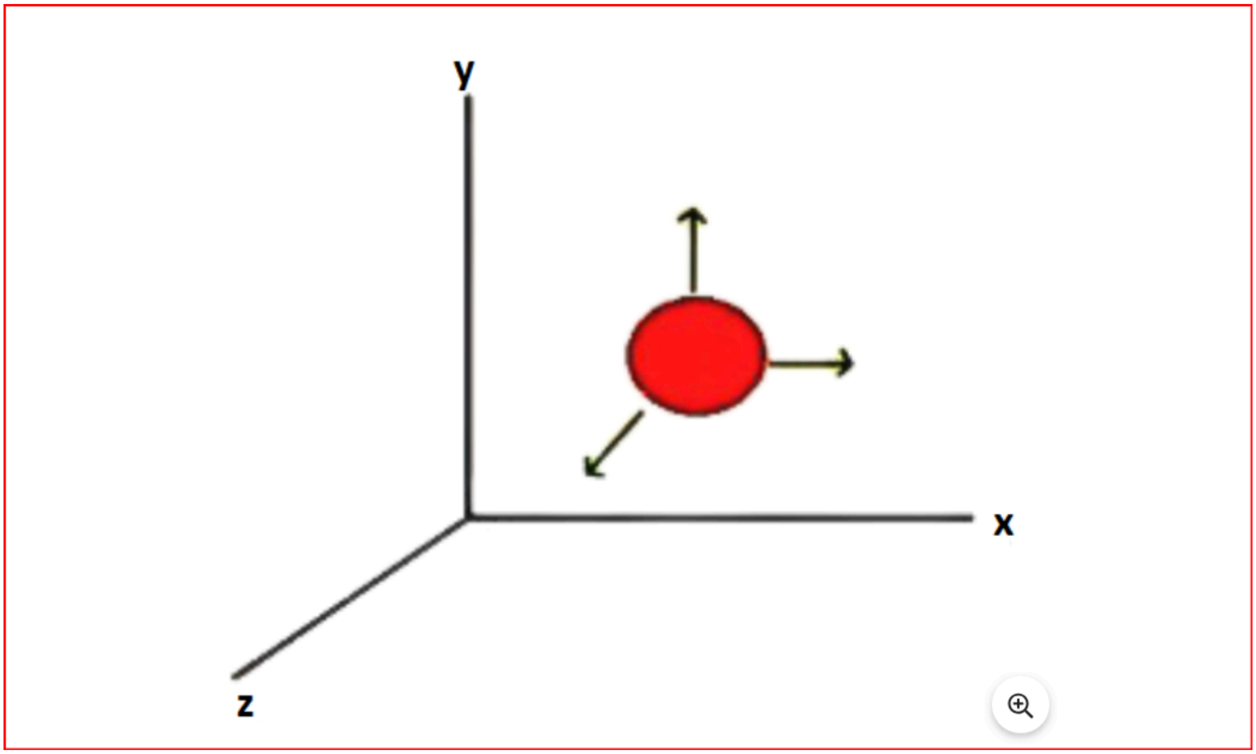
This is known as Dalton’s law of partial pressure.

Law of Equipartition of Energy: Degrees of Freedom

Degrees of freedom represent the independent ways a body or system can move, rotate, or vibrate. They are crucial in understanding the energy distribution of a system, as outlined in the Law of Equipartition of Energy. Here is the different types of degrees of freedom:

Translational Degree of Freedom: This refers to the movement of the entire body from one position to another. For example, in a gas molecule, translational movement occurs when the entire molecule shifts its position in space.

The number of translational degrees of freedom depends on the dimensionality of the movement. In three-dimensional space, a molecule has three translational degrees of freedom (x, y, and z axes). In two-dimensional space, it has two degrees of freedom, and in a straight line, it has only one. In monoatomic gas molecules, which consist of a single atom, only translational degrees of freedom exist.



Rotational Degree of Freedom: This refers to the independent rotations that specify a body's or system's orientation. In diatomic gas molecules, rotational degrees of freedom exist in addition to translational degrees.

For example, in a diatomic molecule like oxygen (O2), the rotation of one oxygen atom relative to the other adds rotational degrees of freedom. There can be rotations along different axes, contributing to the total rotational degrees of freedom.

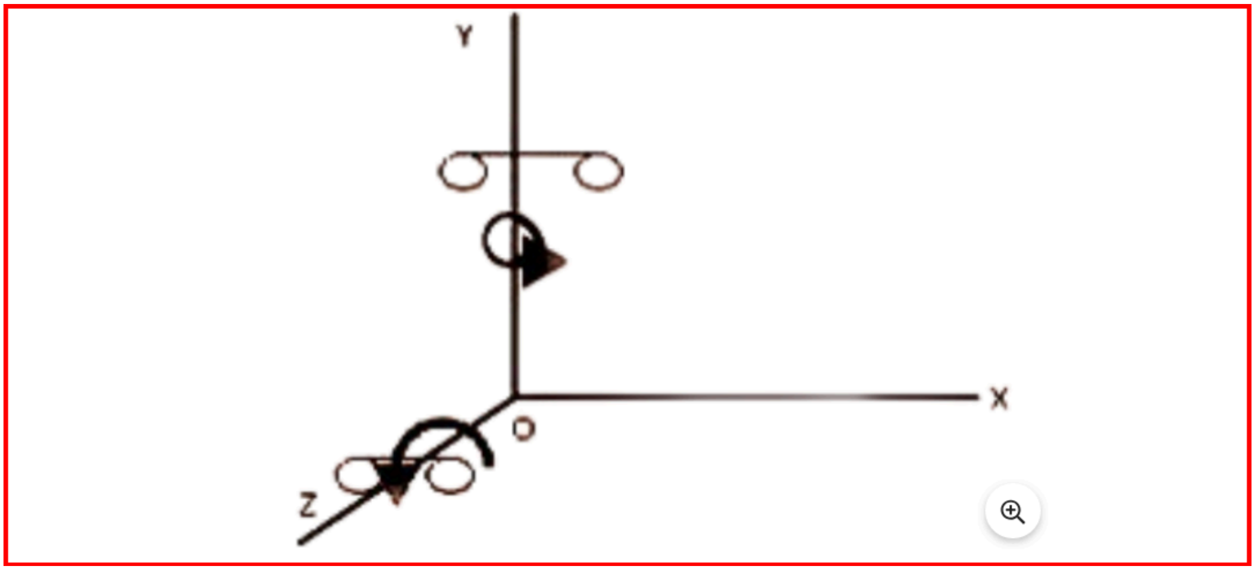
Vibrational Degree of Freedom: In certain molecules, atoms exhibit a specific type of vibration characterized by oscillations along the inter-atomic axis, akin to a one-dimensional oscillator.

This vibration phenomenon is observable in certain molecules, such as in the case of carbon monoxide (CO), where atoms oscillate along the inter-atomic axis like a one-dimensional oscillator.

Consider a scenario where two atoms vibrate in the same direction along the inter-atomic axis.

The expression for total vibrational energy includes terms representing the square of vibrational motion variables.

The total vibrational energy (Ev) is given by the equation: Ev=12m(dydt)2+12ky2, where the first term represents kinetic energy (12m(dydt)2) and the second term represents potential energy (12ky2).



Two distinct terms contribute to the vibrational degree of freedom:

1. The first term represents two axes perpendicular to the line connecting the two particles that rotate, as depicted in the diagram.
2. The second term illustrates the vibrational motion along the line joining the two atoms, as shown in the diagram.

Law of Equipartition of Energy

The Law of Equipartition of Energy states that in equilibrium, the total energy is evenly distributed among all possible energy modes, with each mode having an average energy equal to 1/2 kBT.

- For translational degrees of freedom, each contributes 1/2 kBT to the total energy.
- For rotational degrees of freedom, each contributes 1/2 kBT.
- For vibrational degrees of freedom, each contributes 1 kBT.

Specific Heat Capacity for Monoatomic Gases

For monoatomic gases, there is only one degree of freedom, which is translation.

However, they can have up to three translational degrees of freedom.

Each degree of freedom contributes 1/2 kBT to the total energy, so for three degrees of freedom, it will be 3/2 kBT.

The total internal energy of 1 mole of gas (U) is calculated using the law of energy equipartition, giving us U = 3/2 kBTNA = 3/2 RT.

At constant volume, the specific heat capacity (CV) will be: CV = dU/dT = 3/2R.

For an ideal gas, CP – CV = R. Therefore, CP = 5/2R.

The ratio of specific heats (γ) will then be: γ = CP/CV = 5/3.

Specific Heat of Diatomic Gases (Rigid)

- A stiff diatomic gas has translational and rotational degrees of freedom but no vibrational degrees of freedom.
- They are oscillators that are stiff.
- There are three translational degrees of freedom and two rotational degrees of freedom in a stiff diatomic molecule. There are a total of 5 degrees of freedom.
- Each degree of freedom will contribute 12kB12kB according to the law of energy equi-partition.
- As a result, the fifth degree of freedom will add 52kB52kB to the total.
- The total internal energy for one mole of gas will be: U=52kBTNA=52RTU=52kBTNA=52RT
- At constant volume the specific heat capacity will be: CV=dUdT=52RCV=dUdT=52R
- For a rigid diatomic, the specific heat capacity at constant pressure will be given as: CP=72RCP=72R
- The ratios of specific heat will be: γ=CPCV=75

Specific Heat of Diatomic Gases (non – rigid)

- Translational, rotational, and vibrational degrees of freedom exist in a no-rigid diatomic gas.
- There will be three degrees of freedom in translation, two degrees of freedom in rotation, and one degree of freedom in vibration.
- The total contribution by translational is 12kBT12kBT , by rotational is two times of translational and by vibrational is kBkB
- Total internal energy for one mole is equal to: 52kBT+kBT=72kBT=72RT52kBT+kBT=72kBT=72RT
- Cv=dUdT=72RCv=dUdT=72R
- CP=CV+R=92RCP=CV+R=92R
- γ=CPCV=97γ=CPCV=97

Specific Heat for Polyatomic Molecules

- Polyatomic gases will have three degrees of freedom in translation, three degrees of freedom in rotation, and a ‘f' number of vibrational modes.
- For one mole of gas the total internal energy will be: $(32kBT+32kBT+fkBT)\times NA=(3+f)RT(32kBT+32kBT+fkBT)\times NA=(3+f)RT$
- $Cv=dUdT=(3+f)RCv=dUdT=(3+f)R$
- $CP=CV+R=(4+f)RCP=CV+R=(4+f)R$
- $\gamma=CPCV=(4+f)(3+f)\gamma=CPCV=(4+f)(3+f)$

Specific Heat Capacity for Solids

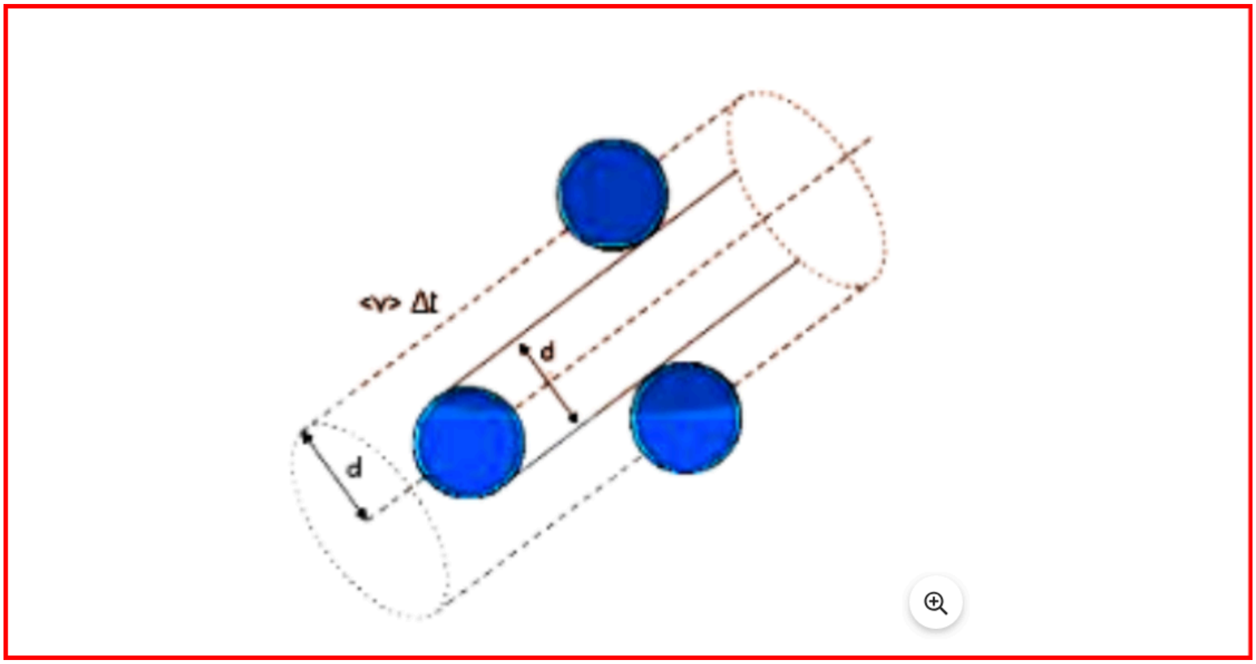
- Consider the number of atoms in a solid, which is N. Each atom has the ability to fluctuate around its mean position.
- The degree of freedom for vibrational motion will be: $kBTkBT$
- The average energy for one – dimensional is $kBTkBT$and for three – dimensional will be $3kBT3kBT$.
- The total internal energy for one mole of solid will be: $3kBTNA=3RT3kBTNA=3RT$
- The change in volume is very less in solids at constant pressure, $\Delta Q=\Delta U+P\Delta V\Delta Q=\Delta U+P\Delta V$

  . So, $\Delta V=0\Delta V=0$ therefore, $\Delta Q=\Delta U\Delta Q=\Delta U$
- $Cv=(dUdT)VCv=(dUdT)V$
- $CP=(dQdT)VCP=(dQdT)V$ as, $\Delta Q=\Delta U\Delta Q=\Delta U$, therefore, $Cv=dUdT=3RCv=dUdT=3R$
- Therefore, $CP=CV=3RCP=CV=3R$

Specific Heat Capacity of Water

- If water is considered as solid, then there will be ‘N’ number of atoms.
- So, the average energy of each atom is: $3kBT3kBT$
- Therefore, there are three molecules in a water molecule $(H2O)(H2O)$.
- So, the total internal energy will be: $U=3kBT\times 3\times NA=9RTU=3kBT\times 3\times NA=9RT$
- $CP=CV=9RCP=CV=9R$

Expression of Mean Free Path



To express the mean free path, let's consider each gas molecule as a sphere with a diameter of dd meters, and the average speed of each molecule as vv.

We assume that a molecule collides with any other molecule within a distance equal to its diameter (dd). Any molecule within this range will collide with it.

The volume in which a molecule suffers collision is given by $v\Delta t\pi d2$vΔtπd2.

Let nn be the number of molecules per unit volume.

Therefore, the total number of collisions will be $v\Delta t\pi d2n$vΔtπd2n.

The collision rate will be $v\Delta t\pi d2n\Delta t=v\pi d2n$ΔtvΔtπd2n=vπd2n.

The average time between collisions will be $\tau=1v\pi d2n$τ=vπd2n1.

And the average distance between collisions will be $\tau v=1\pi d2n$τv=πd2n1.

This value is modified by adding a factor to it.

So, the mean free path will be equal to $l=12-1\pi d2n$l=21−πd2n1.

Conclusion

The Kinetic Theory of Gases provides a powerful framework for understanding the behavior of gases and their interactions. By applying its principles, physicists and engineers can make predictions, design experiments, and develop technologies that rely on the properties of gases. Understanding this theory is essential for mastering the fundamentals of physics and its applications in various fields.

Benefits of CBSE Class 11 Physics Notes Chapter 13 Kinetic Theory of Gases

Conceptual Understanding: The chapter helps students develop a deep conceptual understanding of the behavior of gases at the molecular level, including concepts such as pressure, temperature, and volume.

Foundation for Higher Studies: It lays a strong foundation for students who wish to pursue higher studies in physics or related fields like chemistry, engineering, or atmospheric sciences.

Practical Applications: Understanding the Kinetic Theory of Gases is crucial for comprehending real-world phenomena such as the operation of engines, refrigeration systems, and atmospheric processes.

Problem-Solving Skills: Solving numerical problems related to the Kinetic Theory of Gases enhances students problem-solving skills and analytical thinking abilities.