Al-Mustaqbal University
College of Science
Forensic Evidence Department
Second Stage

جامـــــعـة المـــــستـقـبـل
AL MUSTAQBAL UNIVERSITY

كلية العلـــوم
قــســـــــــم الادلة الجنائية

Lecture (6)

Molecular Speeds and Kinetic Energy

Second stage

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Molecular Speeds and Kinetic Energy

The speed of molecules in thermodynamics relates to the motion of the particles that make up the system, where temperature is a measure of the average kinetic energy of these molecules. As the temperature increases, the speed of the molecules and their average kinetic energy also increase, leading to a broader and faster distribution of velocities, with the peak of the distribution shifting toward higher speeds. These velocities and their motion affect the pressure by increasing the rate and intensity of collisions with the walls of the container.

The relationship between temperature and molecular speed

- Increase in temperature: Leads to an increase in the average kinetic energy of the molecules, and consequently, their speed increases.
- Velocity distribution: As the temperature rises, the velocity distribution shifts toward higher speeds, meaning that fewer molecules will have low speeds while a larger number will move at higher speeds. The overall shape of the velocity distribution changes accordingly, as described in... The Chemistry textbook for the University of Calgary.

• Kinetic energy: It is the energy possessed by molecules as a result of their motion and is directly proportional to the square of their velocity.

Ek=1/2mv2$_{cap}$ E sub k equals one-half m v squared

$$Ek=1/2mv^2$$

• Maxwell–Boltzmann distribution of speeds: It describes the distribution of molecular speeds in a gas at a given temperature, where all molecules have different speeds, but the average speed and some most probable speeds can be determined.

- The previous discussion showed that the Maxwell–Boltzmann theory qualitatively explains the behaviors described by various gas laws. The assumptions of this theory can also be applied quantitatively to derive these individual laws.
- To do this, we must first study the speeds of gas molecules, their kinetic energies, and the temperature of a gas sample.
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- In a gas sample, the speeds of individual molecules vary greatly; however, due to the enormous number of molecules and collisions involved, the distribution of molecular speeds and the average speed remain constant. This distribution of molecular speeds is known as the Maxwell–Boltzmann distribution, which represents the relative numbers of molecules in a large gaseous sample that possess a specific speed.

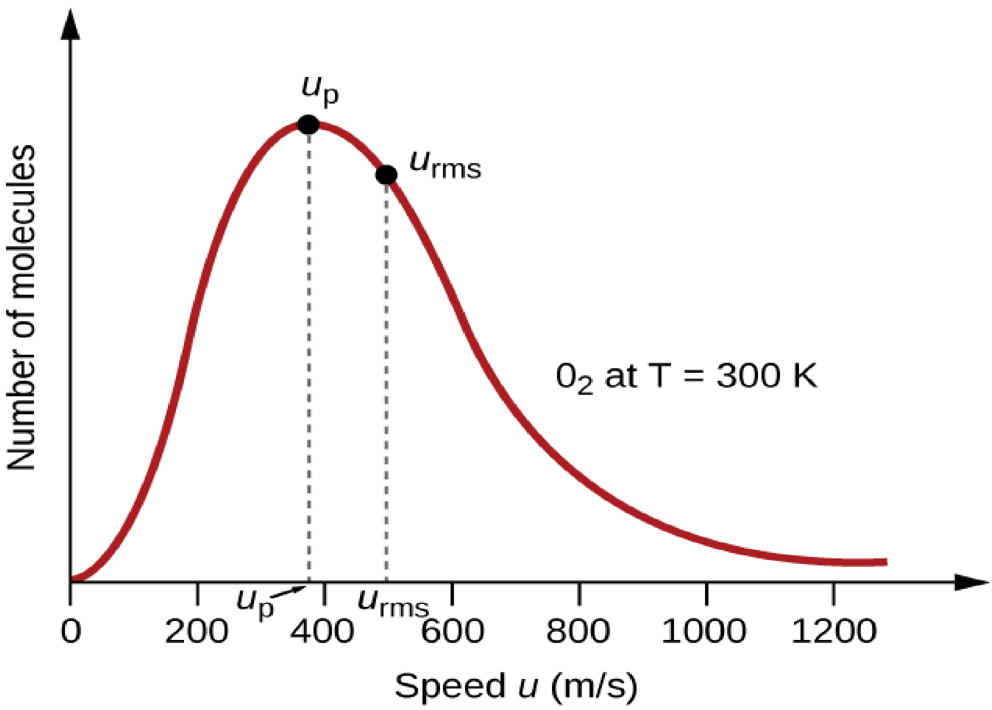


FIGURE-1-The molecular speed distribution for oxygen gas at 300 K is shown here. Very few molecules move at either very low or very high speeds. The number of molecules with intermediate speeds increases rapidly up to a maximum, which is the most probable speed, then drops off rapidly. Note that the most probable speed, vp, is a little less than 400 m/s, while the root mean square speed, urms, is closer to 500 m/s.

When mass is expressed in kilograms and speed in meters per second, the resulting energy values are obtained in joules ($J = kg{\cdot}m^2/s^2$).To deal with the large number of gas molecules, averages of speed and kinetic energy are used.

In the **Kinetic Molecular Theory (KMT)**, the **root mean square speed ($u_{(rms)}$)** of the particles is defined as the square root of the average of the squares of their speeds, where *n* is the number of particles:

$$u_{rms} = \sqrt{\overline{u^2}} = \sqrt{\frac{u_1^2 + u_2^2 + u_3^2 + u_4^2 + \ldots}{n}}$$

Therefore, the average kinetic energy of one mole of particles, **KE_{avg}**, is equal to:

$KE_{avg}=1/2Mu^2_{rms}$

where *M* is the molar mass expressed in units of kg/mol. The KE_{avg} of a mole of gas molecules is also directly proportional to the temperature of the gas and may be described by the equation:

$$\mathbf{KE_{avg}=3/\ 2RT}$$

where *R* is the gas constant and T is the kelvin temperature. When used in this equation, the appropriate form of the gas constant is 8.314 J/mol·K (8.314 kg $m^2s^{-2}mol^{-1}K^{-1}$). These two separate equations for KE_{avg} may be combined and rearranged to yield a relation between molecular speed and temperature:

$$\frac{1}{2}Mu_{rms}^2 = \frac{3}{2}RT$$

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

Calculation of u_{rms}

Calculate the root-mean-square velocity for a nitrogen molecule at 30 °C.

Solution

Convert the temperature into Kelvin:

$$\mathbf{30°C+273=303K}$$

Determine the molar mass of nitrogen in kilograms:

$$\frac{28.0\ \cancel{g}}{1\ mol} \times \frac{1\ kg}{1000\ \cancel{g}} = 0.028\ kg/mol$$

Replace the variables and constants in the root-mean-square velocity equation, replacing Joules with the equivalent kg m^2s^{-2}:

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

$$u_{rms} = \sqrt{\frac{3(8.314\ J/mol\ K)(303\ K)}{(0.028\ kg/mol)}} = \sqrt{2.70 \times 10^5\ m^2s^{-2}} = 519\ m/s$$

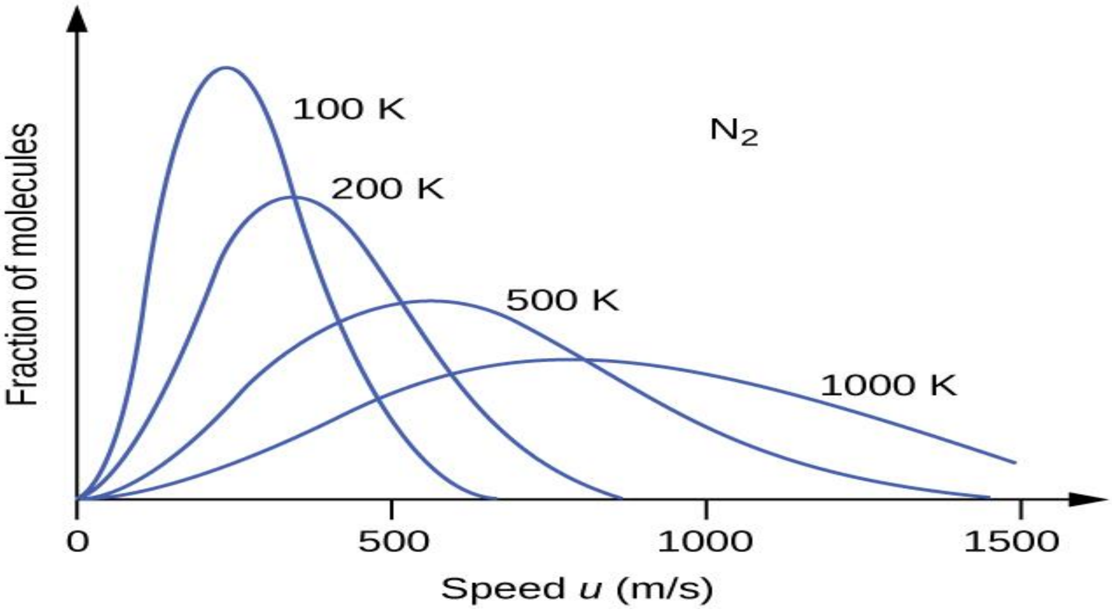
Check Your Learning

Calculate the root-mean-square velocity for a mole of oxygen molecules at –23 °C

Answer:

441 m/s

If the temperature of a gas increases, its KE_{avg} increases, more molecules have higher speeds and fewer molecules have lower speeds, and the distribution shifts toward higher speeds overall, that is, to the right. If temperature decreases, KE_{avg} decreases, more molecules have lower speeds and fewer molecules have higher speeds, and the distribution shifts toward lower speeds overall, that is, to the left. This behavior is illustrated for nitrogen gas in At a given temperature, all gases have the same KE_{avg} for their molecules. Gases composed of lighter molecules have more high-speed particles and



a higher u_{rms}, with a speed distribution that peaks at relatively higher velocities. Gases consisting of heavier molecules have more low-speed particles, a lower u_{rms}, and a speed distribution that peaks at relatively lower velocities. This trend is demonstrated by the data for a series of noble gases shown in .

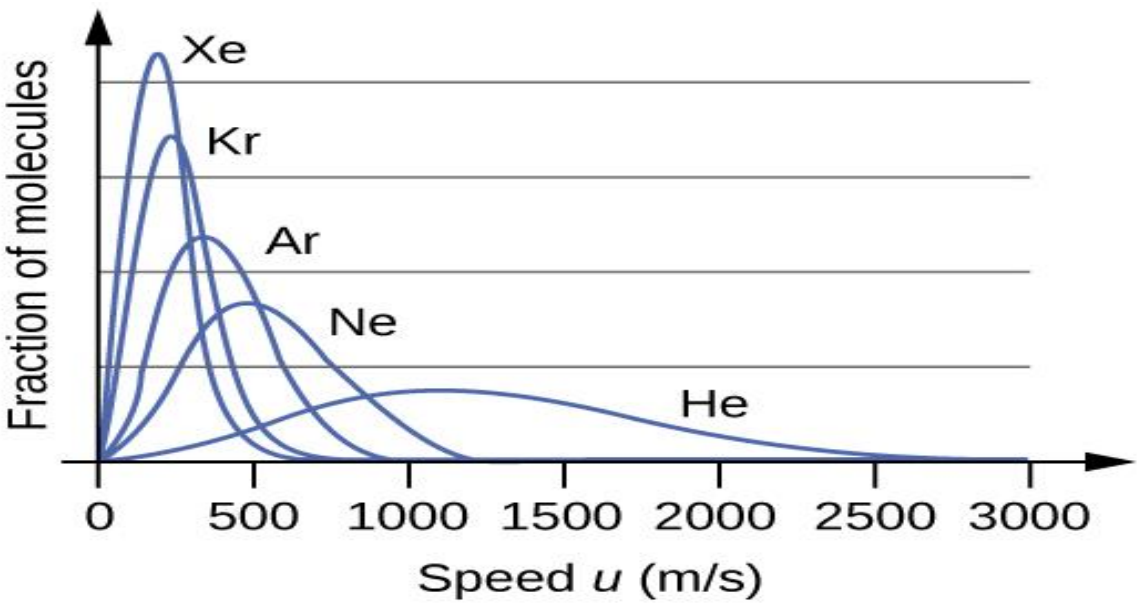


FIGURE-3- Molecular velocity is directly related to molecular mass. At a given temperature, lighter molecules move faster on average than heavier molecules.

The Kinetic-Molecular Theory Explains the Behavior of Gases, Part II

According to Graham's law, the molecules of a gas are in rapid motion and the molecules themselves are small. The average distance between the molecules of a gas is large compared to the size of the molecules. As a consequence, gas molecules can move past each other easily and diffuse at relatively fast rates.

The rate of effusion of a gas depends directly on the (average) speed of its molecules:

$$\text{effusionrate} \propto u_{rms}$$

Using this relation, and the equation relating molecular speed to mass, Graham's law may be easily derived as shown here:

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

$$M = \frac{3RT}{u_{rms}^2} = \frac{3RT}{\bar{u}^2}$$

$$\frac{effusion\ rate\ A}{effusion\ rate\ B} = \frac{u_{rms\ A}}{u_{rms\ B}} = \frac{\sqrt{\frac{3RT}{M_A}}}{\sqrt{\frac{3RT}{M_B}}} = \sqrt{\frac{M_B}{M_A}}$$

The ratio of the rates of effusion is thus derived to be inversely proportional to the ratio of the square roots of their masses. This is the same relation observed experimentally and expressed as Graham's law.

Kinetic Molecular Theory Summary

Key Concepts and Summary

The kinetic molecular theory is a simple but very effective model that effectively explains ideal gas behavior. The theory assumes that gases consist of widely separated molecules of negligible volume that are in constant motion, colliding elastically with one another and the walls of their container with average velocities determined by their absolute temperatures. The individual molecules of a gas exhibit a range of velocities, the distribution of these

Key Equations

$u_{rms} = \sqrt{\overline{u^2}} = \sqrt{\frac{u_1^2 + u_2^2 + u_3^2 + u_4^2 + \dots}{n}}$
$KE_{avg} = \frac{3}{2}RT$
$u_{rms} = \sqrt{\frac{3RT}{M}}$