

GASES

Section A: Introduction

Gas behaviour is well understood and can be expressed in terms of simple mathematical models.

One objective of scientists is to develop models of natural phenomena, and a study of gas behaviour allows us to achieve this.

Further some common elements and compounds exist in the gaseous state under normal conditions of pressure and temperature, thus a study of gas behaviour will allow us to comprehend the nature of gaseous elements and compounds.

Furthermore, many common liquids can be vaporized, and the properties of these vapours are important.

(FYI: a gas is a substance that is normally in the gaseous state at ordinary pressures and temperatures. A vapour is the gaseous form of a substance that is normally a liquid or solid at ordinary pressures and temperatures. It is therefore correct to refer to oxygen gas and water vapour.)

Our gaseous atmosphere provides one means of transferring energy and material throughout the globe, and it is the source of life-giving chemicals, hence a study of the nature of the gas phase will further enhance our understanding of the atmosphere.

Section B: Gas Pressure

A gas is a state of matter. A substance in the gaseous state takes that shape of the container and its volume. To describe the gaseous state, four quantities are needed:

1. the *quantity* of gas, n (in moles);
2. the *temperature* of the gas, T (in kelvins);
3. the *volume* of gas, V (in liters); and
4. the *pressure* of the gas, P (in kPa)

Gas molecules are in constant motion, colliding with each other and with the walls of their container. These impacts cause gas pressure. The pressure of a gas is a measure of the force it exerts on its container.

A gas, such as our atmosphere, exerts a pressure on every surface.

Atmospheric pressure can be measured with a **barometer**, which can be made by filling a tube with a liquid and inverting the tube in a dish containing the same liquid.

A mercury-filled barometer is made by inverting a column of completely filled with mercury in a mercury dish. Some of the mercury flows out when the tube is inverted, the space in the tube above the mercury is nearly a vacuum; only a negligible amount of mercury vapour occupies it.

Draw a diagram of the mercury barometer:

At sea level the height of the mercury column is about 760 mm above the surface of the mercury in the dish. The pressure of the atmosphere on the surface of the mercury in the dish is balanced by the downward pressure exerted by the mercury in the column.

Pressure exerted by a mercury barometer is recorded in units of millimeters of mercury (**mm Hg**), a unit sometimes called the **torr** in honour of Evangelista Torricelli (1608 – 1647), who invented the mercury barometer in 1643.

The **standard atmosphere (atm)** is defined as:

$$1 \text{ standard atmosphere} = 1 \text{ atm} = 760 \text{ mm Hg}$$

Newton's first law states that force = mass x acceleration.

The SI units of mass and acceleration are the kilogram (kg) and the meter per second squared (m/s^2), respectively, so the units of force are $\text{kg} \cdot \text{m/s}^2$. This derived unit is given the name **newton (N)**.

Pressure is defined as the force exerted on an object divided by the area over which the force is exerted.

$$\text{Pressure} = \frac{\text{force}}{\text{Area}} \quad (\text{N/m}^2)$$

The SI unit of pressure is the **pascal (Pa)**, named for the French mathematician and philosopher Blaise Pascal (1623–1662). It is the only pressure unit that is defined directly in terms of force per unit area.

$$1 \text{ pascal (Pa)} = 1 \text{ newton/meter}^2$$

This is a very small unit compared with ordinary pressures, so the **kilopascal (kPa)** is more often used. Hence the relationship between the units of pressure is:

$$\begin{aligned} \text{One atmosphere (atm)} &= 101.325 \text{ kilopascals (kPa)} \\ &= 760 \text{ mm Hg} \\ &= 76 \text{ cm Hg} \\ &= 760 \text{ torr} \\ &= 1.013 \text{ bar} \\ &= 14.7 \text{ lb / in}^2 \text{ (psi)} \end{aligned}$$

Atmospheric pressures are sometimes reported in the unit called the **bar**, where:

$$\begin{aligned} 1 \text{ bar} &= 100,000 \text{ Pa.} \\ &= 10^5 \text{ Pa} \\ &= 100 \text{ kPa} \end{aligned}$$

A **barometer** is any device that is used to measure the Earth's atmospheric pressure (dry air creates more pressure than wet air, it is heavier than the same amount of wet air. Consequently, high barometer readings indicate dry air- fair weather. Low reading show an increased chance of rain/showers.)

Note: Any liquid may be used to measure atmospheric pressure, the height of the column depends on the density of the liquid. A comparison of mercury, with a density of 13.6 g/cm^3 , and water (density = 1 g/cm^3) indicates that if a barometer were filled with water, the column would be almost 34 feet high!!!

Assignment:

1. Convert a pressure of 635 mm Hg into its corresponding value in units of atmospheres (atm), bars, and kilopascals (kPa). (Answer: $635 \text{ mm Hg} = 0.836 \text{ atm} = 0.847 \text{ bar} = 84.7 \text{ kPa}$)
2. Convert (a) 75 cm Hg into kPa, (b) 0.63 atm into kPa (c) 0.923 bar into atm
3. On Titan, the largest moon of Saturn, the atmospheric pressure is 1.60 atm, what is the atmospheric pressure of Titan in kilopascals?

Pressure of gases is measured by an instrument called a **manometer**, whose principle of operation is similar to that of a barometer.

1. Closed-tube manometer

This is used to measure pressures below atmospheric pressures.

The pressure is simply the difference in the heights of the mercury levels in the two arms.

2. Open-tube manometer

This instrument is used to measure gas pressures that are near atmospheric pressures.

The difference in the heights of the mercury levels in the two arms of the manometer relates the gas pressure to the atmospheric pressure.

- a. If the pressure of the gas in the bulb is the same as the atmospheric pressure, then the levels in the two arms are equal:

$$P_{\text{gas}} = h \quad (h = \text{height difference in the arms})$$

- b. If the pressure of the gas in the bulb is greater than the atmospheric pressure, then the mercury will be forced higher in the arm exposed to the atmosphere:

$$P_{\text{gas}} = P(\text{atmospheric}) + h$$

- c. If the atmospheric pressure is greater than the gas pressure, the mercury will be higher in the arm exposed to the gas:

$$P_{\text{gas}} = P(\text{atmospheric}) - h$$

Now Do: Assignment: Reading Manometers

Section C: Gas Laws

Experimentation in the 17th and 18th centuries led to the understanding of gas behaviour and to the formulation of important gas laws.

Changes in temperature and pressure have little effect on the volume of a liquid or a solid, but a considerable effect on the volume of a gas.

Various gas laws have been devised following experiments and they are summarised below.

The pressure - Volume Relationship: Boyle's Law (British chemist: Robert Boyle, 1627 – 1691)
Robert Boyle studied the compressibility of gases in 1661 and observed that *the volume of a fixed amount of gas at a given temperature is inversely proportional to the pressure exerted on the gas.*

When the pressure of the gas is plotted as a function of $1/V$, a straight line is observed. This type of plot demonstrates that the volume and pressure of a gas are inversely proportional; they change in opposite directions.

Plot of Pressure vs $1/V$:

Mathematically, we can write this two ways:

$$P \propto \frac{1}{V} \qquad V \propto \frac{1}{P} \qquad [n, T \text{ constant}]$$

(the symbol \propto means proportional to)

Both of the above relationship show that for a fixed mass of gas at a given temperature, the gas volume increases if pressure on the gas decreases, or conversely if the pressure on the gas increases then its volume decreases.

(If pressure on a fixed mass of gas at a given temperature is doubled, then volume is halved, or if the volume is tripled then the pressure is one-third of the original, etc; $2 \times P = \frac{1}{2} v$ or $\frac{1}{2} v = 2P$)

Mathematically, Boyle's Law, the relationship between pressure and volume of a gas may be expressed as:

$$P \times V = \text{constant}$$

i.e. for a given quantity of gas at a given temperature, the product of pressure and volume is a constant. The value of the constant depends on the temperature and the amount of gas in the sample

This means that if the pressure-volume product is known for one set of conditions (P_1 and V_1), it is known for all other conditions of pressure and volume, (P_2 and V_2), (P_3 and V_3), and so on ...

$$P_1V_1 = P_2V_2 = P_3V_3$$

This form of Boyle's law is useful when we want to know, for example, what happens to the volume of a given quantity of gas when the pressure changes (at a constant temperature).

The graph of P versus V is a curve : $P V = \text{constant}$

Sample Problem

A fixed quantity of gas at a constant temperature exhibits a pressure of 98.8 kPa and occupies 10.7 L. Calculate the volume the gas will occupy if the pressure is increased to 102.2 kPa.

Solution

original conditions

$$P_1 = 98.8 \text{ kPa}$$

$$V_1 = 10.7 \text{ L}$$

final conditions

$$P_2 = 102.2 \text{ kPa}$$

$$V_2 = ?$$

We know that: $P_1V_1 = P_2V_2$

$$\text{therefore: } V_2 = \frac{P_1V_1}{P_2} = \frac{(98.8 \text{ kPa})(10.7 \text{ L})}{102.2 \text{ kPa}} = \quad \text{L}$$

(Logically the pressure on the gas was increased, thus we would expect the volume to decrease compared to the original volume.)

Now Do Assignment: 1. Pressure – Volume Relationship
2. Pressure — Volume Relationship II

The Temperature – Volume relationship: Charles's Law

(French scientist: Jacques Charles, 1746 – 1823)

In 1787, Jacques Charles discovered that the volume of a fixed quantity of gas at constant pressure increased with increasing temperature.

A graph of volume of gas versus temperature is plotted, they do not show a direct linear relationship, however when the graphs are extended (extrapolated) to lower temperatures, the line reaches zero volume at -273.15°C .

Plot of Volume of a fixed mass of gas as a function of Temperature at constant pressure:

[note: $V = 0$ at absolute zero]

The gas is predicted to have zero volume at this temperature, -273.15°C , however gases do not actually reach zero volume because all gases liquefy or solidify before reaching this temperature.

In 1848, William Thomson (1824 – 1907), a British physicist, whose title was lord Kelvin, proposed that it would be convenient to have a temperature scale in which the zero point was -273.15°C .

This temperature scale is referred to as the Kelvin scale, and the units of the scale are known as Kelvins, K.

The Kelvin degree, is equivalent in size to the Celsius degree. [Every $^{\circ}\text{C} = 1\text{K}$]

When the Kelvin temperature scale is used, the volume-temperature relationship, now known as **Charles's law**, can be stated as: *If a given quantity of gas is held at a constant pressure, its volume is directly proportional to the absolute temperature.*

Mathematically, this can be expressed as:

$$V \propto T \quad [n, P \text{ constant}]$$

$$V = \text{constant} \times T$$

$$\frac{V}{T} = \text{constant}$$

For a given sample of gas at constant pressure, the quotient V/T is always equal to the same constant.

Thus, doubling the temperature, the kelvin temperature will cause the gas volume to double.

($T = \text{Kelvin}$, $2 \times T = 2 \times V$)

Hence, if we know the volume (V_1) and temperature (T_1) of a gas sample, we can determine the new volume (V_2) at the new temperature (T_2) from the relationship:

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

Note: In calculations temperature, T , must be expressed in kelvins.

Sample Problem

A balloon is inflated with helium to a volume of 40.0 L at room temperature, 25.0 °C. If the balloon is inflated with the same quantity of helium on a very cold day — 10.0 °C, what is the new volume of the balloon?

Solution

Original conditions	Final conditions
$V_1 = 40.0 \text{ L}$	$V_2 = ?$
$T_1 = 25.0 \text{ }^\circ\text{C} + 273 = 298 \text{ K}$	$T_2 = -10.0 \text{ }^\circ\text{C} + 273 = 263 \text{ K}$

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(40.0 \text{ L})(263 \text{ K})}{298 \text{ K}} = \text{L}$$

Logically, the volume of the gas will decrease with a decrease in temperature.

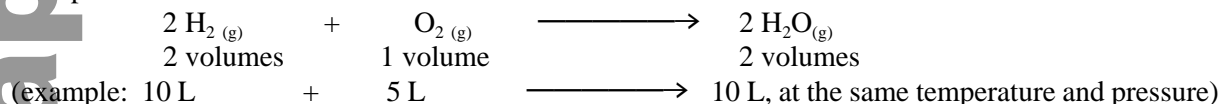
Now Do: Assignment: 1. Volume – Temperature Relationship 2. Gas Law Problems: PV, VT, PT

The Quantity — Volume Relationship: Avogadro's Law

The volume of a gas in a container is affected not only by pressure and temperature but by the amount of gas as well. The relationship between the quantity of a gas and its volume was investigated by Joseph Louis Gay-Lussac (1778 — 1823) and by Amadeo Avogadro.

Joseph Gay-Lussac experimented on the properties of gases and he found *volumes of gases always combine in simple whole number ratios, as long as the volumes are measured at the same temperature and same pressure*. This is now known as the **Law of Combining Volumes**.

For example, two volumes of hydrogen gas react with one volume of oxygen gas to form two volumes of water vapour:



Gay-Lussac's law was based on experimental observation was explained in 1811 by Amadeo Avogadro and is now known as **Avogadro's hypothesis**:

Equal volumes of gases at the same temperature and pressure contain equal number of molecules.

Example: Complete the following chart:

	Hydrogen	Nitrogen	Helium	Oxygen	carbon dioxide
Volume (L)	1	2	5	10	100
Pressure (kPa)	100	100	100	100	100
Temperature (°C)	25	25	25	25	25
Number of mol	1	2	5		
Number of molecules					

Experiments have shown that *1 mol of any gas (i.e. 6.02×10^{23} gas molecules) at 101.3 kPa and 0°C occupies 22.4 L*. This is known as the **Molar Volume** of a gas at **Standard temperature and pressure, STP**.

At, standard temperature and pressure, STP (0°C, 101.3 kPa), the Molar Volume of any gas = 22.4 L

1mole = 6.02×10^{23} molecules = Molar Volume of 6×10^{23} molecules = 22.4 L

Avogadro's Law follows from Avogadro's hypothesis:

The volume of a gas, at a given temperature and pressure, is directly proportional to the quantity of gas.

Hence, volume of a gas, V is proportional to the number of mols of gas, n:

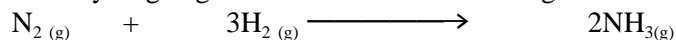
$$\frac{V}{n} = \text{constant}$$

Thus:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Sample Problems

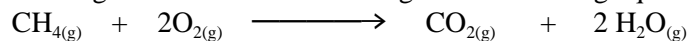
1. Nitrogen and hydrogen gases react to form ammonia gas according to the following equation:



At a certain temperature and pressure, if we begin with 15.0 L of hydrogen gas, and it is all consumed, calculate the volume of nitrogen gas required and the volume of ammonia produced, at the same temperature and pressure.

Solution

2. Methane undergoes combustion according to the following equation:



If 24.0 L of carbon dioxide undergo complete combustion at a given temperature and pressure, determine:

- the volume of oxygen required for complete combustion,
- the volume of carbon dioxide gas produced
- the volume of water vapour produced.

Assume all gases are measured at the same temperature and pressure.

3. Consider a 1.0 L flask containing helium gas and a 2.0 L flask containing nitrogen gas. Both gases are at the same temperature and pressure. According to Avogadro's Law, what can be said about the:

- the number of moles of each gas
- the number of molecules of each gas,
- the mass of each gas in each flask.

Now Do: Assignment: Molar Volume Practice I, Practice II, and Practice III

General Gas Law or Combined Gas Law:

Boyle's Law and Charles's Law can be combined in a general gas equation.

Because generally a change in pressure is accompanied by a change in volume and change in temperature, we use:

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad (T = \text{Kelvin})$$

Used generally when a gas is described under **2 sets of different conditions.**

Now Do: Assignment: Problems on Combined Law: Gas Law Practice IV, and Gas Law Problems V

Section D: IDEAL GAS EQUATION

According to Section C above, we discovered:

$$\text{Boyle's Law: } V \propto \frac{1}{P} \quad (\text{constant } n, T)$$

$$\text{Charles's Law: } V \propto T \quad (\text{constant } n, P)$$

$$\text{Avogadro's Law: } V \propto n \quad (\text{constant } P, T)$$

These relationships can be combined to make a more general gas law:

$$V \propto \frac{nT}{P}$$
$$V = \frac{R(nT)}{P}$$

Rearranging, we have the relationship:

$$PV = nRT$$

This equation is known as the **ideal - gas equation**.

Any gas that obeys this equation is called an **ideal gas**.

(Note that there is no such thing as an ideal-gas, however at low pressures, around 1 atm or lower, and room temperature, gases will behave like an ideal gas and obey $PV = nRT$)

$PV = nRT$ is applied to a single set of conditions

The constant R in the ideal gas equation is called the ideal gas constant, (aka: universal gas constant, Avogadro's constant).

The value of R depends on the units of P , V , n , and T . (Temperature must always be expressed in Kelvin).

The quantity of gas, n , is expressed in mols, P is often expressed in kPa, and volume, V in litres. Therefore units of R can be;

$$R = 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K}, \quad = 8.314 \text{ L} \cdot \text{kPa} / \text{mol} \cdot \text{K} \quad = 8.314 \text{ J/mol} \cdot \text{K}$$

(Note: the product of PV has the units of energy, therefore the units of R can include joules)
(The value for 'R' is always given on the examination paper.)

Now suppose we have 1.00 mol of an ideal gas at 101.3 kPa and 0.00°C ($= 273.15 \text{ K}$), then from the ideal-gas equation, the volume of the gas is $= 22.4 \text{ L}$

The conditions 0°C and 101.3 kPa ($= 1 \text{ atm}$) are referred to as the **standard temperature and pressure (STP)**.

The volume occupied by 1 mol of an ideal gas at STP, 22.4 L is known as the **molar volume** of an ideal gas at STP, (see above).

Some important derived relationships from the ideal-gas equation:

1. Since, we know: $n = \frac{m}{M_R}$ therefore, $PV = \frac{m}{M_R} \cdot R T$

2. For a mixture of gases: $P_T V = n_T R_T$

3. Dry air, with an average molar mass of about 29 g/mol, has a density of about 1.2 g/L (at 1 atm and 25 °C). This means that gases or vapours with molar masses greater than 29 g/mol have densities larger than 1.2 g/L (at 1 atm and 25 °C). Therefore, gases such as CO₂, SO₂, and gasoline vapour settle along the ground if released into the atmosphere. Conversely, gases such as H₂, He, CO, CH₄ (methane), and NH₃ rise if released into the atmosphere.

Since density, $d = \text{mass} / \text{volume}$:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad \text{therefore,} \quad P = \frac{d \cdot R \cdot T}{M_R}$$

Therefore density and/or the molar Mass of an unknown gas may be easily calculated.

Sample Problems

1. A deep breath of air has a volume of 1.05 L at a pressure of 99.0 kPa and body temperature of 37.0 °C. Calculate the number of molecules in the breath. (Answer: 2.42×10^{22} gas molecules)

Solution

2. Fluorine gas is dangerously reactive and is shipped in stainless steel cylinders of 30.0 l capacity at a pressure of 195.5 kPa at 26.0 °C. Determine the mass of fluorine in a cylinder.

3. Cyclopropane, is a gas used with oxygen as a general anaesthetic, and is composed of 85.7 % C, 14.3 % H by mass. If 1.56 g of cyclopropane has a volume of 1.00 L at 99.6 kPa and 50.0 °C, what is the molecular formula of cyclopropane?

Now Do: Assignment: Ideal Gas Law Problems

Section E: The Gas Laws and Chemical Reactions

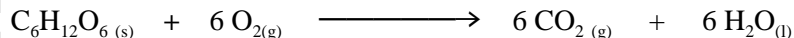
Many important chemical reactions involve gases. Manufactures of fertilisers require large quantities of ammonia gas. The large-scale manufacture of ammonia is carried out using the Haber process. In this process nitrogen gas is reacted with hydrogen gas:



It is important to know quantities involved in chemical reactions, thus principles of stoichiometry are used to determine quantities.

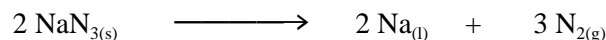
Sample Problem

1. Glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, is broken down by the body, expelling carbon dioxide gas by our lungs:



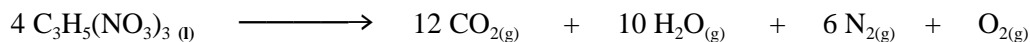
Calculate the volume of carbon dioxide produced at body temperature, 37.0°C and 101.3 kPa when 5.50 g of glucose is consumed.

2. Air bags in cars inflate in the event of an accident (usually in about 0.04 s), they use a fast gas-producing reaction, the decomposition of sodium azide, NaN_3 , according to the following reaction:



What mass of sodium azide, NaN_3 is required to fill an air bag with nitrogen gas which has a volume of 45.5 L at a temperature of 22.0°C and a pressure of 110.5 kPa ?

3. All explosives use the same principle of turning a small volume of liquid or solid into a very large volume of gas in a short time period. For example, the explosive component of dynamite is nitroglycerine, $\text{C}_3\text{H}_5(\text{NO}_3)_3$. It decomposes according to the following equation:



Calculate the pressure produced in a 1.0 L steel container by the explosion of 10.0 g of nitroglycerine at an internal temperature of 50.0°C

Now Do: Assignment: Gas Stoichiometry

Section F: Dalton's Law of Partial Pressures

We have only so far discussed the behaviour of pure gases, however normally we have mixture of gases. For example: air is a mixture of gases, oxygen, nitrogen, carbon dioxide, water vapour and small amounts of other gases.

John Dalton was the first to observe:

The total pressure of a mixture of gases equals the sum of the partial pressure of each gas in the mixture.

$$P_{\text{TOTAL}} = P_1 + P_2 + P_3 + \dots \quad (P_1, P_2) = \text{Partial pressure of each gas.}$$

The partial pressure of a gas is the pressure that the gas would exert if it were the only gas in the container.

(There is so much space between the molecules in a gas that the molecules of another gas can readily share the space). Therefore each gas behaves independently of the other, and makes its own contribution to the total pressure.

Partial pressure of a gas is independent of the type of the gas itself, BUT is dependent on the number of moles of gas particles present.

$$\frac{\text{mols of gas 1, } n_1}{\text{Total mols, } n_T} = \frac{\text{partial press. of gas 1, } P_1}{\text{Total pressure, } P_T}$$

therefore:

$$\text{Partial pressure of gas 1, } P_1 = \frac{(\text{mols gas 1, } n_1) \cdot P_T}{\text{Total mols, } n_T}$$

$$\text{Partial pressure of gas 1, } P_1 = (\text{Mole fraction of gas 1}) \cdot P_T$$

$$P_1 = \frac{n_1}{n_T} \cdot P_T$$

Also we can use the ideal gas equation:

$$P_T V = n_T \cdot R T$$

P_T = sum of partial pressures of the gaseous mixture.

n_T = total number of mols of all gases present in the mixture.

Sample Problem

1. A mixture consists of 0.500 mol of $\text{H}_{2(g)}$, 1.00 mol $\text{CO}_{2(g)}$, 0.500 mol $\text{Ar}_{(g)}$. If the total pressure exerted by the mixture is 200 kPa, determine the partial pressure of exerted by each gas.
2. What is the partial pressure of each gas and the total pressure exerted by a gaseous mixture of 6.00 g O_2 , 9.00 g $\text{CH}_4_{(g)}$?
3. What is the total pressure exerted by a mixture of 2.00 g of H_2 and 28.0 g of N_2 at 298 K in a 10.0 L container?

Now Do: Assignment: Dalton's Law of Partial Pressures

Dalton's Law and Collecting Gases over Water

Section G: GRAHAM'S LAW OF DIFFUSION

Diffusion is the process whereby a substance spreads from a region of high concentration to one of lower concentration.

This is to say, it is the migration of intermingling of molecules as a result of random molecular motion.

The rate at which diffusion occurs depends directly on the molecular speed i.e. Molecules with high speed diffuse faster than those with low speeds.

Graham's Law:

The rates of diffusion of two gases are inversely proportional to the square root of their Molar Mass. i.e.

$$\frac{\text{Rate of diffusion of A}}{\text{Rate of diffusion of B}} = \frac{\sqrt{\text{Molar mass of B}}}{\sqrt{\text{Molar mass of A}}}$$

Therefore the molecules with smaller mass have a higher speed, and hence the lighter gas diffuses faster than a heavier gas, and therefore travels the farthest. Conversely, the gas that diffuses the fastest takes the shortest time and is the lightest.

Application of Graham's Law in the determination of the Molar Mass of an unknown gas.

Example: Rate of diffusion of an unknown gas (x) is four times faster than oxygen. Calculate the MM of the gas x.

$$\frac{\text{Rate of x}}{\text{Rate of O}_2} = \frac{\sqrt{\text{MM x of O}_2}}{\sqrt{\text{MM of x}}}$$

$$\text{Is therefore: } \frac{4}{1} = \frac{\sqrt{32}}{\sqrt{\text{MMx}}}$$

$$\text{Which is therefore: } \frac{16}{1} = \frac{32}{\text{MMx}} \quad \text{the MM of x} = 2\text{g/mol}$$

Thus, the gas is identified as H₂.

Section H: Kinetic Molecular Theory

The ideal -gas equation demonstrates how gases behave, however it does not explain why they behave as they do.

The kinetic molecular theory, (developed over a period of over 100 years, culminating in 1857 finally by the publication by Rudolf Clausius, 1822-1888), describes the behaviour of gases at the molecular or the atomic level.

The kinetic molecular theory states:

1. Gas molecules are in constant, random and rapid motion.
2. The volume of all the molecules of the gas is negligible compared to the total volume in which the gas is contained.
3. Attractive and repulsive forces between gas molecules are negligible.
4. Gas molecules collide with one another and with the walls of the container without loss of any energy. In other words, the collisions are perfectly elastic.
5. The average kinetic energy of the gas molecules is determined by the gas temperature. All molecules regardless of their size, have the same average kinetic energy at the same temperature.

The kinetic molecular theory explains the concept of pressure and temperature at the molecular level.

The pressure of a gas is caused by:

The magnitude of the pressure depends on how often and how hard the gas molecules hit the wall of their container.

The absolute temperature of a gas is a measure of the average kinetic energy of the gas molecules.

Section I: The Ideal Gas

An ideal gas so far has been considered as a collection of rebounding particles, separated by great distances, so that the gas is mostly empty space, and there are no forces of attraction between the molecules.

The above gas laws and the ideal gas equation were used to predict gas behaviour.

They are Natural or IDEAL gas laws.

However, we find that **no gas is truly “ideal”, that there is indeed attraction between the molecules and the magnitude of these intermolecular forces cause deviation from the ideal gas laws and finally to the formation of liquids.**

Intermolecular forces: are forces of attraction between molecules as opposed to intramolecular forces of attraction, these are forces of attraction within a molecule.

Further, the volume occupied by gas molecules can not be negligible at high pressures.

Now Do: Assignment:

1. **Kinetic Molecular Theory Review**
2. **Gas Problems Practice**
3. **Review Problems I , II**