

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{(101 \text{ kPa})(20 \text{ L})}{(8.3145 \text{ L} \cdot \text{kPa} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}) \cdot 300 \text{ K}} = 0.8 \text{ mol}$$

Example 2

Calculate the number of moles of gas contained within a bouncy house with a volume of 20.63 cubic meters, a temperature of 300 Kelvin, and a pressure of 101 kPa.

$$PV = nRT$$

$$\frac{PV}{RT} = n \cdot n = \frac{101 \text{ kPa} \cdot (20.63 \text{ cubic meters})}{(8.3143 \text{ J/mol}) \cdot K(300K)} \cdot n = 835.34 \text{ mols}$$

The ideal gas equation enables us to examine the relationship between the non-constant properties of ideal gases (n , P , V , T) as long as three of these properties remain fixed.

For the ideal gas equation, note that the product PV is directly proportional to T . This means that if the gas' temperature remains constant, pressure or volume can increase as long as the complementary variable decreases; this also means that if the gas' temperature changes, it may be due in part to a change in the variable of pressure or volume.

The ideal gas equation is a valuable tool that can give a very good approximation of gases at high temperatures and low pressures.

5.3.2: Density Calculations

A reformulation of the Ideal Gas Equation involving density allows us to evaluate the behaviors of ideal gases of unknown quantity.

Learning Objective

Apply the reformulated Ideal Gas Equation in your calculations

Key Points

- Density calculations allow us to evaluate the behaviors of gases of unknown volume.
- We can determine the density of an ideal gas using knowledge of three properties of the evaluated ideal gas.
- This reformulation of the Ideal Gas Equation relates pressure, density, and temperature of an ideal gas independent of the volume or quantity of gas.

Key Term

density

a measure of the amount of matter
contained by a given volume

Key Points

- Density calculations allow us to evaluate the behaviors of gases of unknown volume.
- We can determine the density of an ideal gas using knowledge of three properties of the evaluated ideal gas.
- This reformulation of the Ideal Gas Equation relates pressure, density, and temperature of an ideal gas independent of the volume or quantity of gas.

Key Term

density

a measure of the amount of matter
contained by a given volume

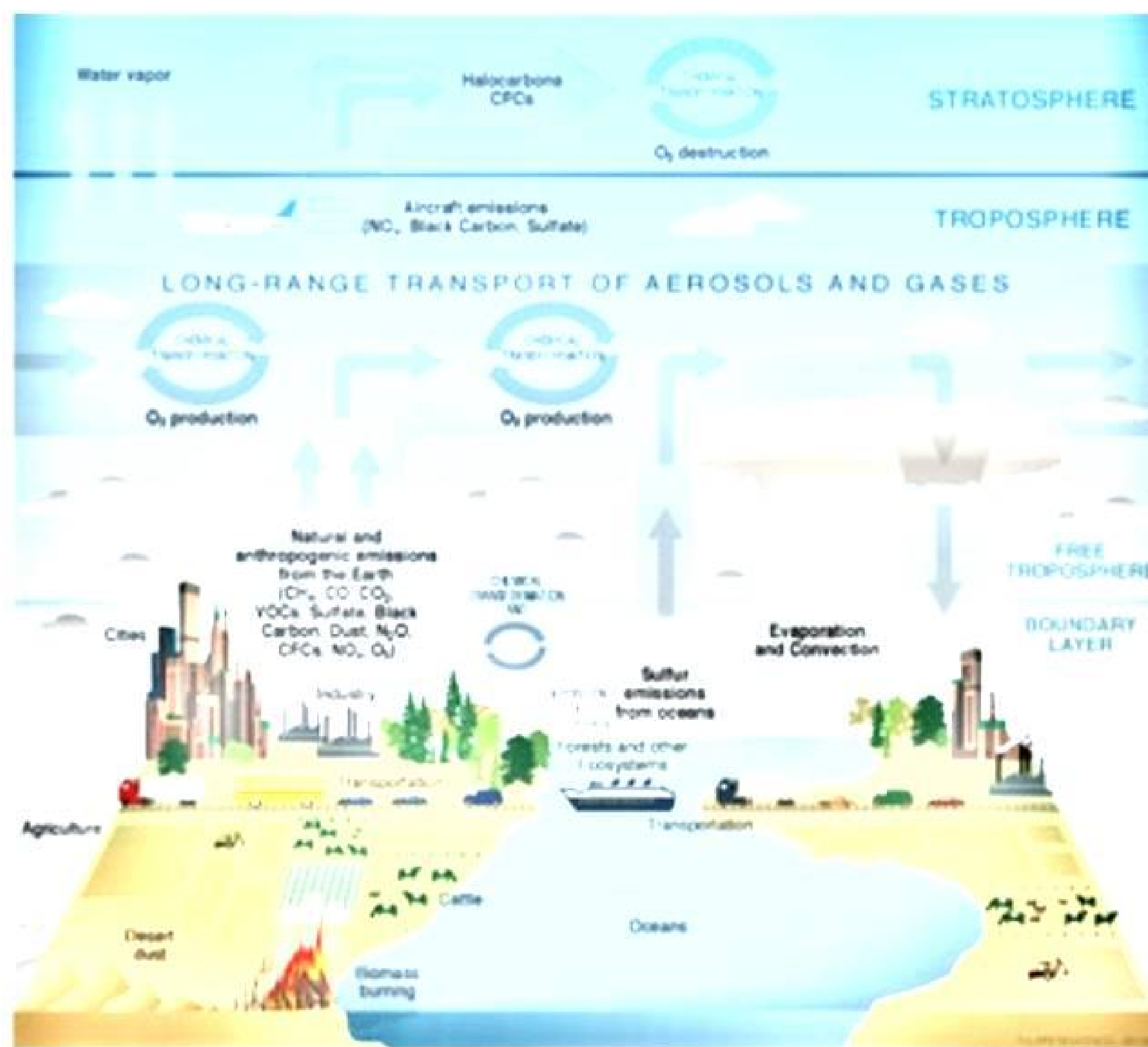
The Ideal Gas Equation in the form $PV = nRT$ is an excellent tool for understanding the relationship between the pressure, volume, amount, and temperature of an ideal gas in a defined environment that can be controlled for constant volume. However, in its most common form, the Ideal Gas Equation is not useful for examining the behavior of gases of undetermined volume, such as the gases in the clouds that surround the stars in our solar system or the atmospheric gases that support life on our planet. To derive a form of the ideal gas equation that has broader applications, we can use calculations that employ the physical property of density.





Astronomical applications of the Ideal Gas Law

The Taurus Molecular Cloud consists of dust and various gases, including hydrogen and helium. The density form of the Ideal Gas Equation may be of theoretical use when studying such astronomical phenomena as star formation.



Atmosphere composition

Atmospheric science offers one plausible real-life application of the density form of the ideal gas equation. Earth's atmosphere is composed of gases that support life—as the image shows, oxygen involved in the water cycle, sulfur emissions from oceans, methane from agriculture, and more. The density form of the Ideal Gas Law enables us to study the behavior of these gases without enclosing them in a container of known volume.

Derivation of the Volume-Independent Ideal Gas Law



We know the Ideal Gas Equation in the form $PV = nRT$. We also know that:

$$n = \# \text{ moles of gas} = \frac{\text{mass of gas (m)}}{\text{molecular weight (M)}} = \frac{m}{M}$$

If we substitute $\frac{m}{M}$ for n :

$$PV = \frac{m}{M}RT$$

Rearranging the above equation, we get:

$$\frac{P}{RT} = \frac{m}{MV}$$

Now, recall that density is equal to mass divided by volume:

$$D = \frac{m}{V}$$

The term $\frac{m}{V}$ appears on the right-hand side of the above rearranged Ideal Gas Law. We can substitute in density, D , and get the following:

$$\frac{P}{RT} = \frac{D}{M}$$

Rearranging in terms of D , we have:

$$D = \frac{MP}{RT}$$

This derivation of the Ideal Gas Equation allows us to characterize the relationship between the pressure, density, and temperature of the gas sample independent of the volume the gas occupies; it also allows us to determine the density of a gas sample given its pressure and temperature, or determine the molar mass of a gas sample given its density.

5.3.3: Molar Mass of Gas

We can derive a form of the Ideal Gas Equation, $PV=nRT$, that incorporates the molar mass of the gas (M , $g * mol^{-1}$).

Learning Objective

Apply knowledge of molar mass to the Ideal Gas Law

Key Points

- The molecular weight (molar mass) of any gas is the mass of one particle of that gas multiplied by Avogadro's number (6.02×10^{23}).
- Knowing the molar mass of an element or compound can help us stoichiometrically balance a reaction equation.
- The average molar mass of a mixture of gases is equal to the sum of the mole fractions of each gas (x_i) multi-

plied by the molar mass (M_i) of that particular gas: $\bar{M} = \sum_i x_i M_i$.

Key Terms

stoichiometry

the study and calculation of quantitative (measurable) relationships of the reactants and products in chemical reactions (chemical equations)

molar mass

the mass of one mole of an element or compound

ideal gas

a hypothetical gas whose molecules exhibit no interaction and undergo elastic collision with each other and with the walls of the container

Molar Mass of Gases and Gas Mixtures

Molar mass (M) is equal to the mass of one mole of a particular element or compound; as such, molar masses are expressed in units of grams per mole (g mol^{-1}) and are often referred to as molecular weights. The molar mass of a particular gas is therefore equal to the mass of

a single particle of that gas multiplied by Avogadro's number (6.02×10^{23}). To find the molar mass of a mixture of gases, you need to take into account the molar mass of each gas in the mixture, as well as their relative proportion.

The average molar mass of a mixture of gases is equal to the sum of the mole fractions of each gas, multiplied by their respective molar masses:

$$\bar{M} = \sum_i x_i M_i$$

The molar volumes of all gases are the same when measured at the same temperature and pressure (22.4 L at STP), but the molar masses of different gases will almost always vary.

Calculating Molar Mass using the Ideal Gas Equation

The molar mass of an ideal gas can be determined using yet another derivation of the Ideal Gas Law: $PV = nRT$.

We can write n , number of moles, as follows:

$$n = \frac{m}{M}$$

where m is the mass of the gas, and M is the

We can write n , number of moles, as follows:

$$n = \frac{m}{M}$$

where m is the mass of the gas, and M is the molar mass. We can plug this into the Ideal Gas Equation:

$$PV = \left(\frac{m}{M}\right) RT$$

Rearranging, we get:

$$\frac{PV}{RT} = \frac{m}{M}$$

Finally, putting the equation in terms of molar mass, we have:

$$M = \frac{mRT}{PV}$$

This derivation of the Ideal Gas Equation is useful in determining the molar mass of an unknown gas.

Example

- An unknown gas with a mass of 205 g occupies a volume of 20.0 L at standard temperature and pressure. What is the molar mass of the gas?
- $M = \frac{mRT}{PV}$





Equation

$$PV = \left(\frac{m}{M}\right) RT$$

Rearranging, we get:

$$\frac{PV}{RT} = \frac{m}{M}$$

Finally, putting the equation in terms of molar mass, we have:

$$M = \frac{mRT}{PV}$$

This derivation of the Ideal Gas Equation is useful in determining the molar mass of an unknown gas.

Example

- An unknown gas with a mass of 205 g occupies a volume of 20.0 L at standard temperature and pressure. What is the molar mass of the gas?
- $M = \frac{mRT}{PV}$

$$M = \frac{(205)(0.0821)(273)}{(1.0)(20.0)} = 230 \frac{g}{mol}$$

Equation



5.4: Gas Stoichiometry

5.4.1: Gas Stoichiometry

At standard temperature and pressure, one mole of any gas will occupy a volume of 22.4 L.

Learning Objective

Calculate volumes of gases consumed/produced in a reaction using gas stoichiometry.

Key Points

- At Standard Temperature and Pressure (STP), 1 mole of any gas will occupy a volume of 22.4 L.
- The Ideal Gas Law, along with a balanced chemical equation, can be used to solve for the amount, either in volume or mass, of gas consumed or produced in a chemical reaction.

Key Term

stoichiometry

the study and calculation of quantitative (measurable) relationships of the reactants and products in chemical reactions (chemical equations)

Stoichiometry is the quantitative study of

the relative amounts of reactants and products in chemical reactions; gas stoichiometry involves chemical reactions that produce gases. Stoichiometry is based on the law of conservation of mass, meaning that the mass of the reactants must be equal to the mass of the products. This assumption can be used to solve for unknown quantities of reactants or products.

Stoichiometric Calculations Involving Ideal Gases at STP

Stoichiometric calculations involving gases allow us to convert between mass, number of moles, and most importantly, *volume* of gases. The following relationship makes this possible:

- 1 mole of *any* gas at standard temperature and pressure (273 K and 1 atm) occupies a volume of 22.4 L.

While the above relationship is an estimation, it is a relatively good approximation at STP, and can be used reliably in calculations.

Example



- According to the above reaction, what volume of $\text{NO}_2(g)$ is produced



Example



- According to the above reaction, what volume of $\text{NO}_2(g)$ is produced from the combustion of 100 g of $\text{NH}_3(g)$, assuming the reaction takes place at standard temperature and pressure?
- From the periodic table, we can determine that the molar mass of ammonia, $\text{NH}_3(g)$, is 17 g/mol, and perform the following stoichiometric calculation:
- $\left(\frac{100 \text{ g NH}_3}{17 \text{ g}}\right) \times \left(\frac{4 \text{ mol NO}_2}{4 \text{ mol NH}_3}\right) \times \left(\frac{22.4 \text{ L}}{1 \text{ mol NO}_2}\right) = 132 \text{ L NO}_2(g)$

Note the final conversion factor. Because we are told that the reaction takes place at STP, we can relate volume, 22.4 L, to 1 mol NO_2 .

Attributions

- Gas Stoichiometry
 - "Stoichiometry."
http://en.wikipedia.org/wiki/Stoichiometry%23Gas_stoichiometry. Wikipedia CC BY-SA 3.0.
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<http://en.wiktionary.org/wiki/stoichiometry>.



5.5: Partial Pressure

5.5.1: Dalton's Law of Partial Pressure

Dalton's Law of Partial Pressure states the total pressure exerted by a mixture of gases is equal to the sum of the partial pressure of each individual gas.

Learning Objective

Demonstrate an understanding of partial pressures and mole fractions.

Key Points

- The total pressure of a mixture of gases can be defined as the sum of the pressures of each individual gas:
$$P_{total} = P_1 + P_2 + \dots + P_n .$$
- The partial pressure of an individual gas is equal to the total pressure multiplied by the mole fraction of that gas.
- Boyle's Law and the Ideal Gas Law tell us the total pressure of a mixture depends solely on the number of moles of gas, and not the kinds of molecules; Dalton's Law allows us to calculate the total pressure in a system from each gas' individual contribution.

Key Terms

Key Terms

mole fraction

number of moles of one particular gas divided by the total moles of gas in the mixture

Dalton's Law of Partial Pressures

the total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of each individual gas; also known as Dalton's Law of Partial Pressures

Because it is dependent solely the number of particles and not the identity of the gas, the Ideal Gas Equation applies just as well to mixtures of gases as it does to pure gases. In fact, it was with a gas mixture—ordinary air—that Boyle, Gay-Lussac, and Charles performed their early experiments. The only new concept we need to deal with gas mixtures is partial pressure, a concept invented by the famous English chemist John Dalton (1766-1844). Dalton correctly reasoned that the low density and high compressibility of gases were indicative of the fact that they consisted mostly of empty space; from this, it Dalton concluded that when two or more different gases occupy the same volume, they behave entirely independently of one another.



Dalton's Law (also called Dalton's Law of Partial Pressures) states that the total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of individual gases. Mathematically, this can be stated as follows:

$$P_{total} = P_1 + P_2 + \dots + P_n$$

where P_1 , P_2 and P_n represent the partial pressures of each compound. It is assumed that the gases do not react with each other.

Example

A 2.0 L container is pressurized with 0.25 atm of oxygen gas and 0.60 atm of nitrogen gas. What is the total pressure inside the container?

$$P_{total} = P_{O_2} + P_{N_2} = 0.25 + 0.60 = 0.85 \text{ atm}$$

The total pressure inside the container is 0.85 atm.

Calculating the Mole Fraction

The mole fraction is a way of expressing the relative proportion of one particular gas within a mixture of gases. We do this by dividing the

a mixture of gases. The mole fraction of a particular gas i is defined as the number of moles of a particular gas i by the total number of moles in the mixture:

$$x_i = \frac{\text{number of moles } i}{\text{total number moles of gas}}$$

Example

A 3.0 L container contains 4 mol He, 2 mol Ne, and 1 mol Ar. What is the mole fraction of neon gas?

$$x_{Ne} = \frac{\text{number of moles Ne}}{\text{total number moles of gas}} = \frac{2}{4+2+1} = \frac{2}{7}$$

The mole fraction of neon gas is $2/7$ or 0.28.

Calculating Partial Pressure

The partial pressure of one individual gas within the overall mixtures, p_i , can be expressed as follows:

$$P_i = P_{total} x_i$$

where x_i is the mole fraction.

Example

A mixture of 2 mol H_2 and 3 mol He exerts a total pressure of 3 atm. What is the partial

pressure of He?

$$P_{He} = P_{total} x_{He} = (3) \left(\frac{3}{5}\right) = \frac{9}{5} \text{ atm}$$

Calculating Total Pressure

We know from Boyle's Law that the total pressure of the mixture depends solely on the number of moles of gas, regardless of the types and amounts of gases in the mixture; the Ideal Gas Law reveals that the pressure exerted by a mole of molecules does not depend on the identity of those particular molecules; Dalton's Law now allows us to calculate the total pressure in a system when we know each gas individual contribution.

Example

Consider a container of fixed volume 25.0 L. We inject into that container 0.78 moles of N_2 gas at 298 K. From the Ideal Gas Law, we can easily calculate the measured pressure of the nitrogen gas to be 0.763 atm.

We now take an identical container of fixed volume 25.0 L, and we inject into that container 0.22 moles of O_2 gas at 298K. The measured pressure of the oxygen gas is 0.215 atm.



As a third measurement, we inject 0.22 moles of O_2 gas at 298K into the first container, which already has 0.78 moles of N_2 . (Note that the mixture of gases we have prepared is very similar to that of air.) The measured pressure in this container is now found to be 0.975 atm.

Our data show that the total pressure of the mixture of N_2 and O_2 in the container is equal to the sum of the pressures of the N_2 and O_2 samples taken separately. We now define the partial pressure of each gas in the mixture to be the pressure of each gas as if it were the only gas present. Our measurements demonstrate that the partial pressure of N_2 as part of the gas P_{N_2} is 0.763 atm, and the partial pressure of O_2 as part of the gas P_{O_2} , is 0.215 atm.

5.5.2: Collecting Gases Over Water

The amount of gas present can be determined by collecting a gas over water and applying Dalton's Law.

Learning Objective

Apply Dalton's Law to determine the partial pressure of a gas collected over water.

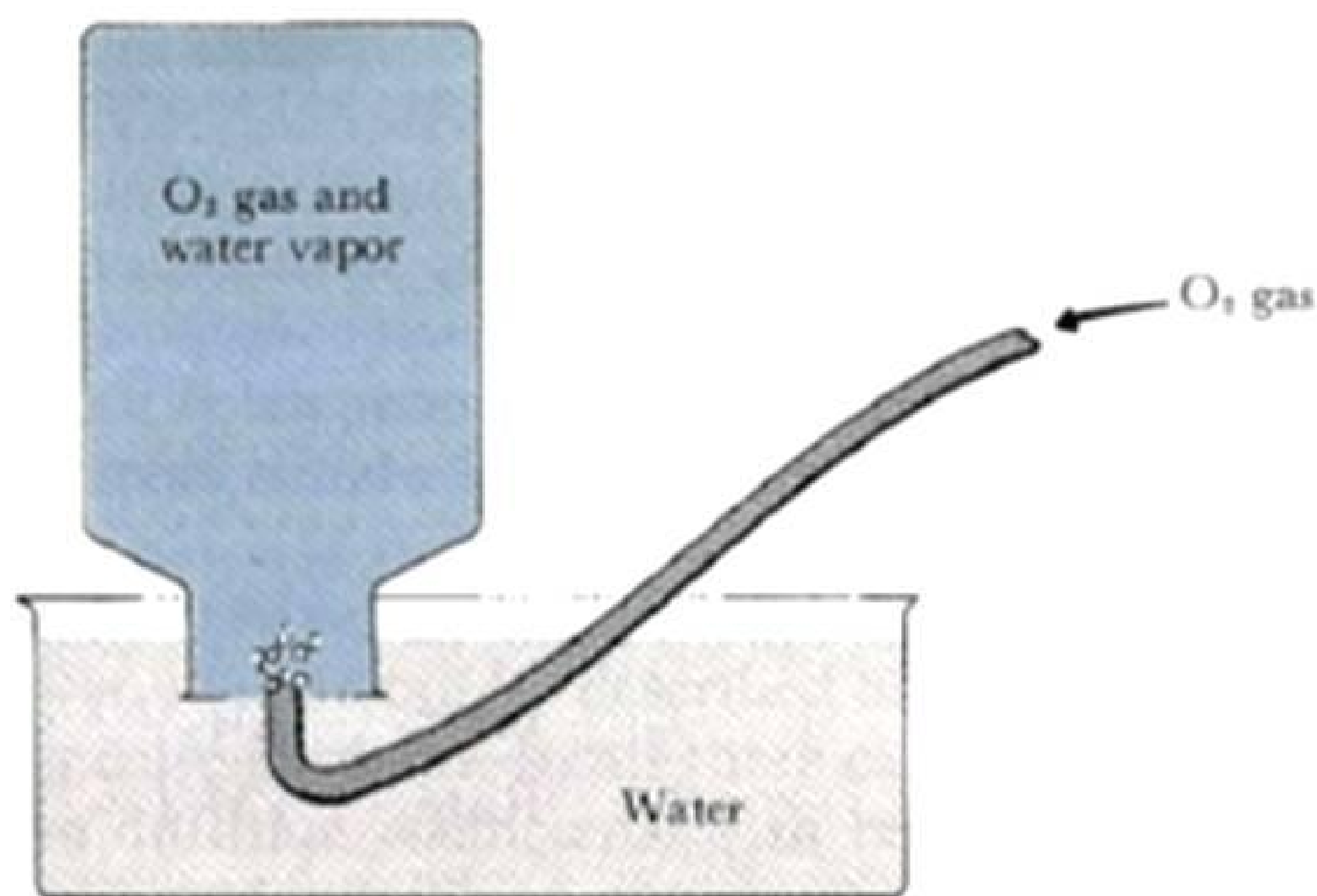
Key Points

Key Term

pneumatic trough

device used to collect a gas over water; the height of water displaced in the tube can be used to determine the total pressure inside the tube

Since gases have such small densities, it can be difficult to measure their mass. A common way to determine the amount of gas present is by collecting it over water and measuring the height of displaced water; this is accomplished by placing a tube into an inverted bottle, the opening of which is immersed in a larger container of water.



Collecting gas over water

When collecting oxygen gas and calculating its partial pressure by displacing water from an



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its partial pressure by displacing water from an inverted bottle, the presence of water vapor in the collecting bottle must be accounted for; this is easily accomplished using Dalton's Law of Partial Pressures.

The Pneumatic Trough

This arrangement is called a pneumatic trough, and it was widely used in the early days of chemistry. As the gas enters the bottle, it displaces the water and becomes trapped in the closed, upper part of the bottle. You can use this method to measure a pure gas (i.e. O_2) or the amount of gas produced from a reaction. The collected gas is not the only gas in the bottle, however; keep in mind that liquid water itself is always in equilibrium with its vapor phase, so the space at the top of the bottle is actually a mixture of two gases: the gas being collected, and gaseous H_2O . The partial pressure of H_2O is known as the vapor pressure of water and is dependent on the temperature. To determine the quantity of gas we have collected alone, we must subtract the vapor pressure of water from the total vapor pressure of the mixture.

Calculating Gas Volume

Example 1

O₂ gas is collected in a pneumatic trough with a volume of 0.155 L until the height of the water inside the trough is equal to the height of the water outside the trough. The atmospheric pressure is 754 torr, and the temperature is 295 K. How many moles of oxygen are present in the trough? (At 295 K, the vapor pressure of water is 19.8 torr.)

The total pressure in the tube can be written using Dalton's Law of Partial Pressures:

$$P_{total} = P_{H_2O} + P_{O_2}$$

Rearranging this in terms of P_{O_2} , we have:

$$P_{O_2} = P_{total} - P_{H_2O}$$

Because the height of the water inside the tube is equal to the height of the water outside the tube, the total pressure inside the tube must be equal to the atmospheric pressure. With substitution, we have:

$$P_{O_2} = P_{total} - P_{H_2O} = 754 - 19.8 = 734 \text{ torr} = .966 \text{ atm}$$

Next, we apply the Ideal Gas Law:

$$n = \frac{PV}{RT}$$

$$= \frac{(.966 \text{ atm})(.155 \text{ L})}{(.082 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})(295 \text{ K})}$$

$$= .00619 \text{ mol } O_2$$

Example 2

Oxygen gas generated in an experiment is collected at 25°C in a bottle inverted in a trough of water. The external laboratory pressure is 1.000 atm. When the water level in the originally full bottle has fallen to the level in the trough, the volume of collected gas is 1750 ml. How many moles of oxygen gas have been collected?

If the water levels inside and outside the bottle are the same, then the total pressure inside the bottle equals 1.000 atm; at 25°C, the vapor pressure of water (or the pressure of water vapor in equilibrium with the liquid) is 23.8 mm Hg or 0.0313 atm.

Therefore, the partial pressure of oxygen gas is 1.000 - 0.031, or 0.969 atm.

The mole fraction of oxygen gas in the bottle is 0.969 (not 1.000), and the partial pressure of oxygen also is 0.969 atm. The number of moles is:

$$n = \frac{PV}{RT} = \frac{(.969 \text{ atm})(1730 \text{ cm}^3)}{(82.054 \text{ cm}^3 \text{ K}^{-1} \text{ mole}^{-1})(298 \text{ K})}$$

$$n = 0.068 \text{ moles O}_2$$

Attributions

- Dalton's Law of Partial Pressure
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<http://www.boundless.com/>. Boundless Learning
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Concept Development

5.6: Kinetic Molecular Theory

5.6.1: Kinetic Molecular Theory and Gas Laws

Kinetic Molecular Theory explains the macroscopic properties of gases and can be used to understand and explain the gas laws.

Learning Objective

Express the five basic assumptions of the Kinetic Molecular Theory of Gases.

Key Points

- Kinetic Molecular Theory states that gas particles are in constant motion and exhibit perfectly elastic collisions.
- Kinetic Molecular Theory can be used to explain both Charles' and Boyle's Laws.
- The average kinetic energy of a collection of gas particles is directly proportional to absolute temperature only.

Key Terms

macroscopic properties

properties that can be visualized or measured by the naked eye; examples in-

clude pressure, temperature, and volume

ideal gas

a hypothetical gas whose molecules exhibit no interaction and undergo elastic collision with each other and the walls of the container

Basic Assumptions of the Kinetic Molecular Theory

By the late 19th century, scientists had begun accepting the atomic theory of matter started relating it to individual molecules. The Kinetic Molecular Theory of Gases comes from observations that scientists made about gases to explain their macroscopic properties. The following are the basic assumptions of the Kinetic Molecular Theory:

1. The volume occupied by the individual particles of a gas is negligible compared to the volume of the gas itself.
2. The particles of an ideal gas exert no attractive forces on each other or on their surroundings.
3. Gas particles are in a constant state of random motion and move in straight lines until they collide with another body.
4. The collisions exhibited by gas particles are completely elastic.

particles are completely elastic; when two molecules collide, total kinetic energy is conserved.

5. The average kinetic energy of gas molecules is directly proportional to absolute temperature only; this implies that all molecular motion ceases if the temperature is reduced to absolute zero.

Applying Kinetic Theory to Gas Laws

Charles' Law states that at constant pressure, the volume of a gas increases or decreases by the same factor as its temperature. This can be written as:

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

According to Kinetic Molecular Theory, an increase in temperature will increase the average kinetic energy of the molecules. As the particles move faster, they will likely hit the edge of the container more often. If the reaction is kept at constant pressure, they must stay farther apart, and an increase in volume will compensate for the increase in particle collision with the surface of the container.

Boyle's Law states that at constant temperature, the absolute pressure and volume of a given mass of confined gas are inversely

At a given mass of confined gas are inversely proportional. This relationship is shown by the following equation:

$$P_1 V_1 = P_2 V_2$$

At a given temperature, the pressure of a container is determined by the number of times gas molecules strike the container walls. If the gas is compressed to a smaller volume, then the same number of molecules will strike against a smaller surface area; the number of collisions against the container will increase, and, by extension, the pressure will increase as well. Increasing the kinetic energy of the particles will increase the pressure of the gas.

5.6.2: Distribution of Molecular Speeds and Collision Frequency

The Maxwell-Boltzmann Distribution describes the average molecular speeds for a collection of gas particles at a given temperature.

Learning Objective

Identify the relationship between velocity distributions and temperature and molecular weight of a gas.

Key Points

- Gaseous particles move at random speeds and in random directions.
- The Maxwell-Boltzmann Distribution describes the average speeds of a collection gaseous particles at a given temperature.
- Temperature and molecular weight can affect the shape of Boltzmann Distributions.
- Average velocities of gases are often expressed as root-mean-square averages.

Key Terms

quanta

the smallest possible packet of energy that can be transferred or absorbed

velocity

a vector quantity that denotes the rate of change of position with respect to time or a speed with a directional component

According to the Kinetic Molecular Theory, all gaseous particles are in constant random motion at temperatures above absolute zero. The movement of gaseous particles is characterized by straight-line trajectories interrupted by collisions with other particles or with a physical boundary. Depending on the nature of the particles' relative kinetic energies, a collision causes a transfer of kinetic energy as well as a change in direction.

Root-Mean-Square Velocities of Gaseous Particles

Measuring the velocities of particles at a given time results in a large distribution of values; some particles may move very slowly, others very quickly, and because they are constantly moving in different directions, the velocity could equal zero. (Velocity is a vector quantity, equal to the speed and direction of a particle) To properly assess the average velocity, average the *squares* of the velocities and take the square root of that value. This is known as the root-mean-square (RMS) velocity, and it is represented as follows:

$$\bar{v} = v_{rms} = \sqrt{\frac{3RT}{M_m}}$$

$$KE = \frac{1}{2}mv^2$$

$$KE = \frac{1}{2}mv^2$$

In the above formula, R is the gas constant, T is absolute temperature, and M_m is the molar mass of the gas particles in kg/mol.

Energy Distribution and Probability



Consider a closed system of gaseous particles with a fixed amount of energy. With no external forces (e.g. a change in temperature) acting on the system, the total energy remains unchanged. In theory, this energy can be distributed among the gaseous particles in many ways, and the distribution constantly changes as the particles collide with each other and with their boundaries. Given the constant changes, it is difficult to gauge the particles' velocities at any given time. By understanding the nature of the particle movement, however, we can predict the probability that a particle will have a certain velocity at a given temperature.

Kinetic energy can be distributed only in discrete amounts known as quanta, so we can assume that any one time, each gaseous particle has a certain amount of quanta of kinetic energy. These quanta can be distributed among the three directions of motions in various ways, resulting in a velocity state for the molecule; therefore, the more kinetic energy, or quanta, a particle has, the more velocity states it has as well.

If we assume that all velocity states are equally probable, higher velocity states are favorable because there are greater in quantity. Although higher velocity states are favored statistically, however, lower energy states are

equally probable, higher velocity states are favorable because there are greater in quantity. Although higher velocity states are favored statistically, however, lower energy states are more likely to be occupied because of the limited kinetic energy available to a particle; a collision may result in a particle with greater kinetic energy, so it must also result in a particle with less kinetic energy than before.

Maxwell-Boltzmann Distributions

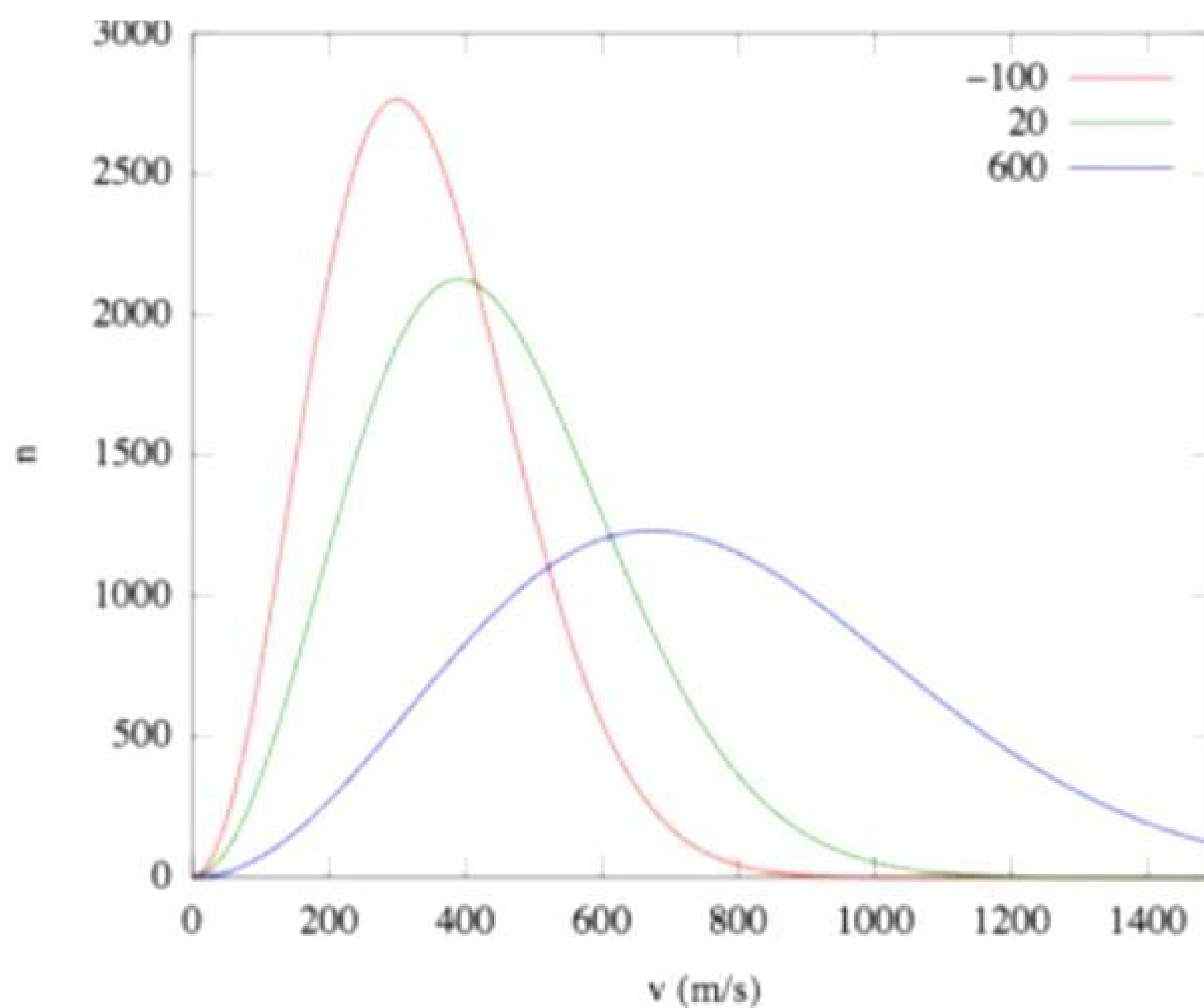
Using the above logic, we can hypothesize the velocity distribution for a given group of particles by plotting the number of molecules whose velocities fall within a series of narrow ranges. This results in an asymmetric curve, known as the Maxwell-Boltzmann distribution. The peak of the curve represents the most probable velocity among a collection of gas particles.

Velocity distributions are dependent on the temperature and mass of the particles. As the temperature increases, the particles acquire more kinetic energy. When we plot this, we see that an increase in temperature causes the Boltzmann plot to spread out, with the relative maximum shifting to the right.

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Effect of temperature on root-mean-square speed distributions

As the temperature increases, so does the average kinetic energy (v), resulting in a wider distribution of possible velocities. n = the fraction of molecules.

Larger molecular weights narrow the velocity distribution because all particles have the same kinetic energy at the same temperature. Therefore, by the equation $KE = \frac{1}{2}mv^2$, the fraction of particles with higher velocities will increase as the molecular weight decreases.

5.6.3: Root-Mean-Square Speed

The root-mean-square speed measures the



The root-mean-square speed measures the average speed of particles in a gas, defined as

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

Learning Objective

Recall the mathematical formulation of the root-mean-square velocity for a gas.

Key Points

- All gas particles move with random speed and direction.
- Solving for the average velocity of gas particles gives us the average velocity of zero, assuming that all particles are moving equally in different directions.
- You can acquire the average speed of gaseous particles by taking the root of the square of the average velocities.
- The root-mean-square speed takes into account both molecular weight and temperature, two factors that directly affect a material's kinetic energy.

Key Term

velocity

a vector quantity that denotes the rate of change of position, with respect to time or a speed with a directional component

Kinetic Molecular Theory and Root-Mean-Square Speed

According to Kinetic Molecular Theory, gaseous particles are in a state of constant random motion; individual particles move at different speeds, constantly colliding and changing directions. We use velocity to describe the movement of gas particles, thereby taking into account both speed and direction.

Although the velocity of gaseous particles is constantly changing, the distribution of velocities does not change. We cannot gauge the velocity of each individual particle, so we often reason in terms of the particles' average behavior. Particles moving in opposite directions have velocities of opposite signs. Since a gas' particles are in random motion, it is plausible that there will be about as many moving in one direction as in the opposite direction, meaning that the average velocity for a collection of gas particles equals zero; as this value is unhelpful, the average of velocities can be determined using an alternative method.

By squaring the velocities and taking the square root, we overcome the "directional" component of velocity and simultaneously acquire the particles' average velocity. Since the value excludes the particles' direction, we now

refer to the value as the average speed. The root-mean-square speed is the measure of the speed of particles in a gas, defined as the square root of the average velocity-squared of the molecules in a gas.

It is represented by the equation:

$v_{rms} = \sqrt{\frac{3RT}{M}}$, where v_{rms} is the root-mean-square of the velocity, M_m is the molar mass of the gas in *kilograms per mole*, R is the molar gas constant, and T is the temperature in Kelvin.

The root-mean-square speed takes into account both molecular weight and temperature, two factors that directly affect the kinetic energy of a material.

Example

- What is the root-mean-square speed for a sample of oxygen gas at 298 K?

$$v_{rms} = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3(8.3145 \frac{J}{K \cdot mol})(298 K)}{32 \times 10^{-3} \frac{kg}{mol}}} = 482 m/s$$

Attributions

- Kinetic Molecular Theory and Gas Laws
 - "Boundless."