Chemistry Year 1

SECTION 5

KINETIC THEORY AND THE STATES OF MATTER

PHYSICAL CHEMISTRY

Matter and its Properties

INTRODUCTION

Hello learner! Welcome to the kinetic theory of matter. This fundamental concept in chemistry explains how different states of matter behave. It's based on the idea that matter is made up of tiny particles—molecules, atoms, or ions—that are always in motion. In this section you will also focus on understanding and applying kinetic theory, which explains how particles behave in gases. You will also explore how gases are prepared in the laboratory and their practical uses in everyday life.

At the end of this section, you will be able to:

- Explain the kinetic theory of matter and apply it to distinguish between the properties of solids, liquids and gases.
- State and perform calculations involving various Gas Laws and analyse graphs based on the laws.
- State Graham's Law of Diffusion/effusion and Dalton's Law of partial pressure and apply them to perform calculations.
- Write the ideal gas equation and apply it in simple calculations using the different numerical values of R and units of Pressure and Volume.
- Explain why gases show deviation from ideal behaviour and suggest how the ideal gas equation could be modified to describe gas behaviour more accurately.
- Design and perform experiments to prepare and test for gases (hydrogen, ammonia and carbon dioxide gases).

Key Ideas

- **Particles** are small constituent units of matter, which can be atoms, molecules, or ions.
- **Intermolecular forces** are forces of attraction or repulsion between particles.
- **Diffusion** is the process of particles spreading out from an area of higher concentration to an area of lower concentration.
- **Brownian motion** is the random movement of particles suspended in a fluid (liquid or gas) resulting from collisions with fast-moving molecules of the fluid.
- **Gas Law** is a mathematical relationship between the pressure, volume, temperature, and quantity of a gas.
- **Mole fraction** is the ratio of the number of moles of a component to the total number of moles in a mixture.
- **Standard temperature and pressure** are defined as 0°C (273.15 K) and 1 atm pressure.
- **Ideal gas equation** is the equation describing the relationship between Pressure, Volume, Temperature, and the number of moles of a gas: $PV = nRT$
- **Universal Gas Constant (R)** is a molar physical quantity.
- **Displacement of water** this method involves collecting a gas by displacing water in a container.
- **Displacement of Air** (Upward or Downward Delivery) gases are collected by displacing air, based on their density relative to air.
- **Using gas syringe**, a gas syringe is used to directly measure and collect a known volume of gas.

KINETIC THEORY OF MATTER

The kinetic theory of matter states that;

1. Matter is made of tiny particles which are in constant random motion.
2. Matter possesses kinetic energy due to the motion of the particles.

3. The difference between the different states of matter is due to the nature and extent of motion and the separation between the particles.

Solid State

1. Solids have a fixed shape and volume at a given temperature.
2. Particles of solids are closely packed in an orderly manner.
3. Solids have the greatest forces of attraction between their particles when compared to the energy possessed by the particles.
4. Particles of solids undergo vibration about their mean position.
5. Increasing the temperature of solids causes faster vibration of particles.

Using the kinetic model to explain the properties of solids

a. Solids tend to have the greatest density because the particles are usually closest together.
b. Solids have a fixed shape and volume because of the strong force of attraction between the
c. Particles hold the particles in their fixed positions.
d. Solids are difficult to compress because of the lack of empty space between the particles.
e. Solids expand on heating, due to the increased energy of the particles at higher temperatures. This allows the particles to exist at greater distances from each other because the intermolecular forces are constant.

The Liquid State

1. Liquids have fixed volumes that take the shape of the container at a given temperature.
2. Particles of liquids are close together and arranged randomly.
3. The particles of liquids move rapidly in all directions.
4. The forces of attraction between the particles are stronger than that of gases, but lower than that of solids when compared to the energy possessed by the particles.
5. Increasing the temperature of liquids makes their particles move faster due to a gain in kinetic energy.

Using the kinetic model to explain the properties of liquids

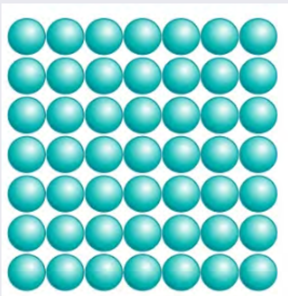
a. Liquids have greater densities than gases because the particles are closer due to the attractive forces.

b. Liquids have fixed volume but take the shape of their container because of the increased particle attraction.

c. Liquids are not easily compressed because there is so little empty space between the particles.

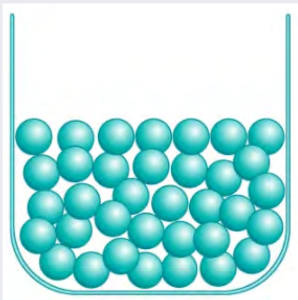
d. Liquids expand on heating, due to the increased energy of the particles at higher temperatures. This allows the particles to exist at greater distances from each other because the intermolecular forces are constant.

Activity 4.1: Representations to depict the arrangement and motion of particles in solids and liquids.



Solid (ice)

Particles vibrating in place but not changing positions



Liquid (water)

Particles moving and sliding past each other.

Use the links below to watch the animation of particles of solid and liquid:

https://phet.colorado.edu/sims/html/states-of-matter-basics/latest/states-of-matter-basics_en.html or

https://www.acs.org/content/acs/en/education/resources/k-8/inquiryinaction/fifth-grade/particles-solid-hammer.html or

https://www.youtube.com/watch?v=gPMVaAnij88

Activity 4.2: Investigating Properties of Solids and Liquids

Experiment 1: Observing the Shape and Volume of Solids and Liquids

Materials needed:

- Transparent containers of different shapes and sizes
- A solid object (e.g. A rubber ball)
- Different liquids (e.g. Water, oil, syrup).

Procedure:

1. Place the solid object in different containers and observe if its volume and shape change.
2. Pour the liquids into different containers and observe how they take the shape of each container.
3. Measure the volume of the liquids before and after transferring to different containers to see if the volume changes.

Presentation

	Solid	Liquid
Shape		
Volume		

4. Explain the differences between the shape and volume of solids and liquids.

Extension task

Consider the material toothpaste, is it a solid or a liquid? Prepare a balanced argument which outlines the reasoning for it being a solid **and** a liquid.

Experiment 2: Compare the density of various solids and liquids.

Materials needed:

- Different solid objects (e.g., metal cubes, wooden blocks, plastic objects),
- Various liquids (e.g., water, oil, syrup)
- Graduated cylinders
- Electronic balance
- Measuring cups

Procedure:

1. Measure the mass of each solid object using the electronic balance.
2. Measure the volume of each solid by water displacement in a graduated cylinder.
3. Calculate the density of each solid using the formula:

   $$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

4. Measure the volume of each liquid using measuring cups.
5. Measure the mass of each liquid by pouring it into the graduated cylinder and weighing it.
6. Calculate the density of each liquid using the same formula.

Presentation

	Solid			Liquid		
	Solid 1	Solid 2	Solid 3	Liquid 1	Liquid 2	Liquid 3
Mass						
Volume						
Density						

7. Compare the densities of solids and liquids.

Extension task

If the density of water is 1000 kgm^{-3}, which of these objects (both solid and liquid) will float or sink?

Experiment 3: Behaviour of particles in solids

Materials needed: Marbles, beads, trays and spoons

Procedure

1. Place marbles or beads tightly packed into a tray.
2. Gently shake the tray to show that the marbles/beads vibrate but do not change their positions significantly.
3. Use a spoon to apply gentle pressure and rearrange the marbles/beads, showing that while their positions can be changed by an external force, they still maintain a fixed arrangement overall.
4. This simulates how particles in a solid can be rearranged by external forces but generally stay in a fixed structure.

Experiment 4: Comparing the compressibility of various solids and liquids.

Materials needed: Flexible plastic bottle with cap, water and beads

Procedure

1. Fill the bottle with water to the very top and put the cap on securely.
2. Squeeze the bottle.
3. Hold the bead between two fingers and squeeze.
4. Were you able to squeeze the bottle filled with water or the bead?

Extension task

When a solid sponge is compressed its volume changes drastically. Why does this not constitute a solid being compressed?

Safety: Wear safety glasses or goggles and be sure to follow all safety instructions given by your teacher. Wash your hands after completing the activity.

Experiment 5: Comparing the Viscosity of Liquids

Materials needed: Different liquids (e.g., water, oil, honey, syrup), stopwatch, inclined planes, measuring cups.

Procedure:

1. Pour equal amounts of each liquid down the inclined plane.
2. Measure how long it takes for each liquid to reach the bottom of the inclined plane.
3. Rank the liquids based on their flow times to determine their viscosity.

Type of liquid	Time taken for liquid to flow
Water	
Oil	
Syrup	
Honey	

4. Discuss why some liquids might flow faster than others.

Extension task

Consider why viscosity is an important property for an engineer to understand about a liquid.

- What applications could there be for liquids with very low viscosities?
- What applications could there be for liquids with very high viscosities?
- What issues might arise from handling such liquids (high or low viscosity) on an industrial scale? Think about any issues that might arise from pumping, pouring or mixing.

THE GASEOUS STATE

1. Gases have no fixed shape or volume but fill a container.
2. Forces of attraction between the particles of gases are negligible.
3. Particles are so small that the actual volume of individual particles is negligible compared to the volume of the container.
4. Particles are widely spaced and scattered and undergo random and rapid motion.
5. The average kinetic energy of the gas particles is directly proportional to the absolute temperature of the particles.
6. The collision of the gas particles with the surface of the container causes gas pressure.

Using the kinetic model to explain the properties of gases

a. Gases have very low densities because the particles will space out in the container.

b. Gases have no fixed volume and shape because of the negligible force of attraction.

c. Gases are easily compressed because of the space between the particles

d. Order of ease of compression: gas > liquid > solid

e. Gases exert pressure because of the collision of the particles with the walls of the container.

Activity 4.3: Gas Behavior Under Different Conditions

Materials needed: software - PhET Interactive Simulations (or similar) computer and internet access

Use the link below to watch the simulation of gases:

https://phet.colorado.edu/sims/html/states-of-matter-basics/latest/states-of-matter-basics_en.html

Procedure:

Temperature and pressure effects

1. Increase and decrease the temperature of the gas in the simulation.
2. Observe how changing temperature affects the speed and kinetic energy of gas particles.
3. Adjust the pressure settings in the simulation.
4. Observe how changes in pressure affect the volume and density of the gas.
5. Discuss the relationship between pressure and the frequency of particle collisions.
6. Discuss how temperature impacts the volume and pressure of the gas.
7. Record your observations and findings from each task.

Volume Effects

1. Change the volume of the gas chamber in the simulation.
2. Observe how altering the volume affects the pressure and density of the gas.
3. Record your observations and findings from each task.

Extension task

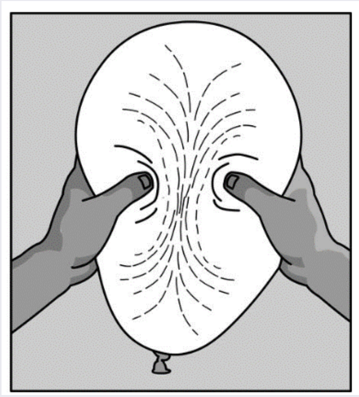
Push the model to its limits. Describe and then explain what you see, how would this look in a real-world scenario?

Activity 4.4: Comparing space, shape and compressibility of gasses

Material needed: Empty balloon

Procedure:

1. Fill the entire space inside the balloon.
2. Discuss what you observed.
3. Squeeze the balloon between.



4. Discuss what you observed.

Extension task

How does the compressibility of a gas change as it is being compressed? Use your knowledge of particles to answer this question.

EXPLANATION OF CHANGE OF STATE PROCESSES

Melting

When solids are heated, particles gain kinetic energy and vibrate more strongly. Attractive forces weakening in comparison to particle energy, particles become free to move around.

Freezing

When liquids are cooled, particles lose kinetic energy. Attractive forces strengthen in comparison to particle energy, causing particles to be restricted to a fixed position they still move but this is now a vibration.

Evaporation and Boiling

1. Evaporation occurs when the particles of a liquid escape to form a vapour.
2. On heating, surface particles gain more kinetic energy, move faster, and break away from intermolecular forces.
3. Boiling is rapid vaporisation anywhere in the bulk liquid at a fixed temperature (the boiling point of the liquid).

Condensation

It is the change in the physical state of matter from the gaseous form into the liquid form. This occurs when the temperature of the gas is lowered to the point where its particles lose enough kinetic energy to form bonds and transition into a liquid. Condensation is the reverse process of vaporisation.

Melting Point Determination

1. Heat one end of the capillary tube in a Bunsen flame.
2. Fill about one-quarter of the capillary tube with the substance and tie it to a thermometer.
3. Insert the tube and thermometer into an oil bath.
4. Heat with constant stirring and record the temperature at which the first crystal melts and the
5. temperature at which the last crystal melts

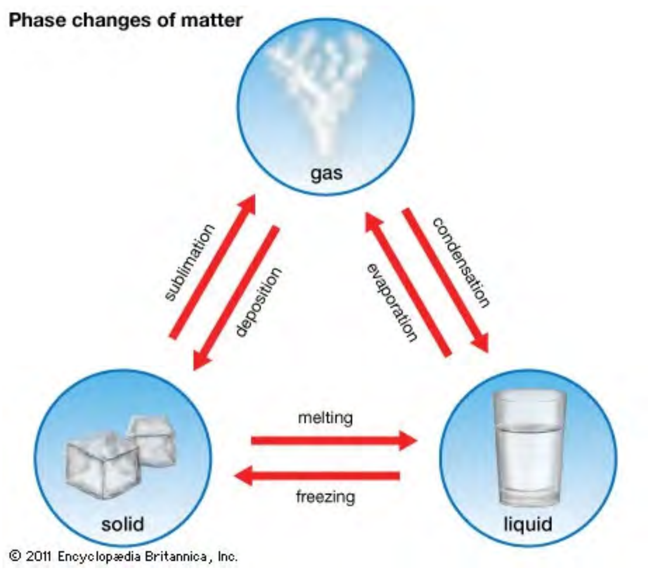


Figure 4.1: Phase changes of matter

Image source: https://www.britannica.com/science/phase-state-of-matter#/media/1/455270/155241

Activity 4.5: Demonstrating melting and boiling processes using ice cubes and a heat source

Materials Needed: Ice cubes, Bunsen burner or hot plate, or electric kettle, thermometers, beakers, stopwatch, safety goggles and gloves.

Melting of ice

1. Place a few ice cubes in a beaker and record its initial temperature.
2. Allow the ice to melt at room temperature first, observing the process and recording the temperature change every minute.
3. Use a heat source to speed up melting after initial observations. (be careful if using a direct heat source such as a Bunsen as the glass may shatter)
4. Note the time it takes for the ice to melt completely with and without the heat source.

Melting Observation

The initial temperature of the ice cube	
Time taken for ice to melt at room temperature	
Time taken for ice to melt with a heat source	
The final temperature of melted water	

Boiling of water

1. Pour the melted water into a beaker and place it on the heat source.
2. Measure and record the temperature of the water every minute until it starts boiling.
3. Observe the boiling process and note the boiling point of water (100°C).

Boiling Observation

The initial temperature of the water	
Time taken to reach boiling point	
Boiling point observed	
Observations	

Extension task

How could this experiment be improved to be more accurate (closer to the value of 100°C)?

Activity 4.6: Demonstrating change of state using simulations of the water cycle or industrial process like distillation.

1. The simulation allows users to explore how water molecules move and change state under different temperatures and pressures. Use the link below to observe the animation.

   https://phet.colorado.edu/sims/html/states-of-matter-basics/latest/states-of-matter-basics_en.html

2. The video provides an engaging animation of the water cycle, demonstrating processes like evaporation, condensation, and precipitation. Use the link below to observe the animation.

   https://www.youtube.com/watch?v=ncORPosDrjI

GAS LAWS

Hello learner! In this lesson, you will learn about the gas laws, which help predict how gases behave under different conditions. You will understand how pressure, volume, and temperature affect gases and relate the amount of gas to these factors. These laws have many practical applications in Engineering, Chemistry, Physics, and more.

Boyle's Law

It states that the volume of a fixed mass of gas at constant temperature is inversely proportional to the pressure of that gas.

Mathematically Boyle's law is expressed as,

$P_1 V_1 = P_2 V_2$

Where P_1 and P_2 are initial and final pressures respectively

And V_1 and V_2 are initial and final volumes respectively.

Graphically Boyle's law is represented as follows;

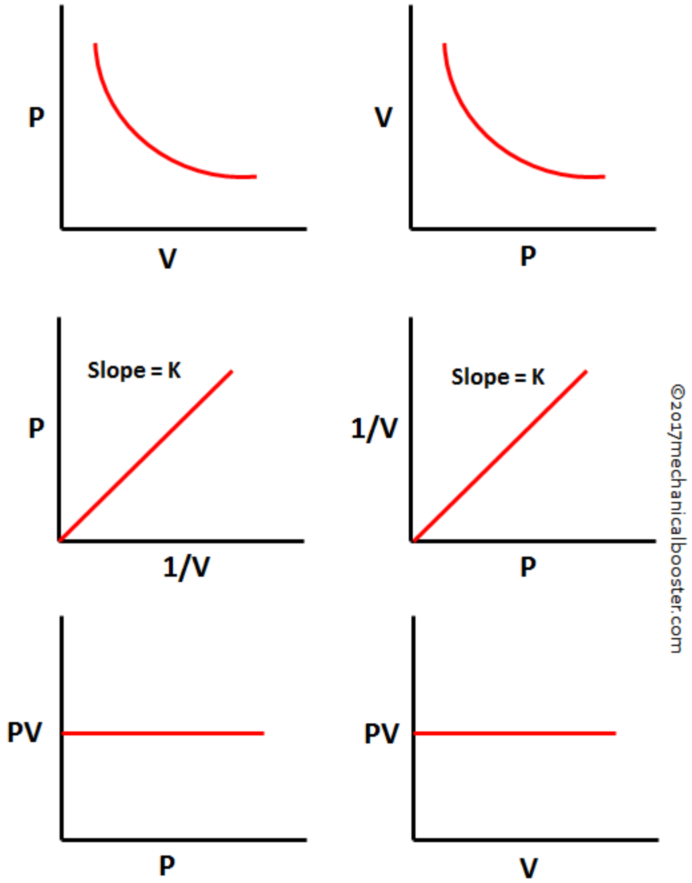


Figure 4.2: Graphs Showing Different Representations of Boyle's Law.

Example 4.1

10 m^3 volume of a gas at a pressure of 101,300 Pa was compressed to a volume of 6 m^3 at constant temperature, calculate the final pressure.

Answer: (Use problem solving strategy)

a. Analyse the question

Known

Initial volume, $V_1 = 10\ m^3$

Initial pressure, $P_1 = 101{,}300$ Pa

Final volume, $V_2 = 6\ m^3$

Unknown

Final pressure, $P_2 =?$

b. Apply the problem-solving strategy

$$P_1V_1 = P_2V_2$$

$$P_2 = \frac{P_1V_1}{V_2}$$

$$= \frac{101{,}300 \times 10}{6}$$

$$= 168{,}833\ Pa$$

Charles' Law

It states that the volume of a fixed mass of a gas at constant pressure is directly proportional to its absolute temperature.

Mathematically Charles' law is represented as,

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

Where V_1 and V_2 are initial and final volumes respectively, and T_1 and T_2 are initial and final temperatures respectively.

NB: The temperature must ***always*** *be converted to Kelvin.*

Graphically Charle's law is represented as follows;

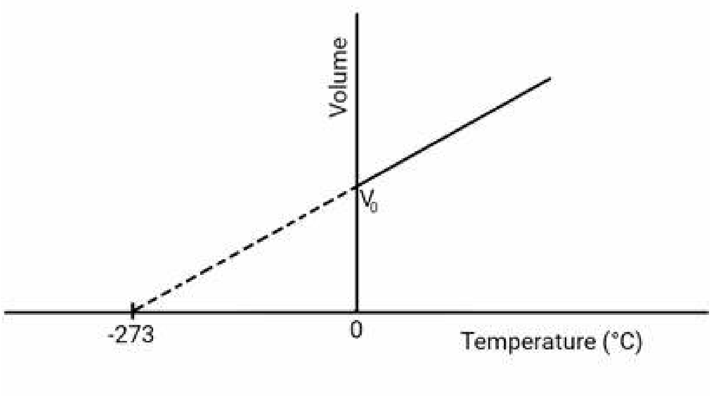


Figure 4.3: Graphical Representation of Charles Law.

Example 4.2

10 m^3 volume of a gas in a cylinder is heated from 250 K to 300 K at constant pressure, calculate the final volume of the gas in the cylinder.

Answer (Use problem solving strategy)

a. Analyse the question

Known

Initial volume, $V_1 = 10\ m^3$

Initial temperature, $T_1 = 250$ K

Final temperature, $T_2 = 300$ K

Unknown

Final volume, $V_2 =?$

b. Apply the problem-solving strategy

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

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

$$= \frac{10 \times 300}{250}$$

$$= 12\ m^3$$

Gay-Lussac's Law

It states that for a fixed mass of gas at constant volume, the pressure of that gas is directly proportional to its absolute temperature (K).

Mathematically Gay-Lussac's law is represented as,

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Example 4.3

A fuel and air mixture in a car engine cylinder of volume 1000 cm^3 increases from 20 ^{0}C to 2000 ^{0}C upon combustion. If the normal atmospheric pressure is 100 kPa, calculate the final pressure.

Answer: (Use problem-solving strategy)

a. Analyse the question

Known

Initial temperature, $T_1 = 20 + 273 = 293$ K

Final temperature, $T_2 = 2000 + 273 = 2{,}373$ K

Initial pressure, $P_1 = 100$ kPa

Unknown

Final pressure, P_2 =?

b. Apply the problem-solving strategy

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$P_2 = \frac{100 \times 2273}{293} = 775.8\ kPa$$

Evaluate: Check the answer to see if it makes sense.

Combined gas Law

It is the combination of the Boyle's law, Charle's law and the Gay-Lussac's law. Mathematically it is expressed as

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Example 4.4

20 cm^3 of a gas at 1 atm and 25 ^{0}C was compressed to 16 cm^3 at 40 ^{0}C, calculate the final pressure of the gas.

Answer: (Use problem-solving strategy)

a. Analyse the question

Known

Initial pressure, P_1 = 1 atm

Initial volume, V_1 = 20 cm^3

Final volume, V_2 = 16 cm^3

Initial temperature, T_1 = 25 + 273 = 298 K

Final temperature, T_2 = 40 + 273 = 313 K

Unknown

Final pressure, P_2 =?

b. Apply the problem-solving strategy

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_2 = \frac{P_1V_1T_2}{T_1V_2}$$

$$= \frac{1 \times 20 \times 313}{298 \times 16} = 1.31\ \text{atm}$$

It is worth noting in all of these mathematical laws that the units that are chosen for each quantity need to be consistent. If the volume is initially measured in m^3, then it should be in m^3 for the final volume. If the volume is, however, in cm^3 initially and in m^3 for the final volume, then one of them will need to be converted to the other.

Avogadro's Law

It states that equal volumes of gases at the same temperature and pressure contain the same number of molecules or moles of gas.

Mathematically Avogadro's law is expressed as,

$$\frac{V}{n} = K$$

Where V = Volume occupied by the gas

n = number of moles of the gas

K = constant of proportionality

Example 4.5

Consider the reaction,

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

If the rate of production of CO in an industrial plant is at 20 dm^3 per minute, what is the rate of production of oxygen to produce CO_2?

Answer

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

$$\frac{V(CO)}{V(O_2)} = \frac{2}{1}$$

$$V(O_2) = \frac{V(CO)}{2}$$

$$= \frac{20}{2} = 10\ dm^3\,min^{-1}$$

Activity 4.7:Using simulations to visualise the relationships between variables in gas laws.

1. Boyle's Law

Use the link below to observe the simulations.

https://phet.colorado.edu/sims/html/gas-properties/latest/gas-properties_en.html

Steps:

1. Open the PhET "Gas Properties" simulation and select the "Boyle's Law" experiment setup.
2. Adjust the volume of the container and observe the changes in pressure while keeping the temperature constant.
3. Take note of pressure and volume readings at different volumes.
4. Plot a graph of Pressure (P) on the y-axis and Volume (V) on the x-axis.

Note: The plot should show that as volume decreases, pressure increases, demonstrating the inverse relationship between P and V.

2. Charles's Law

Steps:

1. Open the PhET "Gas Properties" simulation and select the "Charles's Law" experiment setup.
2. Adjust the temperature while keeping the pressure constant and observe the changes in volume.
3. Take note of volume and temperature readings at different temperatures.
4. Plot a graph of Volume (V) on the y-axis and Temperature (T) on the x-axis (in Kelvin).

Note: The plot should show that as temperature increases, volume increases, demonstrating the direct relationship between V and T.

3. Avogadro's Law

Steps:

1. Open the PhET "Gas Properties" simulation and select the "Avogadro's Law" experiment setup.

2. Add or remove gas particles (moles of gas) while keeping the temperature and pressure constant and observe the changes in volume.
3. Take note of volume and number of moles readings at different amounts of gas.
4. Plot a graph of Volume (V) on the y-axis and Number of Moles (n) on the x-axis.

Note: The plot should show that as the number of moles increases, volume increases, demonstrating the direct relationship between V and n.

4. Gay-Lussac's Law

Steps:

1. Open the PhET "Gas Properties" simulation and select the "Gay-Lussac's Law" experiment setup.
2. Adjust the temperature while keeping the volume constant and observe the changes in pressure.
3. Plot a graph of Pressure (P) on the y-axis and Temperature (T) on the x-axis (in Kelvin).

Note: The plot should show that as temperature increases, pressure increases, demonstrating the direct relationship between P and T. Take note of pressure and temperature readings at different temperatures.

DIFFUSION

Hello learner! Get ready to explore Graham's law of diffusion, which explains how gases spread, and Dalton's law of partial pressures, which shows how gases behave in a mixture. Exciting insights into the world of gases await you!

What is Diffusion?

It is the random motion of molecules by which there is a net flow of matter from a region of high concentration to a region of low concentration. A familiar example is the perfume of a flower that quickly permeates the still air of a room.

Graham's law of diffusion/effusion

It states that the rate of diffusion, or effusion, for a gas is inversely proportional to the square root of its density at constant temperature and pressure.

Mathematically, $R \propto \frac{1}{\sqrt{D}}$ or $R \propto \frac{1}{\sqrt{M}}$

R = Rate of diffusion or effusion

D = Density of the gas

M = Molecular mass of the gas

Diffusion involves the movement of molecules from an area of high concentration to an area of low concentration due to random motion of molecules. It occurs in gases, liquids and solids and does not require a boundary. On the other hand, effusion specifically refers to the escape of gaseous molecules through a tiny hole into a vacuum or region of lower pressure. It is a type of diffusion, but it is specifically about gaseous molecules moving through a tiny opening.

For 2 gases diffusing at the same time:

$$\frac{R_1}{R_2} = \sqrt{\frac{D_2}{D_1}} \quad \text{and} \quad \frac{R_1}{R_2} = \sqrt{\frac{M_2}{M_1}}$$

Where R_1 and R_2 are rates of diffusion of gases 1 and 2.
Where D_1 and D_2 are densities of gases 1 and 2.
Where M_1 and M_2 are molecular masses of gases 1 and 2
Graham's law of diffusion can also be expressed in terms of time.

For two gases,

$$\frac{t_1}{t_2} = \sqrt{\frac{M_2}{M_1}}$$

Where t_1 and t_2 are times for the diffusion of gases 1 and 2
Graham's law of diffusion can also be expressed in terms of volume. For two gases,

$$\frac{V_1}{V_2} = \sqrt{\frac{M_2}{M_1}}$$

Where V_1 and V_2 are volumes for gases 1 and 2

Activity 4.8: Experiment to Demonstrate Graham's law of Diffusion

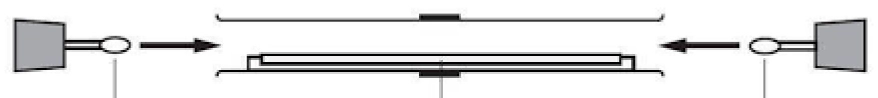


Figure 4.4: A diagram Showing Graham's Law of Diffusion.

Procedure

1. Soak a piece of cotton wool in concentrated NH_3 solution and soak another piece of cotton wool in concentrated HCl.
2. Put at opposite ends of a dry glass tube.
3. Ensure that the glass tube is horizontally mounted to ensure that the diffusion of the gas is not under the influence of gravity.
4. After a few minutes, a white cloud of ammonium chloride appears. This shows the position at which the two gases react.
5. The white cloud of NH_4Cl forms at a position closer to the cotton wool soaked in HCl because particles of NH_3 are lighter than that of HCl and so move faster and where they meet react to give that white fume.

Example 4.6

If 100 cm^3 of methane (CH_4) gas diffuses through a membrane in 40 s, what time will it take 150 cm^3 of ammonia (NH_3) gas to diffuse through the same membrane? [C = 12, H = 1, N = 14]

Answer: (Use problem-solving strategy)

a. Analyse the question

Known

[C = 12, H = 1, N = 14]

Time, t (CH_4) = 40 s

Volume, V(CH_4) = 100 cm^3

Volume, V(NH_3) = 150 cm^3

Unknown

Time, t (NH_3) =?

b. Apply the problem-solving strategy

$M_r(NH_3) = 14 + 3(1) = 17$

$M_r(CH_4) = 12 + 4(1) = 16$

$R_{(NH_4)} = V-$

$= \frac{100}{40} = 2.5\ cm^3 s^{-1}$

$$\frac{R_{(NH_3)}}{R_{(CH_4)}} = \sqrt{\frac{M_{(CH_4)}}{M_{(NH_3)}}}$$

$$\frac{R_{(NH_3)}}{2.5} = \sqrt{\frac{16}{17}}$$

$$R_{(NH_3)} = 0.9696 \times 2.5$$

$$= 2.4\ cm^3 s^{-1}$$

$$t_{(NH_3)} = 1\frac{50}{2.4}$$

$$= 62.5\ s$$

Activity 4.9: Discussing Graham's Law of Diffusion/Effusion

Materials needed: Whiteboard, markers, calculators, sample problems for practice, gas diffusion simulation software or videos.

Group discussion

1. Discuss the meaning of diffusion and effusion.
2. Explain how Graham's Law mathematically relates the rate of diffusion/effusion to the molar mass of gases.

   Mathematical Expression:

   $$\frac{R_1}{R_2} = \sqrt{\frac{D_2}{D_1}} \text{ and}$$

   $$\frac{R_1}{R_2} = \sqrt{\frac{M_2}{M_1}}$$

3. Discuss real-life examples where diffusion or effusion occurs (e.g., the smell of perfume spreading in a room, gases escaping from a balloon).
4. Watch a simulation demonstrating gas diffusion/effusion

   https://phet.colorado.edu/sims/html/diffusion/latest/diffusion_all.html

Extension task

If gases diffuse from high concentrations to low concentrations why do the gases in the Earth's atmosphere now diffuse into space?

Activity 4.10: Effect of Relative Molecular Mass on the Rate of Diffusion/ Effusion

1. Lighter gases or gases with lower molecular mass diffuse or effuse more rapidly. For example, Hydrogen (H_2) with a molar mass of 2 g/mol diffuses faster than Oxygen (O_2) with a molar mass of 32 g/mol.
2. Heavier gases or gases with higher molecular mass diffuse or effuse more slowly.

   This relationship arises because lighter gas molecules move faster at a given temperature than heavier molecules due to their lower mass.

Illustration

Consider two gases, Helium (He) with a molar mass of 4 g/mol, and Argon (Ar) with a molar mass of 40 g/mol. Using Graham's Law:

$$\frac{\text{Rate of diffusion of He}}{\text{Rate of diffusion of Ar}} = \sqrt{\frac{M_{Ar}}{M_{He}}}$$

$$\frac{\text{Rate of diffusion of He}}{\text{Rate of diffusion of Ar}} = \sqrt{\frac{40}{4}} = \sqrt{10} = 3.16$$

This means helium diffuses approximately 3.16 times faster than argon.

Activity 4.11: Investigating the Rate of Diffusion of Ammonia and Hydrogen Chloride Gas

Materials needed: Cotton wool, ammonia solution (NH_3), hydrochloric acid (HCl) solution, two droppers, long glass tube, rubber stoppers or corks, measuring tape or ruler, stopwatch, safety goggles and gloves

Safety Precautions:

- Conduct the experiment in a well-ventilated area or under a fume hood.
- Wear safety goggles and gloves to protect against chemical splashes.
- Ammonia and hydrochloric acid are corrosive and can cause irritation. Handle with care.

Setting up the apparatus

1. Use a long glass tube for the experiment
2. Insert a piece of cotton wool soaked in ammonia solution at one end of the tube.
3. Insert a piece of cotton wool soaked in hydrochloric acid solution at the other end of the tube.
4. Seal both ends of the tube with rubber stoppers or corks immediately after placing the cotton wool.
5. Observe the formation of a white ring inside the tube, which is the product of the reaction between NH_3 and HCl, forming ammonium chloride (NH_4Cl).
6. Use a stopwatch to time how long it takes for the white ring to form after sealing the tube.
7. Measure the distance from each end of the tube to the white ring.
8. Record the distances and time.

Distances travelled by NH_3	
Distances travelled by HCl	
time taken for the white ring to form	

Calculation

$$\text{Rate of diffusion of } NH_3 = \frac{\text{Distances travelled by } NH_3}{\text{time}}$$

$$\text{Rate of diffusion of HCl} = \frac{\text{Distances travelled by HCl}}{\text{time}}$$

Extension task

Use your periodic table to determine the heaviest and lightest gaseous elements, then determine their relative rates of diffusion.

Example 4.8

Hydrogen gas (H_2) and Nitrogen gas (N_2) are allowed to effuse through a small hole. Given that the molar mass of hydrogen is 2 g/mol and the molar mass of nitrogen is 28 g/mol, calculate the rate of effusion of Hydrogen compared to Nitrogen.

Answer

$M_{H_2} = 2 \text{ g mol}^{-1}$, $M_{N_2} = 28 \text{ g mol}^{-1}$

Write the formula for Graham's Law:

$$\frac{\text{Rate of diffusion of } H_2}{\text{Rate of diffusion of } N_2} = \sqrt{\frac{M_{N_2}}{M_{H_2}}}$$

Substitute the given values and calculate the ratio:

$$\frac{\text{Rate of diffusion of } H_2}{\text{Rate of diffusion of } N_2} = \sqrt{\frac{28}{2}} = \sqrt{14} = 3.74$$

Conclusion: Hydrogen effuses approximately 3.74 times faster than Nitrogen.

Example 4.9

Given:

100 cm³ of Nitrogen (N_2) effuses through a membrane in 50 seconds. How long will it take for 100 cm³ of Chlorine to effuse through the same membrane?

$M_{N_2} = 28 \text{ g mol}^{-1}$, $M_{Cl_2} = 71 \text{ g mol}^{-1}$

Answer

$$\frac{\text{Rate of effusion of } N_2}{\text{Rate of effusion of } Cl_2} = \sqrt{\frac{M_{Cl_2}}{M_{N_2}}}$$

$$\frac{\text{Rate of effusion of } N_2}{\text{Rate of effusion of } Cl_2} = \sqrt{\frac{71}{28}} = 1.59$$

$$\text{Rate of diffusion of } N_2 = \frac{\textit{Volume} \text{ of } N_2}{\textit{time for} \text{ of } N_2}$$

$$= \frac{100 \text{ cm}^3}{50 \text{ s}} = 2 \text{ cm s}^{-1}$$

$$\Longrightarrow \frac{2\text{cm s}^{-1}}{\text{Rate of effusion of } Cl_2} = 1.59$$

$$\therefore \text{Rate of effusion of } Cl_2 = \frac{2\text{cm s}^{-1}}{1.59}$$

$$= 1.26 \text{ cm s}^{-1}$$

$$\textit{time for } Cl_2 = V\frac{\textit{olume of } Cl_2}{\text{Rate of diffusion of } N_2}$$

$$= \frac{100 \text{ cm}^3}{1.26 \text{ cm s}^{-1}} = 79.4 \ s$$

DALTON'S LAW OF PARTIAL PRESSURES

It states that, in a mixture of gases which do not react, the total pressure exerted is equal to the sum of the partial pressures of the individual gases at constant temperature.

For a mixture of gases 1 and 2,

$P_T = P_1 + P_2$

where P_T = total pressure

P_1 and P_2 are the partial presures of gases 1 and 2

$P_1 = X_1 P_T$

$P_2 = X_2 P_T$

where X_1 and X_2 are the mole fractions of gases 1 and 2

$X_1 = \frac{n_1}{n_1 + n_2}$

$X_2 = \frac{n_2}{n_1 + n_2}$, where n_1 and n_2 are the number of moles of gases 1 and 2

Example 4.10

Consider the reaction,

$N_2 + 3H_2 \rightarrow 2NH_3$

If the total pressure is 150 atm, calculate the partial pressure of Nitrogen gas.

Answer: Use problem – solving approach

Determine the mole ratio of the gaseous species involved

Mole fraction of N_2,

$$X_{N_2} = \frac{n_{N_2}}{n_{N_2} + n_{H_2} + n_{NH_3}}$$

$$= \frac{1}{1 + 3 + 2}$$

$$= \frac{1}{6}$$

$$P_{N_2} = X_{N_2} \times P_T$$

$$= \frac{1}{6} \times 150$$

$$= 25.05\ atm$$

Example 4.11

Gases A, B and C have partial pressures 2 atm, 3 atm and 1 atm respectively. Calculate the total pressure of the gases.

Answers

$P_{total} = P_A + P_B + P_C$

$= 2 + 3 + 1$

$= 6$ atm

Finding partial pressure

Example 4.12

The total pressure of a mixture of gases A, B and C is 10 atm. The partial pressure of A is 4 atm and B is 5 atm. Calculate the partial pressure of gas C.

Answer

$P_{total} = P_A + P_B + P_C$

$10 = 4 + 5 + P_C$

$P_C = 10 - 9$

$= 1\ atm$

Example 4.13

A gas mixture contains 20% Oxygen (O_2), 30% Nitrogen (N_2) and 50% Carbon dioxide (CO_2). If the total pressure is 750 mmHg,

Calculate the partial pressures of Oxygen, Nitrogen and Carbon dioxide

Answers

$P_i = X_i \times P_{total}$

$P_{O_2} = X_{O_2} \times P_{total}$

$= 0.20 \times 750$ mmHg

$= 150$ mmHg

$P_{N_2} = X_{N_2} \times P_{total}$

$= 0.30 \times 750$ mmHg

$= 225$ mmHg

$P_{CO_2} = 0.50 \times 750$ mmHg

$= 375$ mmHg

THE IDEAL GAS EQUATION

The ideal gas equation, also known as the ideal gas law, is a fundamental equation in chemistry that describes the behaviour of gases under certain conditions. It relates the Pressure (P), Volume (V), Temperature (T) and amount of gas in moles (n) of an ideal gas sample. It combines Boyle's law, Charles's Law and Avogadro's law.

$PV = nRT$

Where:

P = Pressure of gas Pascals;

V = Volume of gas meters cubed;

n = Moles of gas (moles);

T = Temperature of gas (Kelvin)

R = Ideal gas constant = 8.314 J mol^{-1} K^{-1}(S.I. unit of R),

Other units of R = 0.082057 L atm mol^{-1} K^{-1}, 62.364 L Torr mol^{-1} K^{-1}

NB: Because of the different units of R, it is important to always match the units of pressure, volume, number of moles and temperature given with the units of R.

If the value of R is given as 0.082057 L atm mol^{-1} K^{-1}, the unit for pressure must be atm, the unit for volume must be litre and for temperature must be Kelvin.

If the value of R is given as 62.364 L Torr mol^{-1} K^{-1}, the unit for pressure must be Torr, for volume the unit must be litre, and for temperature must be Kelvin.

The ideal gas equation is essential for understanding and predicting the behaviour of gases in various chemical reactions and processes.

Example 4.14

What is the volume of 10 g of nitrogen gas at 25 ^{0}C and 101 kPa? [N =14, R = 8.314 $Jmol^{-1}K^{-1}$]

Answer (Use problem-solving strategy)

Known

N =14, R = 8.314 $Jmol^{-1}K^{-1}$; Temperature, T = 25 + 273 = 298 K; Mass, m = 10 g

Pressure, P = 101 kPa = 101 000 Pa; R = 8.314 $Jmol^{-1}K^{-1}$

Unknown

Volume, V =?

$$M(N_2) = 2(14)$$
$$= 28 \text{ g}$$

$$n = \frac{m}{M}$$
$$= \frac{10}{28}$$
$$= 0.357 \text{ mol}$$

$$PV = nRT$$

$$V = \frac{nRT}{P}$$
$$= \frac{0.357 \times 8.314 \times 298}{101000}$$
$$= 0.0088 \text{ m}^3$$

$\therefore$ volume = 0.0088 m^3

Activity 4.12: To share understanding of Boyle's Law, Charles' Law, and Avogadro's Law.

In your group:

1. Discuss and explain (Boyle's, Charles', or Avogadro's) law in your own words.
   - **a.** How does pressure relate to volume in Boyle's Law?
   - **b.** What happens to the volume of gas when the temperature increases according to Charles' Law?

   c. How does adding more gas particles affect volume according to Avogadro's Law?

2. Give a simple example to illustrate the law.

3. Think of a real-life situation where Boyle's Law applies.

4. How do hot air balloons relate to Charles' Law?

5. When might Avogadro's Law be important in the real world?

6. Share your understanding and examples with the whole class.

7. Discuss how these laws (Boyle's, Charles', or Avogadro's) relate to one another.

   a. How might changing the temperature in Charles' Law affect the pressure in Boyle's Law?

   b. If you increase the number of particles according to Avogadro's Law, what might happen to the pressure or volume?

   c. How do these laws together help us understand the behaviour of gases?

Activity 4.13: Deriving the Ideal Gas equation

Boyle's Law: Pressure increases as volume decreases (like squeezing a balloon).

PV = constant, when T and n are constant.

Charles' Law: Volume increases as temperature rises (like a hot air balloon).

$V \propto T$, so $V = kT$, when P and n are constant.

Avogadro's Law: Volume increases with more gas particles (like inflating a balloon).

$V \propto n$, so $V = kn$, when P and T are constant.

Combining these three relationships into one equation that relates pressure (P), volume (V), temperature (T), and the number of moles (n) of a gas:

$$V \propto \frac{nT}{P}$$

To remove the proportionality sign, introduce a constant, R, so:

$$PV = nRT$$

Example 4.15

A car tyre is filled with 0.050 moles of air at a temperature of 20°C. The tyre has a volume of 10.0 liters. What is the pressure in the tyre in kPa?

Answer

Given Data:

$n = 0.050\,\text{mol}$; $V = 10.0\,\text{L}$; $T = 20 \circ C = 20 + 273.15 = 293.15\,\text{K}$

$R = 8.314\ \text{J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} = 8.314\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$

Use the Ideal Gas Law:

$PV = nRT$

Solve for *P*

P (pressure):

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

Substitute the values:

$$P = \frac{0.050\,\text{mol} \times 8.314\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 293.15\,\text{K}}{10.0\,\text{L}}$$

$$= 12.2\ \text{kPa}$$

The pressure in the tyre is **12.2 kPa**.

Example 4.16

You need to inflate a car tyre to a pressure of 2.5 atm. The tyre has a volume of 30.0 litres, and the temperature is 300 K. How many moles of air are required to achieve the desired pressure?

Answer

Use the Ideal Gas Law:

$PV = nRT$

Given:

$P = 2.5\,\text{atm}$; $V = 30.0\,\text{L}$; $T = 300\,\text{K}$; $R = 0.0821\,\text{L} \cdot \text{atm/mol} \cdot \text{K}$

Solve for n:

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

$$= \frac{2.5\,\text{atm} \times 30.0\,\text{L}}{0.0821\,\text{L} \cdot \text{atm/mol} \cdot \text{K} \times 300\,\text{K}}$$

$$= 3.05\ \text{mol}$$

The implications of using the ideal gas equation to predict the behaviour of a gas at extremely high pressures or low temperatures

The Ideal Gas Law is a simplified model that works well for predicting gas behavior under a wide range of conditions but fails under extremes of high pressure and low temperature. In these situations, molecular interactions and phase changes cause significant deviations from ideal behavior. To accurately describe gases under such conditions, more complex equations like the Van der Waals equation or empirical models that account for real gas behavior must be used. Understanding these limitations is crucial for applications where precision is necessary, such as in chemical engineering, material science, and thermodynamics.

NON-IDEAL GAS BEHAVIOUR

A real gas behaves differently from what is expected in ideal conditions at:

1. High-pressure
2. Low temperature

This is due to the following:

a. Real gases have an actual volume of molecules, which is significant at very high pressure. At extremely high pressures, the value of PV becomes greater than the ideal value. The deviation increases when the relative molecular mass (Mr) increases.

b. Intermolecular forces always exist in real gases. At lower temperatures, the kinetic energy of the molecules is at their lower value, intermolecular forces increase, which reduces the pressure P, making the PV in value less than the ideal value. Examples of gases with greater intermolecular forces are those which are more polar.

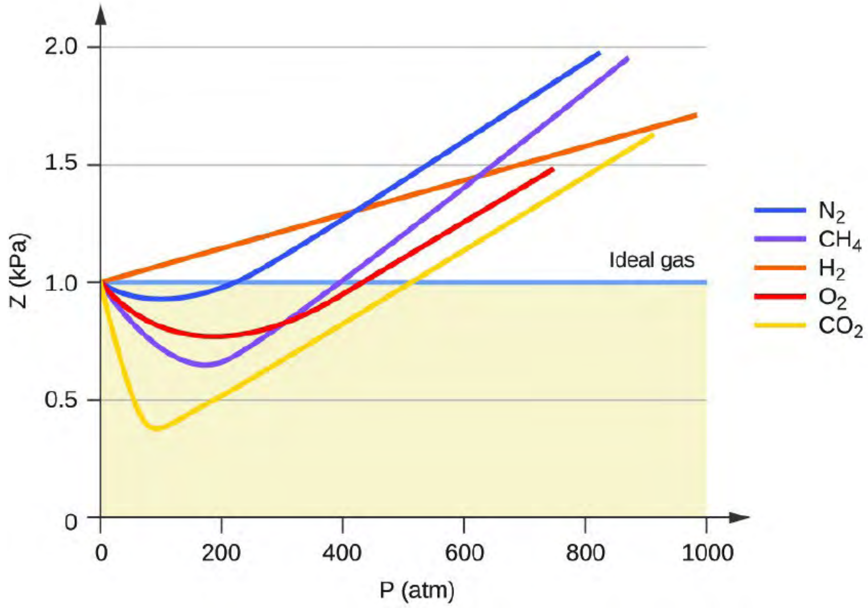


Figure 4.5: Deviations from ideal gas law for select gases.

Example 4.17

Arrange these gases in their order of deviation from ideal behaviour and explain the order: O_2, Ne, NH_3

Answer

Ne < O_2 < NH_3

Increasing order of deviation

Reasons: Neon, a noble gas, exhibits the least deviation from ideal behaviour. This is because noble gases have very weak intermolecular forces (it is non-polar), and their molecules are far apart, resulting in minimal interactions between them. Oxygen, while also a non-polar molecule like neon, has slightly stronger intermolecular forces due to its higher molar mass. However, these forces are still relatively weak, compared to other gases, resulting in a moderate deviation from ideal behaviour. Ammonia, a polar molecule with hydrogen bonding, experiences the highest deviation from ideal behaviour. Hydrogen bonding leads to stronger intermolecular attraction between ammonia molecules, compared to neon and oxygen, causing greater deviation from the ideal gas law.

Assumptions of the Ideal Gas Law

Ideal Gas Law: PV = nRT

1. **Particles have zero volume:** Gas particles are considered as point particles with no volume, meaning they occupy no space.
2. **No intermolecular forces:** The law assumes that there are no attractive or repulsive forces between gas particles, so they move independently of one another.
3. **The gas consists of a large number of molecules:** These molecules are in constant random motion.
4. **Collisions between molecules (and with the walls) are perfectly elastic:** No kinetic energy is lost during collisions, meaning the total kinetic energy of the system remains constant.
5. **The time taken for a collision is negligible:** The duration of collisions between molecules is so short that it can be ignored compared to the time between collisions.
6. **The motion of molecules follows Newton's laws:** Specifically, Newtons second law applies to the motion of the molecules between collisions, meaning their motion is governed by classical mechanics.

Use the link below to watch a video on ideal gases:

https://www.youtube.com/watch?v=Hr5Baj3lXFA&t=106s

Relating assumptions of Ideal Gas Law to Real-Gases

The ideal gas law works well under many conditions, it doesn't perfectly describe real gases, especially at very high pressures or very low temperatures, where gas particles do interact and have volume.

Activity 4.14: The Ideal Gas Law equation and its underlying assumptions.

Group Discussion: In small groups discuss the conditions under which real gases deviate from ideal behavior

1. *High Pressure*: Discuss why real gases deviate from ideal behaviour at high pressures.

   a. What happens to the volume of gas particles under high pressure?

   b. How does this affect the distance between particles and their interactions?

2. *Low Temperature*: Discuss why real gases deviate from ideal behaviour at low temperatures.

   a. How does a decrease in temperature affect the kinetic energy of gas particles?

   b. What role do intermolecular forces play when gas particles move more slowly?

3. *Large Gas Molecules*: Discuss how the size of gas molecules can cause deviations from ideal behaviour.

   a. How does the physical size of gas molecules affect their volume?

   b. Why might larger molecules not behave ideally compared to smaller molecules?

4. *Strong Intermolecular Forces*: Discuss how strong intermolecular forces lead to deviations from ideal behaviour.

   a. What are intermolecular forces, and how do they affect gas particles?

   b. In what way do these forces influence the behaviour of gases under different conditions?

Activity 4.15: Demonstrating Deviations from Ideal Gas Behavior

High-Pressure Demonstration with a Carbonated Drink

Objective: To show how gases dissolve in liquids under high pressure, which deviates from the Ideal Gas Law.

Materials Needed: Carbonated drink (e.g., soda), safety goggles and gloves. Carbonated drinks contain dissolved CO_2 gas. At high pressure (inside the sealed bottle), CO_2 stays dissolved in the liquid.

Procedure:

1. Shake a sealed bottle of soda and then carefully open it.

2. State what you observe.

Discussion: Explain your observation.

Use the link below to watch a video comparison of real gas behavior versus ideal gas law predictions:

https://www.khanacademy.org/science/ap-chemistry-beta/x2eef969c74e0d802:intermolecular-forces-and-properties/x2eef969c74e0d802:deviation-from-ideal-gas-law/v/real-gases-deviations-from-ideal-behavior

Example 4.18

Explain why real gases deviate from ideal behaviour?

Answer

Real gases deviate from ideal behaviour because the assumptions made in the ideal gas law do not hold under certain conditions. These assumptions are:

1. Particles have zero volume
2. No intermolecular forces
3. The gas consists of a large number of molecules
4. Collisions between molecules (and with the walls) are perfectly elastic
5. The time taken for a collision is negligible
6. The motion of molecules follows Newton's laws

VAN DER WAALS EQUATION

The Van der Waals equation is an equation of state that extends the ideal gas law to include the non-zero size of gas molecules and the interactions between them.

For one mole of gas, the equation is

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

and for n moles of gas, the Van der Waals equation is

$$(P + \frac{an^2}{V^2})(V - nb) = n\,RT$$

Where;

a = a measure of the strength of the intermolecular forces.

b = The excluded molar volume

Activity 4.16: Understanding the Van der Waals Equation and Its Purpose in Describing the Behaviour of Real Gases

Objective: To explore the Van der Waals equation, understand how it accounts for the behaviour of real gases, and learn why it differs from the Ideal Gas Law.

Steps:

1. Write the Ideal Gas Law
2. Discuss the assumptions of the Ideal Gas Law.
3. State the assumptions might not hold true for real gases.
4. Write the Van der Waals Equation.
5. Explain the significance of each term in the equation.
6. Discuss how the Van der Waals equation modifies the Ideal Gas Law to describe real gases more accurately.
7. Think about how the constants a and b would change under different conditions.

Extended task

1. State differences between the ideal gas law and the Van der Waals equation.
2. Use the Van der Waals equation to calculate the pressure of 1 mole of carbon dioxide gas in a 2-litre container at 273 K. Given: $a = 3.592$, $b = 0.0427$ and $\mathbf{R} = 8.314$ J $mol^{-1}K^{-1}$
3. Compare this value to that for the ideal gas equation, express the difference as a percentage.

PREPARATION AND TEST FOR HYDROGEN GAS

Equation for preparation:

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

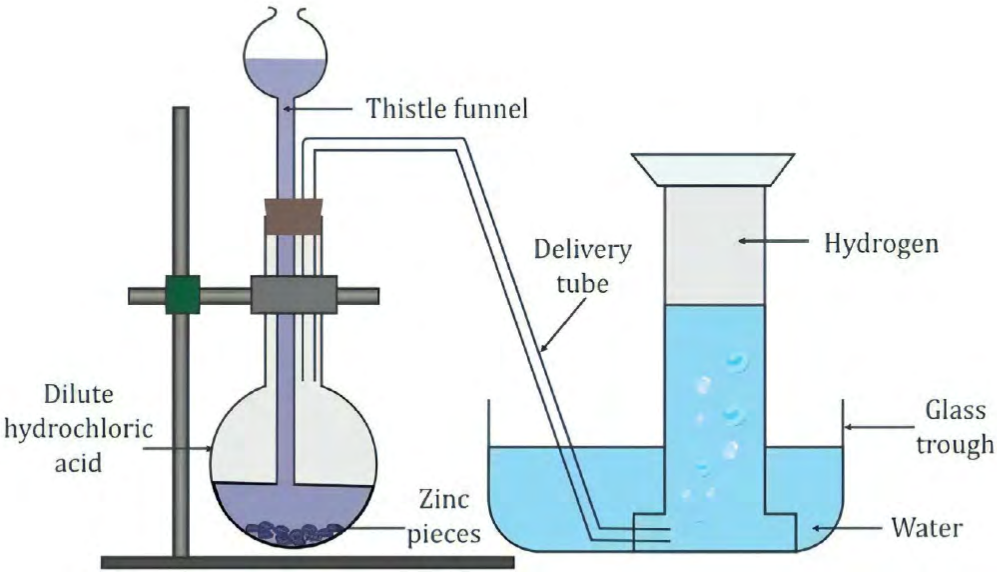


Figure 4.6: Preparation of Hydrogen gas

Test for hydrogen gas: Put a burning/lighted splint in the gas; a "pop" sound indicates the presence

of hydrogen gas.

How to dry hydrogen gas: Pass the gas produced through anhydrous $CaCl_2$.

Physical properties of hydrogen gas

- **a.** It is the lightest gas and has the lowest density.
- **b.** It is colourless and odourless.
- **c.** It is insoluble in water.

Chemical properties of hydrogen gas

- **a.** It is neutral to litmus.
- **b.** It is unreactive under normal conditions.
- **c.** It does not support combustion.
- **d.** It burns in oxygen (air) with the pop sound to produce water.

  $2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$
- **e.** It reacts with halogens

  e.g. $H_2\,(g) + F_2\,(g) \rightarrow 2HF\,(g)$

Uses of hydrogen gas in everyday life

a. It is used to produce ammonia gas in the Haber process.
b. It is used to manufacture margarine by hydrogenation of unsaturated fats.
c. It is used in oxy-hydrogen flame for cutting and welding of metals.
d. It is used in fuel cells.

Activity 4.17: Exploring the uses of hydrogen gas in everyday life

Objective: Investigate and present the various uses of hydrogen gas in everyday life.

Materials: Internet access for research (if available), handouts or printed resources on hydrogen gas.

Steps

In small groups find the use of hydrogen gas in areas such as:

1. Fuel and energy
2. Industrial processes
3. Food industry
4. Chemical industry

Activity 4.18: Preparation and Collection of Hydrogen Gas

Materials Needed: Zinc granules or magnesium ribbon, dilute hydrochloric acid (HCl) or sulfuric acid (H_2SO_4), conical flask (250 mL), thistle funnel or dropping funnel, rubber stopper with a single hole, delivery tube, water-filled trough or basin, gas jar or test tubes, wooden splints, matches or lighter, safety goggles, lab coat or apron, gloves

Safety Precautions:

- Ensure you wear safety goggles and a lab coat.
- Handle acids with care.
- Conduct the experiment in a well-ventilated area or under a fume hood.

Procedure

1. Weigh out a defined mass of zinc granules or a piece of magnesium ribbon and place into the conical flask.

2. Insert the thistle funnel into the flask through the rubber stopper, ensuring the end of the funnel is submerged in the acid when added.
3. Connect the delivery tube to the rubber stopper and place the other end under the water in the trough or basin.
4. Position a gas jar or test tube upside down over the end of the delivery tube in the trough to collect the gas.
5. Slowly add a defined volume dilute hydrochloric acid (of known concentration) through the thistle funnel. The acid will react with the zinc, producing hydrogen gas.
6. Observe the formation of bubbles as the gas is generated.
7. The hydrogen gas will travel through the delivery tube and displace water in the gas jar or test tube, filling it with hydrogen gas.
8. Once enough gas is collected, carefully remove the gas jar or test tube from the water, keeping it upside down.
9. Light a wooden splint and bring it near the mouth of the jar/test tube. The presence of hydrogen gas will be confirmed by a characteristic "pop" sound when it ignites.

Chemical reaction

$$\mathrm{Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)}$$

$$\mathrm{Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)}$$

$$\mathrm{Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)}$$

Extended task: Discussions on Activity 4.18

1. What did you observe when the acid was added to the zinc?
2. How did you confirm the presence of hydrogen gas?
3. Why does hydrogen gas produce a "pop" sound when ignited?
4. State the method of collecting hydrogen gas and explain why it is collected by this method.
5. Calculate the number of moles of each reactant and determine which was in excess.
6. Calculate the number of moles of gas produced using the ideal gas equation.

7. Calculate the number of moles of gas produced using a mass and mole balance calculation
8. Compare solution from question 6 with that in question 7
9. State **two** properties of hydrogen that enable it to be used to fill balloons.
10. State **three** practical uses of hydrogen in everyday life.
11. You are provided with the thistle funnel, delivery tube, split cork, conical flask, gas jar, beehive stand, water trough, Magnesium ribbon and dilute HCl. Design an experiment to prepare and test for hydrogen gas.

PREPARATION AND TEST FOR CARBON DIOXIDE GAS

Equation for preparation: $CaCO_3(s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + H_2O(l) + CO_2(g)$:

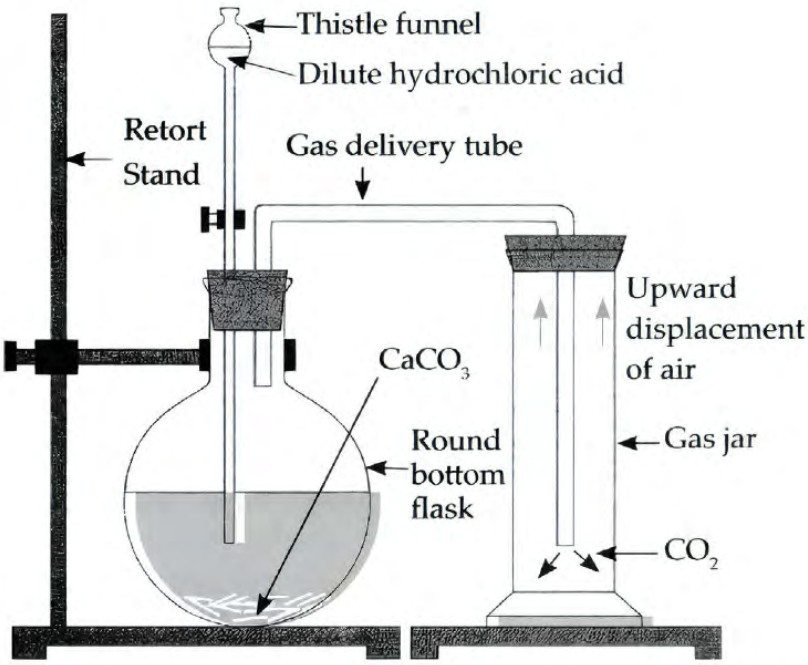


Figure 4.7: Preparation of Carbon dioxide

Test for carbon dioxide gas: Pass the gas through lime water (saturated solution of calcium hydroxide), which turns milky.

How to dry carbon dioxide gas: Pass the gas produced through concentrated sulphuric acid. As an alternative (as concentrated sulphuric acid is hazardous to handle) it can be passed over anhydrous $CaCl_2$.

Physical properties of carbon dioxide gas

1. It is colourless and odourless.
2. It is denser than air.
3. It is soluble in water.
4. It condenses into a white solid called dry ice.

Chemical properties of carbon dioxide gas

1. It turns moist blue litmus red.
2. It turns lime water ($Ca(OH)_2$ solution) milky.
   $$CO_2\,(g) + Ca(OH)_2(aq) \rightarrow CaCO_3\,(s) + H_2O(l)$$
3. Excess passage of CO_2 causes milkiness to disappear.
   $$CaCO_3\left(s\right) + H_2O\left(l\right) + CO_2\left(g\right) \rightarrow Ca(HCO_3)_2\left(aq\right)$$
4. It reacts with water in the presence of sunlight and chlorophyll to produce glucose and oxygen (photosynthesis).
   $$6CO_2\,(g) + 6H_2O\,(l) \rightarrow C_6H_{12}O_6\,(s) + 6O_2(g)$$
5. It does not support combustion.

Uses of Carbon dioxide in everyday life

1. It is used in photosynthesis to produce glucose and oxygen.
2. It is used in fire extinguishers to extinguish fires.
3. It is dissolved into fizzy drinks.
4. It is used to manufacture refrigerants.

Activity 4.19: Preparation and Collection of Carbon Dioxide Gas

Materials Needed: marble chips ($CaCO_3$), dilute hydrochloric acid (HCl), Conical flask (250 mL), thistle funnel or dropping funnel, rubber stopper with a single hole, delivery tube, water-filled trough or basin, gas jar or test tubes, Wooden splints, matches or lighter, safety goggles, lab coat.

Procedure:

1. Weight out a defined mass of marble chips (calcium carbonate) and place them into the conical flask.
2. Insert the thistle funnel into the flask through the rubber stopper, ensuring the end of the funnel is submerged in the acid when added.

3. Connect the delivery tube to the rubber stopper and position the other end under the water in the trough or basin.
4. Place a gas jar or test tube over the end of the delivery tube to collect the gas.
5. Slowly add a defined volume of dilute hydrochloric acid (of a known concentration) through the thistle funnel. The acid will react with the marble chips, producing carbon dioxide gas.
6. Observe the formation of bubbles as the gas is generated.
7. The carbon dioxide gas will travel through the delivery tube and displace water in the gas jar or test tube, filling it with carbon dioxide gas.
8. *Test for Carbon Dioxide:*

Limewater Test: Pour a small amount of limewater into another test tube. Use the gas collected to bubble through the limewater. The presence of carbon dioxide will be confirmed if the limewater turns milky or cloudy.

9. **Reaction:** $CaCO_3\,(s) + 2HCl\,(aq) \rightarrow CaCl_2\,(aq) + H_2O\,(l) + CO_2(g)$

Discussion:

10. What did you observe when the acid was added to the marble chips?
11. How did you confirm the presence of carbon dioxide gas?
12. Why does carbon dioxide gas make limewater turn cloudy?
13. Why does carbon dioxide gas extinguish a flame?
14. Calculate the number of moles of each reactant and determine which was in excess.
15. Calculate the number of moles of gas produced using the ideal gas equation.
16. Calculate the number of moles of gas produced using a mass and mole balance calculation
17. Compare solution from question 6 with that in question 7
18. What are some uses of carbon dioxide gas in daily life or industry?

PREPARATION AND TEST FOR AMMONIA GAS

Equation for preparation: $Ca(OH)_2(aq) + 2NH_4Cl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l) + 2NH_3(g)$

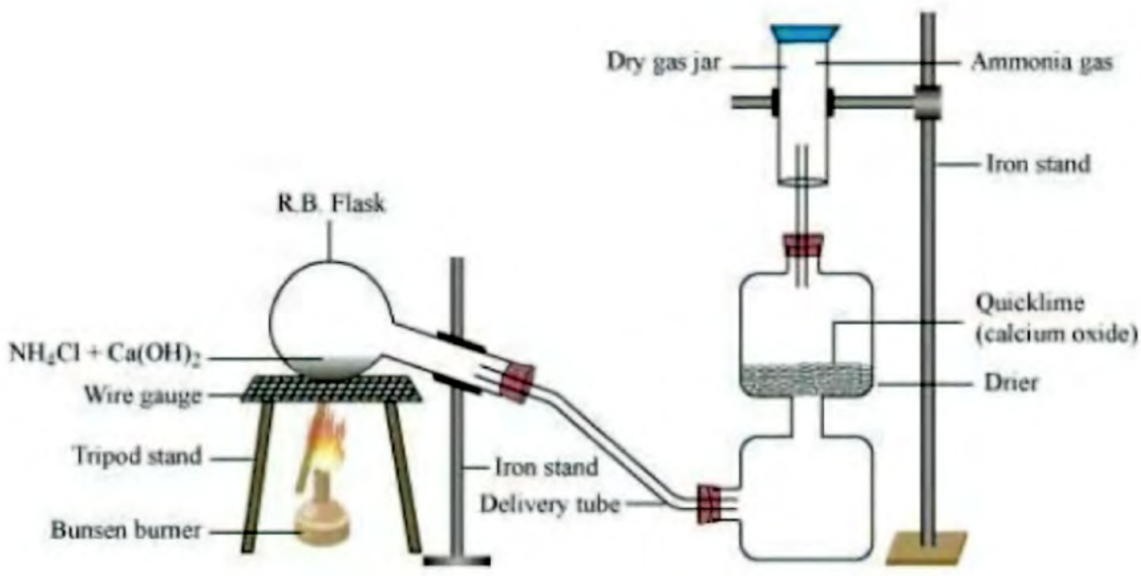


Figure 4.8: Preparation of Carbon dioxide

Test for gas:

a. It turns moist red litmus blue.
b. It forms white fumes with concentrated HCl vapour to form NH_4Cl.

How to dry the gas: Pass the gas produced over solid calcium oxide (CaO).

Physical properties

a. It is a colourless gas with a pungent, choking smell.
b. It is less dense than air.
c. It is soluble in water.

Chemical properties

a. It turns moist red litmus blue.
b. It forms dense white fumes with HCl; $NH_3 + HCl \rightarrow NH_4Cl$
c. It burns in oxygen with a pale-yellow flame.

Uses of Ammonia gas in everyday life

a. It is used to manufacture fertilisers.
b. It is used to manufacture explosives.

c. It is used to manufacture nylon.

d. It is used to manufacture plastics.

e. It is used to manufacture pigment.

Activity 4.20: Preparation and Collection of Ammonia Gas

Materials Needed: Ammonium chloride (NH_4Cl), Calcium hydroxide ($Ca(OH)_2$) or sodium hydroxide (NaOH) pellets, test tube or small conical flask, Bunsen burner, rubber stopper with a single hole, delivery tube, water-filled trough or basin, gas jar, red litmus paper, concentrated hydrochloric acid (HCl) glass rod or dropper, safety goggles, Lab coat or apron, gloves.

Safety Precautions:

- Ensure you wear safety goggles and a lab coat.
- Handle chemicals, especially concentrated acids and bases, with care.
- Conduct the experiment in a well-ventilated area or under a fume hood, as ammonia gas is pungent and irritating.

Procedure

1. Mix ammonium chloride and calcium hydroxide (or sodium hydroxide) in a test tube or small conical flask.
2. Insert the delivery tube into the test tube or flask through the rubber stopper, ensuring a tight fit.
3. Collect the gas by upward displacement of air (since ammonia is less dense than air).
4. **Discuss the Reaction:**
   $Ca(OH)_2(aq) + 2NH_3Cl(aq) \rightarrow CaCl_2(aq) + 2H_2O\,(l) + 2NH_3(g)$
5. **Heat the Mixture:**
   a. Gently heat the test tube containing the mixture using a Bunsen burner. The heat will cause the ammonium chloride and calcium hydroxide to react, releasing ammonia gas.
   b. Observe the release of gas, which can be recognized by its pungent smell.
6. The ammonia gas will travel through the delivery tube. If using water displacement, it will collect in the gas jar or test tube, displacing the water. If collecting by upward displacement of air, it will fill the gas jar.

7. Observe the smell of the gas, the reaction with litmus paper, and the formation of ammonium chloride fumes.

8. **Test for Ammonia:**
   a. **Litmus Paper Test:** Hold a piece of moist red litmus paper near the mouth of the gas jar or test tube. The paper will turn blue, indicating the presence of ammonia gas (a basic gas).
   b. **Fume Formation:** Dip a glass rod or dropper into concentrated hydrochloric acid and hold it near the mouth of the gas jar. White fumes of ammonium chloride will form, confirming the presence of ammonia.

Discussion:

1. What did you observe when the ammonium chloride and calcium hydroxide were heated?
2. How did you confirm the presence of ammonia gas?
3. Why does ammonia gas turn red litmus paper blue?
4. Why is ammonia collected by upward displacement of air?
5. What are some uses of ammonia gas in daily life or industry?
6. Why is it particularly important to conduct this experiment under a fume hood?

EXTENDED READING

Read about the gas laws using the recommended material below:

- General, Organic, and Biochemistry, Tenth Edition, Katherine J. Denniston, Joseph J. Topping,

REVIEW QUESTIONS

Review Questions 4.1

1. Draw a simple diagram illustrating the arrangement and motion of particles in a solid and a liquid.
2. Compare and contrast the behaviour of particles in solids and liquids according to the kinetic theory of matter.
3. In a laboratory experiment, students investigate the effect of temperature on the viscosity of a liquid. Describe how the kinetic theory of matter can be applied to explain the observed changes in viscosity as temperature increases.
4. What real-world phenomena can be explained by the Kinetic Theory of Matter, and how does understanding this theory help in practical applications?

Review Questions 4.2

1. Describe how the kinetic theory of matter explains the behaviour of gases. Provide one example to illustrate how gas particles move and interact with each other.
2. A sealed container of gas is heated, causing an increase in temperature. Explain how the kinetic theory of matter can be used to predict the changes in gas pressure and volume inside the container.
3. A gas cylinder is compressed to half its original volume while maintaining a constant temperature. Using the principles of the kinetic theory of matter, explain how the gas pressure changes because of compression.

Review Questions 4.3

1. Explain the difference between condensation and sublimation, providing examples of substances that undergo each of these changes of state.
2. Describe how a solid can change directly into a gas without passing through the liquid state, using examples from everyday life.

3. Analyse the impact of changes in atmospheric pressure on the boiling point of a liquid. Explain how altitude affects the boiling point of water and its implications for cooking and food preparation.

Review Questions 4.4

1. State the definition of the following gas laws:
   a. Boyles' law
   b. Charles' law
   c. Gay-Lussac's law
   d. Combined gas law
2. Express any two (2) of the gas laws mathematically taking into consideration the conditions associated with each law.
   a. Identify the variables that are related to each law.
   b. Illustrate the relationships among the various variables of the law.
3.
   a. Explain how Boyles' law can be applied in real life situation.
   b. Derive the combined gas law from Boyle's, Charles' and Gay–Lussac's laws.
4.
   a. 20 m3 of gas at a pressure of 100,000 Pa was compressed to a pressure of 400,000 Pa at constant temperature. Calculate the final volume of the gas.
   b. 20 dm3 of gas in a cylinder is heated from 150 K to a certain temperature. If its volume expands to 35 dm3, Calculate the final temperature of the gas.
   c. 35 cm3 of a gas at 1.5 atm at 40 0C was compressed to 20 cm3at 50 0C, calculate the final pressure of the gas.
   d. A balloon was filled to a volume of 2.0 dm3 with 0.082 moles of helium gas. Suppose 0.015 moles of helium is added to the balloon with constant temperature and pressure, what will be the new volume of the balloon?
   e. The temperature of a gas contained in a cylinder undergoes combustion in constant volume, increasing from 30 0C to 1500 0C.

If the normal atmospheric pressure is 101 kPa, calculate the peak pressure reached after combustion.

5. A 250 cm3 sample of a gas has a pressure of 20 Pa. What will be its volume if the pressure is raised to 45 Pa at the same temperature?
6. A balloon can hold 800 cm3 of air before bursting. If the balloon contains 500 cm3 of air at 10 0C, determine whether the balloon would burst when it is taken into a room of temperature of 20 0C, assuming that the pressure of the gas in the balloon is kept constant.

Review Questions 4.5

1. State Graham's law of diffusion or effusion.
2. Given that 100 cm3 of ethane (C2H6) diffuses through a membrane in 40 s, what time will it take 80 cm3 of propane (C3H8) to diffuse through the same membrane at the same temperature and pressure? [C = 12, H = 1].
3. Given a cylindrical glass tube, stoppers, two retort stands, cotton wool, concentrated NH3, concentrated HCl, and tweezers, design an experiment to determine the rate of diffusion of a gas.
4. A gas mixture contains methane (CH_4) and Sulfur dioxide (SO_2). The molar mass of methane is 16 g/mol, and the molar mass of Sulfur dioxide is 64 g/mol. Determine the ratio of their effusion rates.
5. Two gases, A and B, have molar masses of 4 g/mol and 36 g/mol respectively. If gas A effuses through a membrane in 30 seconds, how long will it take gas B to effuse the same amount?
6. An unknown gas X effuses at half the rate of Oxygen. What is the molar mass of gas X?

Review Questions 4.6

1. State Dalton's law of partial pressures.
2. A mixture of gases contains 4.76 mole of Ne, 0.74 mole of Ar and 2.5 mole of Xe. Calculate the partial pressure of the gases if the total pressure is 2 atm, at a fixed temperature.
3. A mixture of 40.0 g of Oxygen and 40.0 g of Helium has a total pressure of 0.900 atm. What is the partial pressure of each gas? [H = 2, O = 16]

4. Three gases (8 g of CH4, 18 g of C2H6, and unknown amount of C3H8) were added to the same 10.0 dm3 container. At 23°C , the total pressure in the container was measured to be 4.43 atm. Calculate the partial pressure of each gas in the container. [C = 12, H = 1]
5. A gaseous mixture of O2 and N2 contains 32.8% Nitrogen by mass. What is the partial pressure of Oxygen in the mixture if the total pressure is 785.0 mmHg? [N = 14, O = 16]
6. A gas mixture contains 40% Helium (He), 30% and 60% Argon (Ar). If the total pressure is 400 mmHg. Calculate the partial pressures of He, and Ar
7. A gas mixture of gases A, B and C contain 2 mol, 3 mol and 5 mol respectively. The total pressure of the gases is 10 atm. Calculate the:
   a. mole fraction of gases A, B and C.
   b. partial pressures of gases A, B and C.

Review Questions 4.7

1. A helium balloon with a volume of 2.5 litres is filled to a pressure of 1.2 atm and a temperature of 290K. How many grams of helium are in the balloon? (Molar mass of helium = 4.00 g/mol)
2. You have a balloon that can hold 2.50 litres of air. If you want to inflate the balloon at room temperature (25°C) and a pressure of 1.00 atm, how many moles of air are needed?
3. A room has a volume of 100 L. Due to a gas leak, the pressure inside the room rises to 1.1 atm. The temperature in the room is 295 K. If the room was initially filled with air at 1.0 atm and the same temperature, how many moles of gas leaked into the room?
4. Under what conditions are deviations from ideal gas behaviour most likely to occur?
5. A real gas behaves more like an ideal gas at high temperatures and low pressures. Explain why this is the case.
6. Evaluate the significance of intermolecular forces in causing deviations from ideal behaviour in real gases. Discuss how these forces affect gas behaviour under different conditions.
7. Why do real gases deviate from the ideal gas law at high pressures?

8. Explain how the volume of gas molecules contributes to the limitations of the ideal gas law at high pressures.
9. How does the Van der Waals equation address the limitations of the ideal gas law?
10. What is the Van der Waals equation, and how does it differ from the ideal gas law?
11. How do you calculate the pressure of 1 mol of a real gas using the Van der Waals equation?
12. What is the significance of the constants a and b in the Van der Waals equation?
13. How do you determine the volume (V) of a gas using the Van der Waals equation?
14. How do the Van der Waals constants a and b vary between different gases, and why is this important?
15. What safety precautions should be taken when preparing hydrogen gas?
16. Why is hydrogen considered a potential clean energy source?
17. What safety precautions should be taken when preparing carbon dioxide gas?
18. How can you test the presence of carbon dioxide gas?
19. How does carbon dioxide contribute to the greenhouse effect?
20. What role does carbon dioxide play in the carbon cycle?
21. What safety precautions should be taken when preparing ammonia gas?
22. How can you confirm the presence of ammonia gas?
23. State the properties of CO2 that allow it to be used in the manufacture of:
    a. fizzy drinks
    b. gas refrigerants.
24. In the laboratory preparation of ammonia gas, the round bottom flask containing the reagents is slanted downwards whilst being heated. Explain why these actions are necessary in the preparation of the gas.
25. What role do intermolecular forces play in causing deviations from ideal gas behaviour?

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GLOSSARY

- **Kinetic Energy** is the energy that particles possess due to their motion.
- **Temperature** is a measure of the average kinetic energy of the particles.
- **Pressure** is the overall force exerted by particles colliding with the walls of their container.
- **Volume** is the space occupied by matter.
- **Diffusion** is the process by which gas molecules spread out (from an area of high pressure/concentration) to fill a container.
- **Effusion** is the process by which gas molecules escape through a small hole into a vacuum.
- **Rate of diffusion** is the speed at which particles diffuse.
- **Partial pressure** is the pressure exerted by a specific gas in a mixture.
- **Mole fraction** is the proportion of a specific gas in a mixture.
- **Total pressure**: is the sum of the partial pressures of all gases in a mixture.
- **Gas mixture** is a combination of two or more gases.
- **Ideal gas** is a hypothetical gas that perfectly follows the ideal gas equation.
- **Mode of gas collection of gas** refers to the method used to collect and store gas.

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