Roseville Secondary School Enugu

A project of Ikota Educational Foundation.

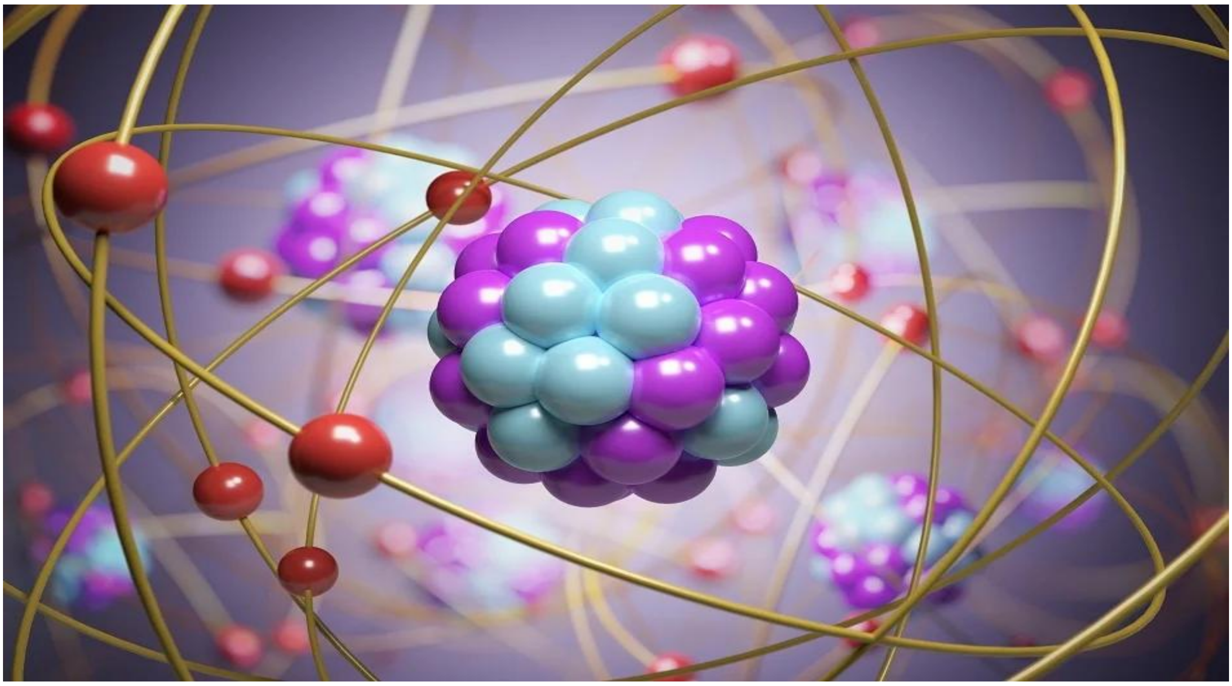
FIRST TERM 2024/2025 ACADEMIC SESSION

SUBJECT : CHEMISTRY

CLASS: SS 1

EXPECTATIONS

1. COPY YOUR NOTE **OR** PRINT AND SPIRAL BIND
2. THREE GRADED ASSESSMENT OF 20 MARKS BEFORE CAT = 60%
3. CAT : 40%
4. NON-GRADED ASSESSMENT
5. PROJECT TO BE SUBMITTED IN FRIDAY OF WEEK 6
6. THREE GRADED ASSESSMENT OF 20 MARKS BEFORE EXAM= 60%
7. EXAMINATION 40%
8. PRACTICALS: FOUR



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SCHEME OF WORK

WEEK	TOPIC
1	Introduction to Chemistry
2	Particulate Nature of Matter I
3	Particulate Nature of Matter II
4	Particulate Nature of Matter III
5	**CAT/PROJECT/MID TERM BREAK**
6	
7	Chemical Symbols and Formulae I
8	Chemical Symbols and Formulae II
9	Chemical Combination
10	Chemical Equation

HOLIDAY ASSIGNMENT:

1. Read and summarize your week one note.
2. Answer the questions that follow

Deadline: Monday of resumption.

PROJECT

TOPIC: <u>Project Title: Investigating the Effects of Different Water Sources on Plant Growth</u>

DEPARTMENT: Science

RATIONALE: This project provides SS1 students with a comprehensive understanding of how different water qualities affect plant growth, integrating concepts from chemistry, physics, and biology.

SDGs INTEGRATED: SDG 1,2 AND 4(NO POVERTY, ZERO HUNGER AND QUALITY EDUCATION)

Objective: To study how different types of water (rainwater, tap water, and saltwater) affect the growth and health of plants.

Materials Needed:

- Potted plants (e.g., beans, tomatoes, or radishes)
- Rainwater, tap water, and saltwater
- Measuring cups or graduated cylinders
- Ruler or measuring tape
- pH meter or pH test strips
- Soil nutrient test kits
- Labels or markers for pots
- Safety gloves and goggles
- Notebook for recording observations
- Camera for documenting plant growth

Experimental Procedure:

1. Preparation:
   - Select a uniform type of plant and pot a consistent number of seeds in each pot with the same soil type.
   - Label each pot according to the water type it will receive: rainwater, tap water, or saltwater.
   - Ensure that the pots are placed in a location with similar light and temperature conditions.
2. Watering and Observation:
   - Prepare the different water sources:
     - Collect rainwater.
     - Use tap water.
     - Prepare saltwater by dissolving table salt in tap water (e.g., 1 teaspoon of salt per liter of water).
   - Water the plants with the respective water types in equal amounts (e.g., 100 ml) at regular intervals (e.g., every 2-3 days).
   - Record the initial height and appearance of the plants.
3. Monitoring Plant Growth:

- Measure and record the height of the plants weekly.
- Observe and document changes in leaf size, color, and overall health.
- Take photographs of the plants at regular intervals to visually track their growth.

4. Testing Soil and Water:
   - Periodically test the pH of the soil and the water sources using pH meters or test strips.
   - Use soil nutrient test kits to check for changes in soil nutrient levels over time.
5. Data Collection:
   - Record plant height, leaf size, and overall plant health in a notebook.
   - Note any differences in growth patterns and health between plants watered with different water sources.
6. Analysis:
   - Compare the growth rates and health of plants watered with rainwater, tap water, and saltwater.
   - Analyze how the different water types impact soil pH and nutrient levels.
   - Discuss the potential reasons behind the observed effects based on the chemical composition of the water sources.

Conclusion: Summarize the findings, highlighting:

- How rainwater, tap water, and saltwater differently impact plant growth and health.
- How changes in soil pH and nutrient levels relate to plant health.
- The potential implications for gardening and agriculture in different water quality conditions.

Safety Notes:

- Ensure proper handling of water and soil testing kits.
- Handle saltwater with care to avoid skin contact and irritation.

Deadline – Friday of week 6.

Week One
Topic: Introduction to Chemistry and Its Applications
Performance Objective: By the end of this lesson, students should be able to :

1. Define chemistry
2. Dissect the career prospects in chemistry;
3. Outline the applications of chemistry;
4. Describe the adverse effect of chemistry;
5. Analyse the scientific method investigation;

Definition of Chemistry:
Chemistry is the scientific study of matter, its properties, composition, structure, and the changes matter undergoes during chemical reactions. Chemistry is an experimental science. It is concerned with the properties of the substances around us and the way it interacts with each other. During these interactions, some substances may undergo certain changes. Chemistry is concerned with the study of the principles governing these changes. Chemistry plays a fundamental role in understanding the natural world and the substances that make up our universe.
Chemistry is the mother of all sciences. This is because subjects such as biology, physics, agricultural science are better understood with adequate knowledge of chemistry. Chemistry also plays vital roles in science related professions such as pharmacy, medicine, nursing, food science, biochemistry, radioactivity and so on. Consequently, students intending to pursue a career in any of these fields must have adequate knowledge and understanding of chemistry.

Branches of Chemistry: The major branches of Chemistry are:

- Physical Chemistry
- Biochemistry
- Organic Chemistry
- Inorganic Chemistry
- Physical Chemistry
- Industrial Chemistry
- Nuclear Chemistry
- Environmental Chemistry
- Analytical Chemistry

Importance of Chemistry

Chemistry can be seen in different aspects of the world such as;

Everyday Life: Chemistry is present in our daily lives, from cooking and cleaning to breathing and digestion.

Advancements: It drives innovations in medicine, technology, and industry.

Environmental Impact: Chemistry plays a crucial role in addressing environmental challenges and sustainability.

Career Prospects Tied to Chemistry

Chemistry-related Professions include

1. Chemist: Research, analyze, and develop new compounds and materials.

2. Pharmacologist: Study drugs and their effects on living organisms.

3. Chemical Engineer: Design and optimize processes in chemical and petrochemical industries.

4. Environmental Scientist: Focus on pollution control and conservation.

5. Biochemist: Study the chemical processes within living organisms.

6. Forensic Scientist: Analyze evidence in criminal investigations.

7. Teaching and Academia: Share knowledge and educate future chemists.

Application/Uses of Chemistry

A. In Hospitals:

- Medicinal Chemistry: Development of drugs and pharmaceuticals.
- Clinical Chemistry: Testing and analysis of blood and urine samples.
- Radiopharmaceuticals: Application of radioactive materials for medical imaging.

B. In Military: Development of explosives and propellants. Chemical detection and protection against chemical warfare.

C. In Teaching: Chemistry educators play a vital role in shaping future scientists.

D. In Chemical and Petrochemical Industries: Production of various chemicals, plastics, and fuels. Petrochemical processing for the creation of petroleum-based products.

E. In Agriculture: Pesticides and fertilizers to enhance crop production. Soil analysis and plant nutrition studies.

F. In Space Science: Analyzing celestial bodies and planetary composition. Research on the effects of space travel on materials.

Adverse Effects of Chemistry

A. Chemicals and Toxins: Improper handling of chemicals can lead to accidents and health hazards. Toxic substances can cause harm to both humans and the environment.

B. Drug Abuse: The misuse of drugs can have severe consequences on individuals and society.

C. Poisoning: Accidental or intentional ingestion of toxic substances.

D. Corrosion: Deterioration of materials due to chemical reactions with the environment.

E. Pollution: Release of harmful chemicals and pollutants into the air, water, and soil.

The Scientific Method

The scientific method is a systematic approach used by scientists to investigate and understand natural phenomena. It involves a series of steps that guide researchers in conducting experiments and making evidence-based conclusions. These steps include;

Step 1 Observation is the act of gathering information through the use of the five senses. keen observation in the scientific process is very important as it often serves as the starting point for scientific investigations. Examples of simple observations, include seeing a plant grow towards sunlight, noticing water boiling at a specific temperature, or observing the moon's phases.

Step 2: Problem Identification/ asking a question: after making observations, scientists identify a specific problem or question they want to investigate further. This problem becomes the foundation of their scientific inquiry. Examples of scientific problems could be "What factors affect plant growth?" or "How does temperature impact the speed of chemical reactions?"

Step 3: Hypothesis: A hypothesis is a proposed explanation for the observed problem. A hypothesis must be testable and falsifiable, meaning it can be proven wrong through experimentation.

If the hypothesis was supported, we might do additional tests to confirm it, or revise it to be more specific. If the hypothesis was not supported, we would come up with a new hypothesis.

Examples of hypotheses include "If plants are given more sunlight, then they will grow taller" or "Increasing the temperature will speed up the chemical reaction."

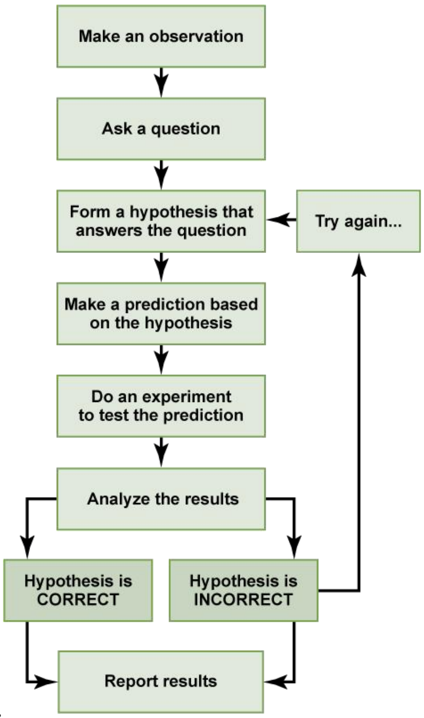
Step 4: Experiment: An experiment is a controlled procedure designed to test the validity of a hypothesis. It involves manipulating variables and measuring the outcomes to determine whether the hypothesis is supported or rejected.

Step 5: Theory: A theory is a well-substantiated explanation based on a substantial body of evidence from multiple experiments and observations. A scientific theory is not just a guess; it is a comprehensive and well-tested framework that explains a broad range of related phenomena.

Examples of scientific theories include the theory of evolution, the atomic theory, or the theory of relativity.

Step 6: Law: A scientific law is a statement that describes a consistent and universal relationship between different variables or phenomena. Unlike theories, laws do not attempt to explain why something happens; instead, they describe what happens. Examples of scientific laws are the Newton's laws of motion or the law of conservation of energy.

The scientific method may be represented schematically as follows



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Conclusion /Summary

Chemistry is a fascinating science that touches almost every aspect of our lives. It offers a wide range of career opportunities and contributes significantly to various fields, such as healthcare, industry, and space exploration. However, it is crucial to be aware of the potential adverse effects of chemistry and practice safe and responsible handling of chemicals. Understanding scientific methods of analysis empowers us to study and comprehend the world around us in a systematic manner.

EVALUATION:

1. Define chemistry
2. List any three branches of chemistry.
3. Explain why chemistry is referred to as the "mother of all sciences."
4. What are the career prospects for someone studying chemistry? Mention two.
5. Describe the role of chemistry in the medical field.
6. Identify and explain one adverse effect of chemistry on the environment.
7. Outline the steps involved in the scientific method.
8. Give an example of a hypothesis that could be tested in a chemical experiment
9. What is the importance of chemistry in everyday life? Provide one example.
10. How does chemistry contribute to environmental sustainability?
11. Describe how chemistry is applied in the agriculture industry.
12. Explain the difference between a scientific theory and a scientific law
13. Why is it important to practice safe and responsible handling of chemicalss.
14. Name one scientific law and briefly describe it.
15. How does the scientific method empower us to understand the world around us?

WEEK 2

TOPIC: Particulate Nature of Matter I

LESSON OBJECTIVES: By the end of this lesson, Students should be able to:
1. Point out physical and chemical properties of matter;
2. Distinguish between physical and chemical changes;
3. Differentiate between atoms, ions and molecules;
4. Evaluate atomicity
5. List out Dalton's Atomic theory;
6. Modify Dalton's Atomic theory;
7. Illustrate how matter is neither created nor destroyed;
8. Distinguish between elements, compounds and mixtures.

Particulate Nature of Matter:

The particulate nature of matter refers to the idea that all matter is made up of tiny discrete particles. These particles can be atoms, molecules, or ions, depending on the type of substance. Understanding the particulate nature of matter is fundamental in chemistry as it helps explain the behavior and properties of different materials.

Definition of Matter:

Matter is anything that has mass and occupies space. It can exist in three states: solid, liquid, and gas.

Matter is composed of atoms, which are the basic building blocks of all elements and compounds.

Properties of matter

Matter can be identified by its characteristics such as colour, smell, density, reaction, and so on. These characteristics are called properties of matter. There are two properties of matter. They are;

1. Physical properties

2. Chemical properties

Physical Properties of Matter: Physical properties of matter are characteristics that can be observed or measured without changing the chemical composition of the substance. These properties can be seen or felt and will not change the identity of the substance. They can be either;
- intensive property (any feature of matter that does not depend on the amount of the substance present) like boiling point, melting point, density, taste, colour, temperature, refractive index and hardness
- extensive property (any feature of matter that depends on the amount of matter being measured) such as volume, mass and weight.

Chemical Properties of Matter: Chemical properties of matter describe how a substance behaves when it undergoes a chemical change or reaction leading to the formation of a new substance. These properties determine the substance's ability to interact with other substances and form new compounds. Examples include reactivity, combustibility, acidity, and ability to rust.

Changes in matter

From the type of properties of matter, matter can undergo two types of changes when subjected to changes in temperature and pressure. These are physical and chemical changes.

Physical Changes:

Physical changes are alterations in the physical properties of a substance without any change in its chemical composition. These changes are usually reversible. Examples of physical changes include changes in state (solid to liquid to gas), changes in shape or size, and changes in density.

Examples of Physical Changes include

●Melting of ice: Ice changing into water without altering its chemical composition.

●Cutting a piece of paper: The paper changes its shape, but it remains the same substance.

●Dissolving salt in water: Salt particles mix with water but remain as salt chemically.

●Magnetisation and demagnetisation of iron

●Changes of state such as boiling, evaporation, freezing, melting, sublimation and deposition

Chemical Changes:

Chemical changes, also known as chemical reactions, involve the transformation of one or more substances into new substances with different chemical properties. In a chemical change, the chemical composition of the substances involved is altered, and the change is usually irreversible.

Examples of Chemical Changes include

●Rusting of iron: Iron reacts with oxygen and water to form iron oxide (rust).

●Burning of a candle: The candle wax reacts with oxygen in the air, producing carbon dioxide and water vapor.

●Baking a cake: The ingredients undergo chemical reactions to form a completely new substance, the cake.

●Addition of water to quick lime ie slaking of lime

●Combustion of methane

●digestion of food,

●neutralization reactions,

●decaying of organic matter,

●tarnishing of silver, etc.

Differences between physical changes and chemical change

	Physical Changes	**Chemical changes**
Nature of Change	Physical changes involves alterations in the physical state or appearance of a substance ie, no new substance is formed	Chemical change results in the formation of a new substance
Reversibility	Generally reversible	Generally irreversible
Energy changes	Usually involve a change in energy (eg. Heat or temperature) but do not produce or absorb energy.	Chemical changes involve the release or absorption of energy
Conservation of mass	The mass of the substance remains the same	The total mass may change due to the formation of new substances.

Particles of matter (Atoms, Molecules, Ions):

Matter is made up of discrete particles. The following experimental evidences show that matter is made of particles. They are;

i. Brownian movement

ii. Diffusion

iii. Evaporation

iv. Osmosis

v. Sublimation

vi. Tyndal effect

vii. Dispersion

Assignment: Write short notes on the seven evidences that matter is made up of particles

Dalton's Atomic Theory:

Dalton's Atomic Theory, proposed by John Dalton in the early 19th century, laid the foundation for modern atomic theory. It consists of the following key points:

1. All matter is composed of tiny indivisible particles called atoms.
2. These atoms can neither be created nor destroyed during chemical reactions. Ie, Chemical reactions involve the rearrangement of atoms;
3. Atoms of a particular elemnt are exactly alike in every aspect but are different from atoms of elements.
4. Atoms combine in fixed ratios to form compounds.

Daltons theory have been verified experimentally and have given rise to the following laws;

1. The law of conservation of mass (or indestructibility of matter)
2. The law of definite proportions (or constant composition)
3. The law of multiple proportions

The particles of matter are atoms, molecules and ions.

1. **Atoms:** Atoms are defined as the smallest indivisible particle of an element that can ever exist and can take part in a chemical reaction.

Atoms are the building blocks of matter. They are incredibly tiny particles that make up all elements. Each element is composed of a unique type of atom. For example, hydrogen is made up of hydrogen atoms (H), oxygen is made up of oxygen atoms (O), and so on. Atoms consist of a central nucleus, containing protons and neutrons, surrounded by electrons that orbit the nucleus.

2. **Molecules:**This is defined as the smallest particle of an element or a compound which can exist alone. When two or more atoms combine chemically, they form molecules. Molecules can

be made up of the same type of atoms or different types of atoms bonded together. For instance, oxygen gas (O_2) is a molecule composed of two oxygen atoms bonded together, while water (H_2O) is a molecule made up of two hydrogen atoms and one oxygen atom. Molecules are the fundamental units of compounds, which are substances formed when elements chemically react with each other.

The combination of atoms of the same type produces a molecule of ***an element***. Examples are; a molecule of bromine (Br_2), a molecule of phosphorus (P_4), a molecule of sulphur (S_8), and so on. The combination of different types of atoms gives rise to a molecule of ***a compound.*** Examples include; nitrogen(IV) oxide (NO_2), tetraoxosulphate(VI) acid (H_2SO_4), sodium oxide (Na_2O), calcium sulphide (CaS), and so on. Molecules can also be

- Monoatomic Molecules: These are molecules which made up of a single atom. Example; Helium (He), Neon (Ne), Argon (Ar), Sodium (Na)
- Diatomic Molecules: These are molecules which made up of two atoms. Example; nitrogen (N_2), iodine (I_2), hydrogen (H_2), fluorine (F_2)
- Triatomic Molecules: These are molecules which made up of three atoms. Example; ozone (O_3) Similarly, carbon dioxide (CO_2), nitrogen dioxide (NO_2).
- Tetraatomic molecules: These molecules consist of four atoms. For instance, P_4 is a tetraatomic molecule of phosphorus.
- Polyatomic molecules: These are molecules which made up of two, three or more atoms. Sulphur (S_8), methane (CH_4), nitric acid (HNO_3) are all examples of polyatomic molecules.

3. Ions: An ion is an electrically charged particle that results from the loss or gain of electrons by an atom or molecule. There are two types of ions;

i. **The positively charged ion or cation:** when an atom loses one or more electrons, it becomes positively charged and is called a **cation.** Examples are hydrogen ion (H^+), lithium (Li^+), magnesium ion (Mg^{2+}), aluminum ion (Al^{3+}), iron(III) ion (Fe^{3+}), ammonium ion (NH_4^+), oxonium ion (H_3O^+) etc.

ii. **The negatively charged ion or anion:** When an atom gains one or more electrons, it becomes negatively charged and is called an **anion.** Examples are iodide ion (I^-), chloride (Cl^-), oxide ion (O^{2-}), sulphide ion (S^{2-}), tetraoxomanganate(VII) ion (MnO_4^-), hydroxide ion (OH^-), etc. Ions play a crucial role in chemical reactions, as they can combine with other ions or molecules to form new compounds.

Atomicity

Atomicity of an element is the number of atoms in a molecule of an element. The atomicity of different molecules are shown below

S/N	Name	Symbol	Atomicity
1.	Group one elements e.g. sodium	Na	1 or monoatomic
2.	Noble gases e.g. neon	Ne	1
3.	Hydrogen molecule	H_2	2 or diatomic
4.	Oxygen molecule	O_2	2
5.	Ozone	O_3	3 or triatomic
6.	Phosphorus	P_4	4 or polyatomic
7.	Sulphur	S_8	8 or polyatomic

Modification of Dalton's Atomic Theory:

Over time, advancements in scientific knowledge led to the modification of Dalton's Atomic Theory. Some of the modifications include:

1. **Subatomic Particles:** The discovery of subatomic particles (protons, neutrons, and electrons) revealed that atoms are not indivisible and have internal structure.

2. **Isotopes:** Isotopes are atoms of the same element with different numbers of neutrons. Dalton's theory did not account for isotopes.

3. Nuclear Reactions: Certain nuclear reactions can change one element into another, implying that atoms can be transformed under specific circumstances.

Classes of Matter

Matter may be classified into three types; **Elements, Compounds, and Mixtures.**

Elements: An element is defined as a substance which cannot be broken down into simpler substances by a chemical process. Elements are pure substances that consist of only one type of atom. They cannot be broken down into simpler substances by ordinary chemical means. Each element is represented by a unique chemical symbol, such as H for hydrogen, O for oxygen, and Au for gold. Elements are divided into:

1. **Metals** have the maximum number of elements in a Periodic Table. They include alkaline metals, alkaline earth metals, transition metals and others. They possess the following characteristics; they are malleable, shiny, sometimes magnetic, conduct electricity. Examples are sodium, beryllium, aluminum, iron.

2. **Non-metals** are found on the right side of the periodic table. They possess the following characteristics; they are typically not conductive, not malleable, not shiny (dull), not magnetic. Examples of non-metals include nitrogen, sulphur, argon, iodine.

3. **Metalloids** possess characteristics of both metals and non-metals. Examples of metalloids are: boron, silicon, arsenic

There exists a total of one hundred and eighteen(118) known elements, only the first ninety-eight (98) are known to occur naturally on Earth. Elements that do not occur naturally on Earth are the synthetic products of man-made nuclear reactions. Eighty (80) out of the 98 naturally-occurring elements are stable; while the rest are radioactive.

Compounds: Compounds are substances composed of two or more elements chemically bonded together in fixed ratios. They have distinct properties different from the elements that compose them. For example, water (H_2O) is a compound composed of hydrogen and oxygen. The properties of water are different from that of hydrogen and oxygen which it is formed from. Other examples of compounds and the elements its made up of are shown in the table below

Mixtures: Mixtures are combinations of two or more substances physically combined together

Compounds	**Elements present**
Carbon (iv) oxide	Carbon and Oxygen
Sand(Silica)	Silicon and oxygen
Calciumtrioxocarbonate(iv) or Chalk	Calcium, carbon and oxygen
Sugar (glucose)	Carbon, hydrogen and oxygen
clay	Aluminium, silicon, oxygen, and hydrogen
Sodium chloride(common salt)	Sodium and chlorine
Rust	Iron, oxygen and hydrogen.

but not chemically bonded. Mixtures can be either homogeneous (uniform throughout) or heterogeneous (non-uniform). Examples of mixtures include air (a mixture of gases), saltwater (a mixture of salt and water), and sand (a mixture of various minerals). Unlike compounds, mixtures can be separated by physical means. Other examples of mixtures include

Mixture	**Components**
Ink	coloured dyes
Seawater:	salt and water
Gunpowder	sulfur, potassium nitrate and carbon
Air	oxygen, nitrogen, carbon dioxide, rare gases, dust and water vapour

Crude oil	Kerosene, diesel, bitumen, petroleum ether, petrol, and so on

Differences between elements and compounds

Elements	Compounds
Elements are made up of only one type of atom	compounds consist of atoms of different elements.
Elements cannot be decomposed further by chemical means	compounds can be broken down into their constituent elements.
Elements have specific chemical properties based on their atomic structure	compounds have distinct properties resulting from the interactions between their constituent elements.
Elements are represented by chemical symbols	compounds are represented by chemical formulas.
Have specific physical and chemical properties.	Have their own set of physical and chemical properties different from their constituent elements.

Differences between compounds and mixtures

Mixtures	Compounds
Combination of two or more substances physically mixed together without forming new chemical bonds	Chemical substances composed of two or more elements in fixed ratios by chemical bonding
Can be homogeneous or heterogeneous	Always homogeneous
Composition varies, substances retain their individual properties	Fixed composition, consistent throughout
Components can be separated by physical means (filtration, distillation, etc.)	Can only be separated through chemical reactions or breaking bonds
No energy changes upon mixing	Energy changes occur during formation and breaking of chemical bonds

Components may have different physical states (solid, liquid, gas)	Homogeneous compounds have uniform physical states

EVALUATION

1. Describe a physical change and provide two examples. Explain why these examples are considered physical changes.
2. Define a chemical change and provide two examples. Explain why these examples are considered chemical changes.
3. Define an atom, a molecule, and an ion. Provide an example of each.
4. Explain how ions are formed from atoms. Provide an example of a common ion.
5. Describe the difference between a diatomic molecule and a polyatomic molecule, giving an example of each.
6. Define atomicity in relation to molecules. Provide examples of molecules with different atomicities.

Week Three

Topic: Particulate Nature of Matter II

Instructional Objectives: By the end of this lesson, the students should be able to

1. Compare the distinguishing properties of solids, liquids and gases.
2. Design the changes applicable to the states of matter;
3. Award an understanding of the random motion of particles in a suspension as evidence for the kinetic particle model of matter;
4. Survey diffusion as a phenomena supporting kinetic theory;
5. Justify the dependence of rate of diffusion on molecular mass;
6. Discuss atomic structure
7. State the rule on how the electrons are arranged in the atom.

States of Matter

Matter can exist in three major states. They are the solid, liquid and gaseous states. The various states which matter exist depends on the way the particles of the matter are arranged as well as the type of forces holding the particles together.

1. Solids: Particles of solids are closely packed in a regular arrangement. They have a fixed shape and volume and vibrate in their positions. Particles of a solid have definite shapes and definite volumes.

2. Liquids: Particles of liquids are close together but not as tightly packed as in solids. They have a definite volume but no definite shape (they assume/take the shape of their container) and move more freely than particles in solids.

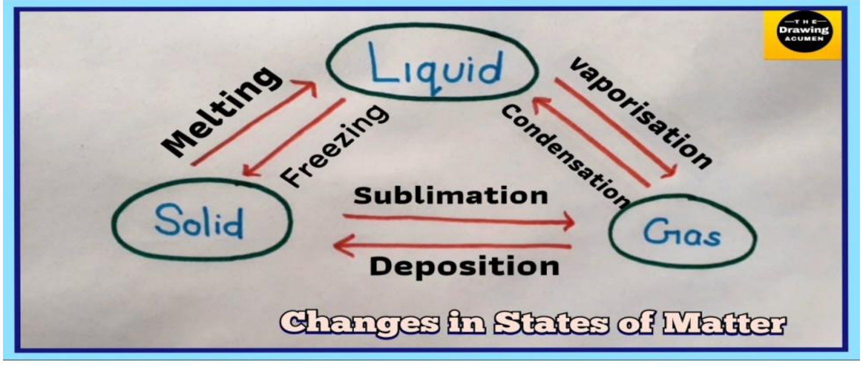
3. Gases: Particles of gases are far apart, move rapidly, and have no fixed shape or volume. They spread to fill the entire available space. Particles of gases have no definite shapes and no definite volumes.

Matter can be changed from one state to another when it is subjected to external factors such as increase or decrease in temperature during heating or cooling. During changes of state, the number of particles remains constant, but their arrangement and movement change.

The following are changes that can occur in the state of a matter.

1. **Melting:** this is a process of heating a solid to a certain temperature (known as its melting point) at which it losses its fixed position and is said to have liquefied (or melted).
2. **Boiling:** this occurs when the saturated vapour pressure of a heated liquid equals the prevailing atmospheric pressure. The temperature at which this occurs is called the **boiling point**.
3. **Evaporation or vaporisation:** this is a process which occurs when the particles in a substance have been energised by heating or by an increase in temperature which leads to the escape of some or all of the liquid molecules in it.
4. **Condensation** is a process where vapour loses its kinetic energy to a cooler body thereby becoming a liquid again.
5. **Freezing:** This occurs when a liquid loses its heat energy to its surrounding continuously until the liquid particles attained a fixed position as of that of a solid. The point at which this happens is called the freezing point and this is also know as the melting point of that same substance (if pure).
6. **Sublimation:** this is a situation where a solid melts into a gas without going through the liquid state.
7. **Deposition:** this is the process by which molecules go directly from the gas phase into the solid phase

The different changes in state of matter can be represented as follows



Kinetic Theory of Matter

This theory explains the behavior of matter based on the motion of particles. It states that particles in matter are in constant motion.

Assumptions of the Kinetic Theory of Gases

This theory describes the behaviour of an ideal or perfect gas. It states the following:

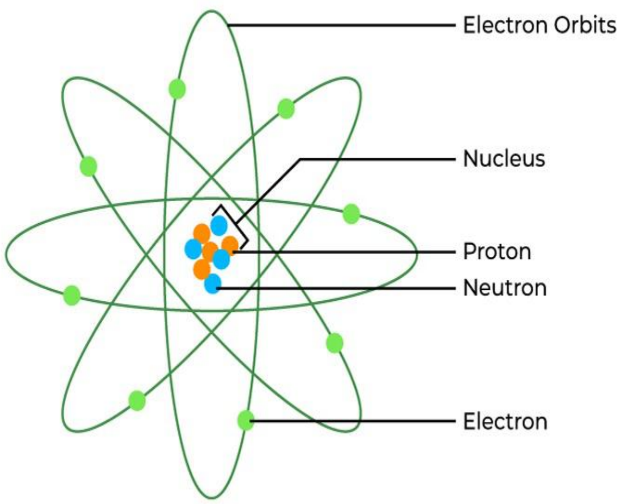
1. The gas molecules move randomly in straight lines colliding with one another and with the walls of the container.
2. The collision of the gas molecules are perfectly elastic
3. The actual volume occupied by the gas molecules are negligible.
4. The cohesive forces between the gas molecules are negligible.
5. The temperature of the gas is a measure of the average kinetic energy of the gas particles.

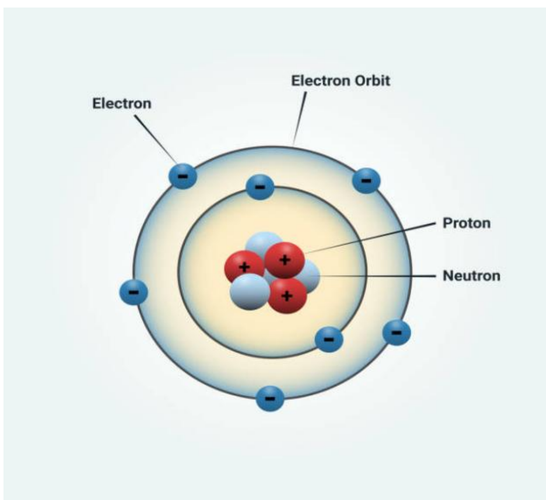
Phenomena Supporting Kinetic Theory

1. **Brownian Movement or Motion:** This is the irregular movement caused as a result of the bombardment of suspended solid particles by the surrounding molecules of the liquid medium.
2. **Diffusion:** This is a movement from a region of higher concentration to a region of lower concentration.
3. **Osmosis:** This is the movement of water molecules through a semi-permeable membrane from a region where they are in higher concentration to one of the lower concentration.
4. **Expansion and Contraction:** Matter expands when heated and contracts when cooled. This can be explained by the increased or decreased kinetic energy of particles.

Atomic Structure: Constituents of Atoms and Electron Arrangement

The atom (The basic unit of matter) consists of three fundamental particles, electron, protons and neutrons. An atom consists of a central nucleus (protons and neutrons) with electrons orbiting around it in orbitals or shells. The structure of the atom is represented below





Protons: Positively charged particles found in the nucleus of an atom. They determine the element's identity.
Neutrons: Neutral particles found in the nucleus. They add mass to the atom but do not significantly affect its chemical properties.
Electrons: Negatively charged particles orbiting the nucleus in energy levels or shells. They determine the atom's chemical behavior and are involved in bonding.

The properties of the sub atomic particles are summarized below

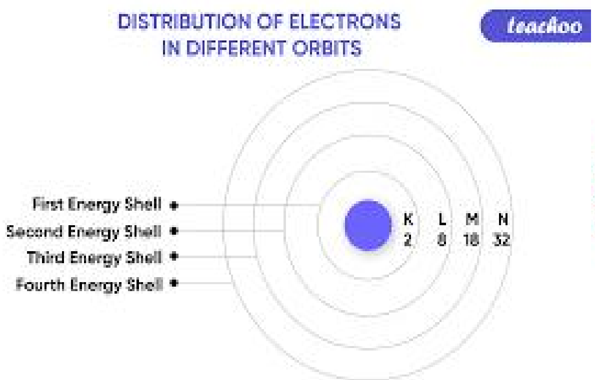
Particles	Location in the atom	Relative mass	Relative charge	Absolute mass(g)	Absolute charge (coulombs)	symbol
Electron	Shells or orbitals	1/1840 (negligible)	-	9.11×10^{-28}	1.602×10^{-19}	e^-
Proton	Nucleus	1	+	1.673×10^{-24}	1.602×10^{-19}	P^+
Neutron	Nucleus	1	Nil (No charge)	1.673×10^{-24}	Nil (ie no charge)	n

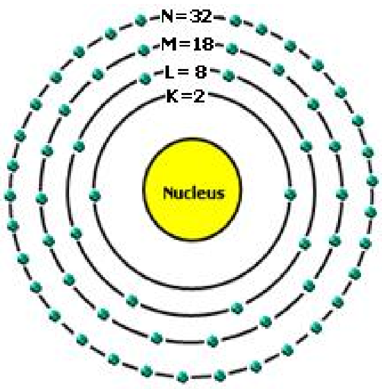
Electron Arrangement in Atoms

Electrons are arranged around the nucleus in shells called energy levels.Electrons occupy energy levels in a specific order. The innermost shell is filled first before moving to higher energy levels. The arrangement determines an element's properties and reactivity.

Bohrs model of the atom

Neil Bohr proposed that the atom is spherical in shape with a small centre called the nucleus. The nucleus houses the proton and the neutrons. Outside the nucleus are the electrons arranged in shells and revolving at various distances from the nucleus. These shells(or orbitals) are denoted by the letters K, L, M, N, O, representing energy levels of 1, 2, 3 ,4 respectively.
The shell closet to the nucleus has the lowest energy. The energy of the shells increase as the distance from the nucleus increase.
The K shell has the lowest energy followed by the L shell and so on.
Bohr also showed that only certain numbers of electrons can stay in a particular shell and this is given by the formular $\mathbf{2n^2}$ where n is the shell number.
Therefore, for the
K shell, n =1, $2 \times 1^2 = 2$ electrons
L shell, n = 2, $2 \times 2^2 = 8$ electrons
M shell, n = 3, $2 \times 3^2 = 18$ electrons
N shell, n = 4, $2 \times 4^2 = 32$ electrons
This means that 2, 8, 18, 32 electrons are the maximum number of electrons that can be found in the K, L, M, N shells respectively.





To fill the shells of the various atoms of elements with electrons, electrons will move into the shell of the lowest energy first until it is completely filled before before electrons can move into a shell of the next higher energy. The arrangement of electrons into shells of the atom is called <u>**electronic configuration**</u>**.**
Example: write the electronic configuration of the following atoms: Na, Mg and
The electronic configuration of the first 20 elements is shown below

Element	Symbol	Number of electrons	1^{st} shell	2^{nd} shell	3^{rd} shell	4^{th} shell	Electron configuration
Hydrogen	H	1	1				1
Helium	He	2	2				2
Lithium	Li	3	2	1			2,1
Beryllium	Be	4	2	2			2,2
Boron	B	5	2	3			2,3
Carbon	C	6	2	4			2,4
Nitrogen	N	7	2	5			2,5
Oxygen	O	8	2	6			2,6
Fluorine	F	9	2	7			2,7
Neon	Ne	10	2	8			2,8
Sodium	Na	11	2	8	1		2,8,1
Magnesium	Mg	12	2	8	2		2,8,2
Aluminium	Al	13	2	8	3		2,8,3
Silicon	Si	14	2	8	4		2,8,4
Phosphorus	P	15	2	8	5		2,8,5
Sulphur	S	16	2	8	6		2,8,6
Chlorine	Cl	17	2	8	7		2,8,7
Argon	A	18	2	8	8		2,8,8
Potassium	K	19	2	8	8	1	2,8,8,1
Calcium	Ca	20	2	8	8	2	2,8,8,2

EVALUATION

1. Describe the basic structure of an atom, including its subatomic particles.
2. Explain the concept of electron shells (energy levels) in an atom.
3. Define electron configuration
4. A.Write the electronic configuration of the following elements; Magnesuim, Neon, Sulphur, Sodium, chlorine. B. Represent the electron distribution in the atom using a diagram.

Week Four

Topic: Particulate Nature of Matter III

Instructional Objectives: By the end of this lesson, the students should be able to

1. Disprove Bohr's model of the atom;

2. Briefly discuss the following; atomic number, mass number, C-12 scale, relative atomic mass and relative molecular mass based on C-12 isotope;

3. Find the relative atomic masses and the relative molecular masses of substances;

4. Discuss isotopy

5. Classify isotopes as being radioactive and non-radioactive along with their uses;

6. Highlight the relative charges and approximate relative masses of protons, neutrons and electrons.

It was discovered by Lord Rutherford in 1911 that the atom is electrically neutral. This is because the number of protons (positive charge) in the nucleus of an atom must equal the number of electrons that revolves round it.

- In 1913, Niels Bohr suggested that the electrons in an atom revolves round the nucleus along certain imaginary circular path known as <u>**orbits or shells**</u> located at various distances from the nucleus. These orbits are named as **K, L, M, N, O, P, Q**.
- the electrons has a quantized (fixed value) energy which emits in form of radiation. But Bohr's model was unable to account for more complicated spectra lines other than that of hydrogen.

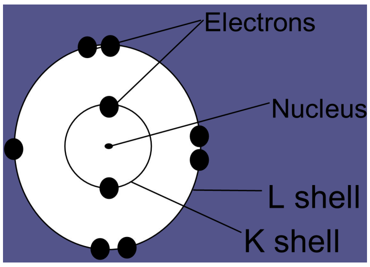


Figure: structure of fluorine atom

Atomic Number, Mass Number, Isotopes, C-12 scale, relative atomic mass and relative molecular mass based on C-12

1. **Atomic number (Z):** *(or proton number)* of an element is the number of protons in the nucleus of an atom of that element. E.g. neon $_{10}Ne$.

Atomic number (Z) = number of protons

2. **Mass number (A):** is the sum of the protons and neutrons in an atom of an element. E.g. neon $^{20.2}Ne$.

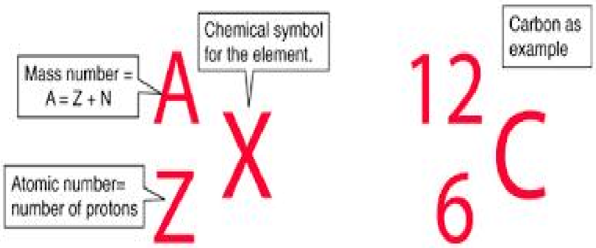
The mass of the atom could be said to be concentrated in the small centre called the nucleus. Since the masses of both proton and neutron are thesame, the mass of the atom could be said to be the total number of protons and neutrons.

Mass number (A) = number of protons + number of neutrons

The atom is electrically neutral. Hence, the number of positively charged protons in the nucleus must be equal to the number of negatively charged electrons outside the nucleus.

Therefore, **number of protons = number of electrons**

The atom of an element can be represented as follows;



Example: Consider the following atoms of different element, $^{12}_{6}C$, $^{23}_{11}Na$, determine the atomic no, mass no, no of neutrons, no of protons, no of electrons.

Solution:

$^{12}_{6}C$, Atomic no = no of protons = 6

Mass no = 12

no of neutrons = 12 - 6 = 6

no of protons = 6

no of electrons = no of protons = 6

$^{23}_{11}Na$, Atomic no = no of protons = 11

Mass no = 23

no of neutrons = 23 - 11 = 12

no of protons = 11

no of electrons = no of protons = 11

Example: a neutral atom of an element has 17 electrons and 18 neutrons. Determine (a)the number of protons (b) the mass number of the atom (c) the name of the atom?

Solution

(a) The atom is electrically neutral, therefore, no of protons = no of electrons

No of electrons = 17, therefore no of protons = 17

(b) Mass number = no protons + no of neutrons

= 17 + 18 = 35

(c) the name of the atom is Chlorine.

Example: for the following cation and anion, $_{13}{}^{27}Al^{3+}$ and $_{17}{}^{35}Cl^{-}$ determine (a) atomic no, (b) mass no, (c) no of electron (d) no of neutron

Solution:

	$_{13}{}^{27}Al^{3+}$	$_{17}{}^{35}Cl^{-}$
(a) atomic no	13	17
(b) mass no	27	35
(c) no of electron	10 (it has lost 3 electrons)	18(it has gained one electron)
(d) no of neutron	27-13 = 14	35-17 = 18

ISOTOPES

Isotopes are atoms of the same element having the same atomic number (no of protons) but different mass numbers. The difference in masses is due to different number of neutron, therefore isotopes have different physical properties but thesame chemical properties.

The existence of isotopes is called **isotopy.** Hence, **Isotopy** is the existence of atoms of an element with different mass numbers but the same atomic number.

Examples of isotopes include

(i) **Hydrogen** has three isotopes; $^{3}_{1}H$(tritium), $^{2}_{1}H$ (deuterium), $^{1}_{1}H$(protium)

(ii) **Carbon** has three isotopes; $^{14}_{6}C$, $^{13}_{6}C$, $^{12}_{6}C$

(iii) **Oxygen** has two isotopes; $^{18}_{8}O$, $^{16}_{8}O$

(iv) **Chlorine** has two isotopes $^{35}_{17}Cl$, $^{37}_{17}Cl$

Similarities amongst isotopes of an element;

A. They posses the same number of proton.

B. They have the same chemical properties.

Differences amongst isotopes of an element;

a. They have different mass number

b. They have slightly different physical properties

Non-radioactive (Stable) Isotopes

These isotopes have a stable proton-neutron ratio with no sign of decay. If an atom possesses too few or too many neutrons, it becomes unstable and tends to disintegrate.

Uses of Non-radioactive Isotopes

A. It is employed in environmental and ecological experiments

B. Non-radioactive isotopes are used for determining geological facts about materials, like their age.

Radioactive Isotopes have an unstable combination of protons and neutrons. These isotopes decay and emit radiation. These isotopes are grouped according to their creation process: long-lived, cosmogenic, anthropogenic and radiogenic.

Uses of Radioactive Isotopes

A. Radioactive isotopes are useful in agriculture, food industry, pest control, archeology and medicine.

B. Radiocarbon dating, which measures the age of carbon-bearing items, uses a radioactive isotope known as carbon-14.

Calculation on Isotopy

Calculate the relative atomic mass of carbon whose istopes are occurring 98.9 % in nature and occurring 1.1 % in nature.

Solution

Relative atomic mass = summation of the multiplication of the individual isotopes by their relative abundance in nature.

= (12 x 98.9 ÷ 100) + (13 x 1.1 ÷ 100) = (12 x 0.989) + (13 x 0.011)

= 11.868 + 0.143

= 12.011

Thus the relative atomic mass of carbon is 12.01

C-12 Scale is a scale that uses the ^{12}C isotope as the standard for comparing the atomic masses of the other elements. **OR**

It is the mass reference scale in which the atom of the ^{12}C isotope is taken to have twelve units of atomic mass.

Relative atomic mass, A_r, **of an element** is the number of times the average mass of one atom of that element is heavier than one-twelfth(1/12)th of the mass of one atom of carbon – 12.

Relative molecular mass, M_r, **of an element or a compound** is the number of times the average mass of one molecule of it is heavier than one-twelfth(1/12)th of the mass of one atom of carbon – 12.

Calculations on relative molecular mass

1.Find the relative atomic masses of an atom boron (B) and nitrogen (N).

Solution

A. The relative atomic mass of boron (B) is = 10.8

B. The relative atomic mass of nitrogen (N) is = 14.0

3. Find the relative molecular masses of a molecule of oxygen (O_2) and that of calcium fluoride (CaF_2).

Solution

a.Relative molecular mass, M_r, of a molecule oxygen (O_2)?

Relative atomic mass, A_r, of oxygen = 16

M_r of O_2 = summation of all the atomic masses present

= A_r of oxygen × 2

= 16 × 2 = **32**

Thus the relative molecular mass of a molecule of oxygen is **32**

b.Relative molecular mass, M_r, of CaF_2?

A_r of calcium = 40; A_r of fluorine = 19.0

M_r of CaF2 = summation of all atomic masses present

= A_r of calcium + (A_r of fluorine × 2)

= 40 + (19.0 × 2) = 40 +38 = **78**

Thus the relative molecular masses of a molecule of calcium fluoride (CaF_2) is **78**

Evaluation

1. State the limitation of Thomson's Model of the atom
2. Calculate the relative molecular masses of the following molecules

a. $K_2Cr_2O_7$

b. H_2SO_3

c. $(NH_4)_2SO_4$

d. $Al(OH)_3$ (take K=39, Cr= 52, O= 16, H=1, S=32, N= 14, Al =27)

Week Seven
Topic: Chemical Symbols and Formulae I
Instructional Objectives: By the end of this lesson, the students should be able to

1. Solve for the oxidation number of substances;

2. Write the chemical formula of named substances;Calculate the empirical and molecular formulae of compounds;

3. Isolate the symbols of the first 20 elements based on their atomic numbers;

4. Establish the relationship proton number and simple structure of atoms;

Chemical Symbols of Elements

A chemical symbol is defined as a sign which consists of one or two letters used to represent an atom of an element.

In order to represent atoms of elements, the first letter (In capital letter) of the name of the element is used. For instance, N for Nitrogen. Where different elements have the same letter beginning their names, to differentiate the atoms of the various elements, the element that was discovered first is assigned the first letter of its name only, the other elements would still retain the first letter of their names written in capital letters together with another letter from its name written in small letter. For instance, Ne for Neon, Ni for Nickel, etc. The symbols of some other atoms of elements are derived from their latin name, for instance, the latin name of sodium is Natrium, the symbol of sodium is represented as Na.

Table 9.1 Elements having first letter as symbol

Element	Symbol	Element	Symbol
Boron	B	Oxygen	O
Carbon	C	Phosphorus	P
Fluorine	F	Sulphur	S
Hydrogen	H	Vanadium	V
Iodine	I	Uranium	U
Nitrogen	N	Yttrium	Y

Table 9.3 Elements with same two letters

Element	Symbol	Element	Symbol
Argon	Ar	Calcium	Ca
Arsenic	As	Cadmium	Cd
Chlorine	Cl	Magnesium	Mg
Chromium	Cr	Manganese	Mn

Table 9.4 Greek or Latin name of elements

Element	Latin Name	Symbol
Sodium	Natrium	Na
Potassium	Kalium	K
Iron	Ferrum	Fe
Copper	Cuprum	Cu
Silver	Argentum	Ag
Gold	Aurum	Au
Mercury	Hydrargyrum	Hg
Lead	Plumbum	Pb
Tin	Stannum	Sn
Antimony	Stibium	Sb
Tungsten	Wolfram	W

Chemical formulae

A chemical formula can be defined as a collection of two or more symbols to represent one molecule of the compound.

Formula are used to represent molecules of compounds, in a formula there must be at least two symbols.

Valency

Valency is the combing power of elements. It is the power with which atoms are combining to form molecule of the compound. The valency of atoms could range from 0 to 7 as deduced experimentally. The valency of different elements are shown below

Table of Valences (valencies) www.vaxasoftware.com

METALS

Name	Symbol	Valence (valency) (+)
Lithium	Li	
Sodium	Na	
Potassium	K	
Rubidium	Rb	1
Cesium	Cs	
Francium	Fr	
Silver	Ag	
Amonium	NH_4^+	
Beryllium	Be	
Magnesium	Mg	
Calcium	Ca	
Strontium	Sr	2
Barium	Ba	
Radium	Ra	
Zinc	Zn	
Cadmium	Cd	
Aluminium	Al	3
Copper	Cu	1 2
Mercury	Hg	
Gold	Au	1 3
Chromium	Cr	
Manganesium	Mn	
Iron	Fe	2 3
Cobalt	Co	
Nickel	Ni	
Tin	Sn	
Lead	Pb	2 4
Platinum	Pt	

NON METALS

Name	Symbol	Valence (valency) (+)	Valence (valency) (−)
Hydrogen	H	1	−1
Fluorine	F		−1
Chlorine	Cl		
Bromine	Br	1 3 5 7	−1
Iodine	I		
Oxygen	O	−2 (−1)	
Sulfur	S		
Selenium	Se	4 6	−2
Tellurium	Te		
Nitrogen	N	1 3 5 (2 4)	−3
Phosphorus	P		
Arsenic	As	3 5	−3
Antimony	Sb		
Boron	B	3	
Bismuth	Bi	3 5	
Carbon	C	2 4	−4
Silicon	Si	4	−4

Name	Symbol	Valence (valency) (+)	Valence (valency) (−)
Manganese	Mn	* 4 6 7	
Chromium	Cr		
Molybdenum	Mo	6	
Tungsten	W		

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Element	Atomic Number	Valency
Hydrogen	1	1
Helium	2	0
lithium	3	1
Beryllium	4	2
Boron	5	3
Carbon	6	4
Nitrogen	7	3
Oxygen	8	2
Fluorine	9	1
Neon	10	0
Sodium	11	1
Magnesium	12	2
Aluminum	13	3
Silicon	14	4
Phosphorus	15	3
Sulphur	16	2
Chlorine	17	1
Argon	18	0
Potassium	19	1
Calcium	20	2
Scandium	21	3
Titanium	22	2,3,4
Vanadium	23	2,3,5,4

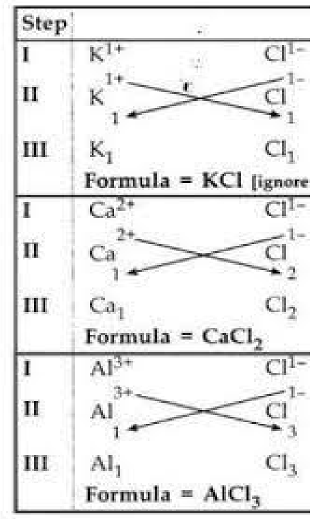
Writing formulae of molecules of compounds

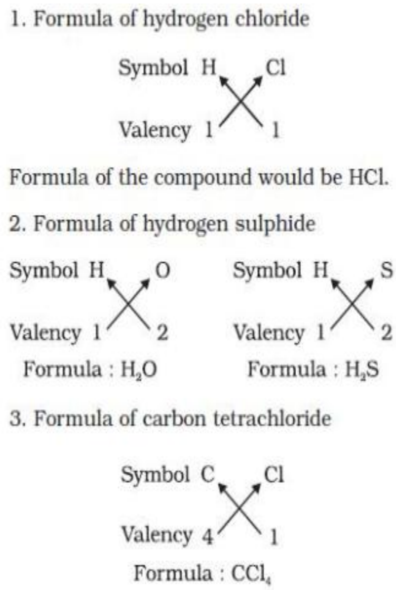
1. Compounds may be classified into two; Binary compounds and Trinary compounds

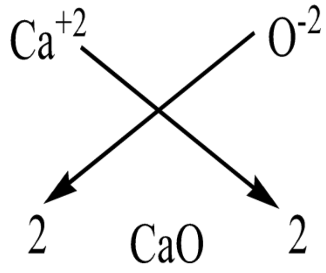
Binary compounds: This contains two different elements chemically combined together.

Writing of binary compounds using valencies

The following steps can be employed in writing the chemical formulae of binary compounds. An easy way to write a correct formula for an ionic compound is to use the crisscross method. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped. Shown below are examples of criss cross method for writing different compounds.







Naming of binary compounds

In naming binary compounds, the following rules are applied

1. The names of the compound must reflect the names of the elements present in the compound. (note, there are exceptions, for instance water, H_2O and ammonia, NH_3)
2. The name of the first element (usually a metal) is written first followed by the name of the second element but changing the ending to end -ide.
3. If the first element have variable valencies, the valency used to form the compound is written as a roman figure, and is attached to the name of the first element before it is followed by the name of the second ending in -ide.

Example. Write the names of the following compounds, KCl, $CaCl_2$, $AlCl_3$, H_2S. HCl. CaO, $FeCl_2$, $FeCl_3$

Solution;

1. KCl = Pottasium chloride
2. $CaCl_2$ = Calcium chloride
3. $AlCl_3$, = Aluminium chloride
4. H_2S. = Hudrogen sulphide
5. HCl. = Hydrogen chloride
6. CaO = Calcium oxide
7. $FeCl_2$, = Iron(II) chloride
8. $FeCl_3$ = Iron (III) chloride

Try these: Write a chemical formulae between aluminium and oxygen, magnesium and sulphur, iron and chlorine, copper and oxygen, Fe^{2+} and sulphur, Fe^{3+} and sulphur. What are the names of the compounds formed?

Trinary Compounds

These are compounds containing three different elements only. They are composed of a metal or hydrogen ion and an acid radical.

A radical is a group of different atoms that react as a single unit. A radical is an ion since it has a charge. The valency of a radical ion is the number of charges on the ion; for instance, the valency of PO_4^{3-} is 3. Examples of radicals are shown below

Radical ion	Name	valency
ClO^-	Oxochhlorate(I) ion	1
ClO_3^-	Trioxochlorate (V) ion	1

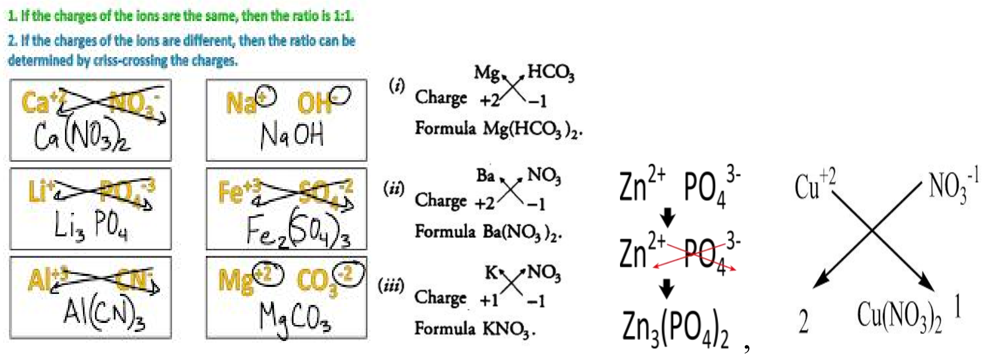
NO_2^-	Dioxonitrate (III) ion	1
NO_3^-	Trioxonitrate(V) ion	1
SO_3^{2-}	Trioxosulphate(IV) ion	2
SO_4^{2-}	Tetraoxosulphate(IV) ion	2
CO_3^{2-}	Trioxocarbonate(IV) ion	2
PO_3^{3-}	Trioxophosphate(III) ion	3
PO_4^{3-}	Tetraoxophosphate(V) ion	3
$Cr_2O_7^{2-}$	Heptaoxodichromate(VI) ion	2
CrO_4^{2-}	Tetraoxochromate(VI) ion	2
MnO_4^-	Tetraoxomanganate(VII) ion	1
OH^-	Hydroxide ion	1
NH_4^+	Ammonium ion	1

Writing the trinary compounds from radicals

This can happen in three ways

1. Metallic and ammonium ions can combine with the negative radicals to form salts
2. Hydrogen ions can combine with the negative radicals to form acids
3. Metallic and ammonium ions can combine with the negative radicals to form bases.

Examples;



Naming of Trinary compounds

Trinary compounds can be named in same way with binary compounds. First name the metal or positive ion followed by the negative radical.

Examples;

$Ca(NO_3)_2$ = Calcuim trioxonitrate(v)

$Fe_2(SO_4)_3$ = Iron (III) tetraoxosulphate (VI)

$MgCO_3$ = Magnesium trioxocarbonta (IV)

$NaOH$ = Sodium hydroxide

Li_3PO_4 = Lithium tetraoxophosphate (v)

$Al(CN)_3$ = Aluminium cyanide

$Mg(HCO_3)_2$ = Magnesium hydrogen carbonate

KNO_3 = Potassium trioxonitate (v)

$Cu(NO_3)_2$ = Copper (II) trioxonitrate (V)

Oxidation Number/State

Oxidation number is the charge an atom would have if electrons were transferred completely. Or The oxidation number of an atom is defined as the charge acquired by the atom when they form ionic bonds with atoms of another element i.e. heteroatoms in a molecule.

Rules for assigning oxidation numbers

a. The oxidation number of an element in its elemental form is 0. For instance, the oxidation number of Na, H_2, Mg, O_2 are 0

b. The oxidation number of a monoatomic ion is its charge. For instance, the oxidation number of Ca^{2+}. O^{2-}, Al^{3+}, H^+ is +2, -2, +3 and +1 respectively.

c. Oxygen usually has an oxidation number of -2 in compounds, except in peroxides where it is -1.

d. Hydrogen usually has an oxidation number of +1 in compounds, except in hydrides where it is -1.
e. The sum of oxidation numbers in a neutral compound is equal to 0.
f. The sum of oxidation number in a radical is equal to the charge on the radical.

To determine the oxidation number of an atom we need to follow the following mentioned steps:

Step 1: Assume the oxidation number of the atom to be X which you need to calculate,

Step 2:Mention the oxidation state of other bonded atoms and multiply it with the number of such atoms present in one molecule

Step 3: Write the oxidation number of all the atoms in the molecule in a linear sum format and equate it to the overall charge of the molecule.

Step 4: Solve for X.

Solved examples:

1. Calculate the Oxidation Number of Sulphur in H_2SO_4

Solution:

Assume the oxidation number of sulphur to be x

Step 2: The oxidation number for Hydrogen is +1 and for O is -2.

Step 3: Since the overall charge on the molecule is 0, therefore $2(+1) + X + 4(-2) = 0$

Step 4: $2 + X – 8 = 0 \Rightarrow X – 6 = 0 \Rightarrow X = +6$

Hence, the oxidation number of Sulphur in H_2SO_4 is **+6**

2. **Determine the oxidation number of Carbon in CO**

CO is a neutral compound, so the sum of its oxidation number is 0

$X + (-2) = 0$

$X – 2 = 0$; $X =$ **+2**

Hence the oxidation state of the carbon atom in CO is **+2**.

3. Oxidation state of chlorine in KCl

KCl is neutral, and so net charge = 0

Oxidation state of KCl = Oxidation state of potassium + oxidation state of chlorine =0.

Oxidation state of potassium = +1

Oxidation state = +1 + x = 0: x = -1

Oxidation state of chlorine in KCl = **-1**

4. Oxidation number of manganese in permanganate ion MnO_4^-

The charge on the permanganate ion is -1

Oxidation state of permanganate ion = Oxidation state of manganese + 4 x oxidation state of oxygen = -1.

Oxidation state of oxygen = -2

Oxidation states → x + (4 x-2) = -1: x = +7

Oxidation state of manganese = **+7**

5. **Cl in Cl_2O:** Cl_2O is neutral, and so net charge = 0.

Net oxidation state of Cl_2O = 2 x Oxidation state of chlorine + 1x Oxidation state of oxygen = 0.

Oxidation state of oxygen = -2.

Oxidation states → 2 x + (-2) = 0: x = +1

Oxidation state of chlorine in Cl_2O= 2/2 = +1

Oxidation state of chromium in dichromate anion.

6. Chromium in Dichromate ion $Cr_2O_7^{2-}$.

$Cr_2O_7^{2-}$ = -2

2Cr + (7*-2) = -2: Cr= +6

Therefore, Oxidation state of chromium = **+6**

Try these:

1. Calculate the oxidation number of nitrogen in ammonium nitrate i.e. NH_4NO_3
2. Mn in $KMnO_4$,
3. Na in NaCl,
4. Mg in MgO
5. Cl in HCl
6. in SO_4^{2-}
7. S in $Na_2S_2O_3$
8. Na in Na_2CO_3
9. C in CN^-
10. N in NH_3

Empirical and molecular formula

Empirical Formula: The empirical formula is the simplest formula for a compound which expresses the ratio of atoms of the various elements present in the compound. It is defined as the ratio of subscripts of the smallest possible whole number of the elements present in the formula. This formula gives information about the ratio of the number of atoms in the compound.

Molecular Formula: The molecular formula specifies the actual number of each type of atom in a molecule. For example, the molecular formula of benzene is C_6H_6, i.e., it is composed of six carbon and six hydrogen atoms. The molecular formula helps in determining whether a chemical compound is a binary compound, ternary compound, quaternary compound, or has even more elements depending upon the number of elements in a molecule. A molecular formula is always a multiple of the empirical formula, where an empirical formula for a chemical compound is defined as a simple expression that represents the ratio of the elements in the compound. For example, the molecular formula of hydrogen peroxide is H_2O_2, whereas its empirical formula is HO. We need the molar mass of a compound to find the molecular formula of a compound, and it is often derived after obtaining the empirical formula. Though molecular formulae are simple and easy to understand, they lack the knowledge concerning the atomic arrangement and bonding that is presented in a molecular formula. A molecular formula gives more information about a molecule than its empirical formula, however, it is more difficult to establish.

Relationship between Molecular Formula and Empirical Formula

The molecular formula specifies the actual number of each type of atom in a molecule, whereas an empirical formula for a chemical compound is a simple expression that represents the ratio of the elements in the compound. For example, the molecular formula of hydrogen peroxide is H_2O_2, whereas its empirical formula is HO. Here, the n-factor is 2. The empirical formula might be the same for different chemical compounds. For example, glucose, formaldehyde, and acetic acid have the same empirical formula, CH_2O, but their molecular formulae are different. For some chemical compounds, both formulae are the same, like water (H_2O), hypochlorous acid (HClO), formaldehyde (CH_2O), methane (CH_4), etc.

Molecular Formula = n × Empirical Formula

Where n is a whole number

To gain a better understanding of the empirical formula, consider the molecular formula of acetylene, C_2H_2. The empirical formula is the simplest ratio of a number of each different atom present in the compound.

Molecular Formula = n × Empirical Formula

$C_2H_2 = 2 \times CH$

So, the Empirical formula of acetylene is CH.

How to Calculate empirical and molecular Formula

1. Empirical formular can be determined from the percentage compositions or masses of the various elements present in the compound.
2. The number of atoms present in that mass can be determined by dividing with the relative atomic masses of the atoms of the element.
3. The ratio of the atoms present is then determined by dividing the values obtained in step 2 above by the smallest ratio.

To determine the molecular formula

1. Find the n-factor by using its formula. n = Molar Mass/Empirical Formula Mass
2. Multiply all the subscripts in the empirical formula by n and the resultant formula is the required molecular formula.

Solved Examples

1. A compound is composed of 82.78% nitrogen and 17.22% hydrogen. If its molecular weight is 17.031 g/mol, then find its empirical and molecular formula. (N= 14. H = 1)

Solution:

Elements present:	N	H
% composition:	82.78	17.22
No of atoms:	82.78/14	17.22/1
	= 5.913	= 17.22
Divide by the smallest atom	5.913/5.913	17.22/5.913
Ratio	1	3

Empirical formula = NH_3

Molecular Formula : $(NH_3)n = 17.031$ g/mol

$(14 + 3x1)n = 17.031$ g/mol

$17n = 17.031$

$n = 17.031/17$

$n = 1$

Molecular Formula : $(NH_3)n = (NH_3) \times 1$

Therefore Molecular Formula = NH_3

2. The Empirical formula of Butane is C_2H_5. Calculate the Molecular formula when the measured mass of the compound is 58.1224 (C = 12.011, H = 1.00784)

Solution:

Given empirical formula = C_2H_5

$(C_2H_5)n = (2 \times 12.011 + 5 \times 1.00784)n = (29.0612)n$

$(29.0612)n = 58.1224$

$n = 58.1224/29.0612 = 2$

$(C_2H_5)n = (C_2H_5) \times 2 = C_4H_{10}$

3. Find the empirical for the compound with molecular formula $C_6H_{12}O_6$.

Solution

$C_6H_{12}O_6 = 6 \times CH2O$

we know that Molecular Formula = n × Empirical Formula

Here n = 6

So empirical formula for the given compound is CH_2O.

4. A compound is composed of 68.29% carbon, 12.02% hydrogen, and 21.69% oxygen. If its molecular weight is 86.136 g/mol, then find its molecular formula.

Solution:

Elements present:	C	H	O
% composition:	68.29	12.02	21.69
No of atoms:	68.29/12	12.02/1	21.69/16
	= 5.691	= 12.02	=1.356
Divide by the smallest atom	5.691/1.356	12.02/1.356	1.356/1.356
Ratio	4	9	1

Therefore Empirical formula = C_4H_9O

Molecular Formula : $(C_4H_9O)n = 86.136$ g/mol

$(12 \times 4 + 1 \times 9 + 16 \times 1)n = 86.136$ g/mol

$73n = 86.136$

$n = 86.136 / 73$

$n = 1$

Molecular Formula : $(C_4H_9O)n = (C_4H_9O) \times 1$

Therefore Molecular Formula = C_4H_9O

EVALUATION;

1. The empirical formula of a compound of carbon, hydrogen, and oxygen is CH2O. If its molar mass is 60.052 g/mol, then determine the molecular formula of the compound.
2. Boric acid is composed of 21.14% boron, 4.65% hydrogen, and 74.21% oxygen. If its molecular weight is 61.83 g/mol, then find its molecular formula.
3. Oxalic acid is composed of 27.42% carbon, 2.33% hydrogen, and 70.25% oxygen. If its molecular weight is 90.035 g/mol, then find its molecular formula.
4. The empirical formula of a compound that is composed of hydrogen, chlorine, and oxygen is HClO. If its molar mass is 52.46 g/mol, then determine the molecular formula of the compound.
5. Find a molecular formula for the compound having the empirical formula CH2 with a molecular weight of 42.08.

Week 8

Topic: Chemical Symbols and Formulae II: The Laws of chemical Combination
Specific Objectives: By the end of the lesson, the students should be able to

1. Determine the percentage compositions of elements in a compound
2. Appraise the laws of chemical combination
3. State and explain the law of conservation of matter
4. State and explain the law of constant composition/definite Proportion
5. State and explain the law of multiple proportion (stating the laws and illustrating them with experiments)

Calculations Using Masses and Formulae

The mass of 1 mole of a compound is the sum of the masses of the moles of its component elements. For example, mass of 1 mole of $CaCO_3$

= mass of 1atom of Calcium + mass of 1atom of Carbon + mass of 3 atoms of Oxygen

= 40 + 12 + (3x16)

= 100g

If we know the right formula of a compound, we can calculate the percentage composition of all the component of a compound.

Example: Calculate the percentage composition of all the component elements in $NaNO_3$ [Na=23, N=14, O=16]

Solution

% mass of Na $= \frac{mass\ of\ Na}{molar\ mass\ of\ NaNO3}$ x 100% $= \frac{23}{(23+14+16x3)}$ x100% = 27.06%

% mass of N $= \frac{mass\ of\ N}{molar\ mass\ of\ NaNO3}$ x 100% $= \frac{14}{(23+14+16x3)}$ x100% = 16.47%

% mass of O $= \frac{mass\ of\ O}{molar\ mass\ of\ NaNO3}$ x 100% $= \frac{16\ x\ 3}{(23+14+16x3)}$ x100% = 56.47%

Try these:

Determine the percentage composition of the following

1. H in H_2O
2. S in H_2SO_4
3. C in $CaCO_3$
4. O in NaOH
5. Cl in NaCl

[H=1, O=16, Na =23, Cl = 35.5, Ca = 20, S=32, C=12)

The Laws of chemical Combination

Matter can be transformed from one form to another. These transformations often occur as a result of the combination of different types of matter. The combination of different elements to form compounds is governed by certain basic rules or laws.

Chemical combinations or reactions are governed by several fundamental laws and principles. These laws help us understand and predict how substances will react with each other. These laws are referred to as laws of chemical combination.

In this lesson, we will explore four important laws of chemical combination; the law of Conservation of Mass, the law of Definite Proportions (or Law of Constant Composition), the Law of Multiple Proportions and the law of reciprocal proportions.

I. Law of Conservation of Mass:

This law was first formulated by Antoine Lavoisier and is a fundamental principle in chemistry as it forms the basis for all chemical reactions.

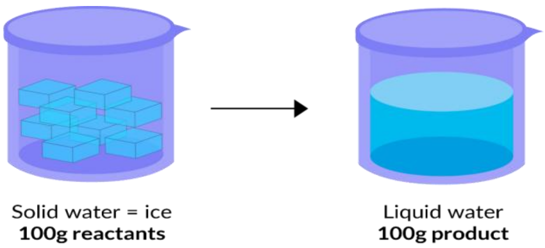
The Law of Conservation of Mass states that "the total mass of substances before a chemical reaction is equal to the total mass of substances after the reaction".

This law is also called the law of **indestructibility of matter** and can also be stated as "matter is neither created nor destroyed during" a chemical reaction but changes from one form to another". In other words, matter it is merely rearranged during chemical reactions.

The law of conservation of mass can be applied to both physical changes(such as melting of ice, etc) and chemical processes (such as rusting, combustion, etc). Burning is a chemical process. The flames are produced due to the fuel undergoing combustion (burning). When a piece of wood burns, the mass of the smoke, ashes, and gases equals the original mass of the total charcoal and the oxygen when it first reacts, so this means that the total mass of any product equals the total mass of the reactant derived.

The Law of conversation of Mass can also be applied to Physical Changes.

If 100 grams of ice melted, it produced 100 grams of water.



In the laboratory, this law can easily be verified by the reaction between silver trioxonitrate(V) ($AgNO_3$) solution and sodium chloride (NaCl). When a solution of silver trioxonitrate(V) is treated with a solution of sodium chloride, a white precipitate of silver chloride (AgCl) and sodium trioxonitrate (V) ($NaNO_3$) are formed.

Equation of reaction

$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

If the law is true, then the total mass of $AgNO_3$ and NaCl before the reaction should be the same as the total mass of AgCl precipitate and $NaNO_3$ solution. The experiment is done in a specially designed H shaped tube called the Landolts tube. NaCl solution is taken in one limb of the tube while $AgNO_3$ is taken in the other limb. The tube is weighed and then inverted so that the solutions react chemically. The reaction takes place and a precipitate of silver chloride is obtained. The tube is weighed again. The mass of the tube is found to be exactly the same as the mass obtained before inverting the tube.

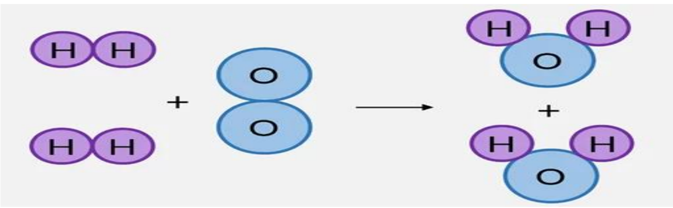
The experiment therefore verifies the law of conservation of mass.

Example 1: consider the reaction between hydrogen and oxygen to form water

Equation of reaction:

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

(H=1, O= 16)



to verify if the reaction follows the law of conservation of mass, the total mass of the reactants must be equal to the products.ie

Mass of reactants = mass of $2H_2$ + mass of O_2

= 4amu + 32amu = 36amu

Mass of product = mass of $2H_2O$

= 2(2 + 16) = 36amu

Mass of reactants = mass of products; therefore the chemical reaction follows the law of conservation of mass

Example 2:

In an experiment, 5.0 grams of $CaCO_3$ on heating produces 2.8g of calcium oxide(CaO) and 2.2g of CO_2 . Prove that these results are in agreement with the law of conservation of mass

Solution:

Mass of the reactants, $CaCO_3$: 5.0 g

Mass of the products = Mass of CaO + Mass of CO_2

= 2.8 g + 2.2 g = 5.0 g.

Mass of the reactants molecules equals the Mass of the product molecules, therefore the results are in agreement with the law of conservation of Mass

Example 3:

Potassium hydroxide (KOH) readily reacts with carbon dioxide (CO_2) to produce potassium trioxocarbonate (IV) (K_2CO_3) and water (H_2O). How many grams of potassium carbonate is formed if 224.4 g of KOH reacted with 88.0 g of CO_2. The reaction also produced 36.0 g of water.

Solution: Let the grams of potassium carbonate produced = X

224.4 g KOH + 88.0 g CO2 = X g K_2CO_3 + 36.0 g H_2O

312.4 g reactants = X g K_2CO3 + 36.0 g H_2O

312.4 g reactants – 36.0 g H_2O = X g K_2CO_3

Therefore X = 276.4 g of potassium carbonate(K_2CO_3)

II. Law of Definite Proportions:

The Law of Definite Proportions, also known as the Law of Constant Composition, was developed by Joseph Proust.

It states that "all pure samples of a particular chemical compound contains similar elements combined in the same proportion by mass".

This means that a chemical compound no matter its origin or its method of preparation always contains the same elements in the same proportion by mass. Simply put; the ratio of the masses of the elements in a compound is always fixed and doesn't change.

In the laboratory, this law is verified by the following experiment.

Prepare pure samples of Copper (II) oxide by two different methods

A. By the decomposition of copper (II) trioxonitrate(V)

$2Cu(NO_3)_2 \rightarrow 2CUO + 4NO2 + O_2$

B. By heating copper(II) trioxocarbonate (IV)

The copper (II) oxide prepared by both methods always contains the same elements copper and oxygen combined together in the same fixed proportion of 4:1 by weight. This illustrates the law of definite proportions.

Example 4:

1.37g of CuO were reduced by hydrogen and 1.098g of Cu were obtained. In another experiment, 1.178g of Ci were dissolved in nitric acid and the resullting copper trioxonitrate (V) was converted into CuO by ignition. The weight of CuO formed was 1.476g. show that these results prove the law of definite proportions

Solution

Experiment 1;

Weight of CuO = 1.375g

Weight of Cu = 1.098g

Weight of oxygen (O) = (1.375 - 1.098)g = 0.277g

Ratio of copper to oxygen = 1.098 : 0.277

= 4: 1

Experiment 2:

Weight of CuO = 1.476g

Weight of Cu = 1.178g

Weight of oxygen (O) = (1.476 - 1.178)g = 0.298g

Ratio of copper to oxygen = 1.178 : 0.298

= 4: 1

In both experiments, the ratio of copper to oxygen is 4:1; hence the results prove law of definite proportion.

III. Law of Multiple Proportions:

The Law of Multiple Proportions, was proposed by John Dalton. It states that "when two elements combine to form more than one compound, the ratios of the masses of one element that combine with a fixed mass of the other element are in small whole-number ratios".

For example, carbon combines with oxygen to form two different compound; carbon (II) oxide (CO) and carbon (IV) oxide (CO_2). The proportion by weight of the two elements in the compound are

CO = C:O; 12:16

CO_2 = C:O 12: 32

The weights of oxygen(16 and 32)g that combine with a fixed weight of carbon (12g) are in the ratio of 16:32, ie, 1:2, a simple numerical ratio.

This law is particularly important in understanding the formation of different compounds by the same elements.

Example 5:

In an experiment, 34.5g oxide of a metal was heated so that oxygen was liberated and 32.1g of metal was obtained. In another experiment, 119.5g of another oxide of the same metal was heated and 103.9g metal was obtained and oxygen gas was liberated. Calculate the mass of O_2 liberated in each experiment. Show that the data explain the law of multiple proportions.

Solution

Experiment 1;

Weight of the metal oxide = 34.5g

Weight of the metal = 32.1g

Weight of oxygen liberated = 2.4g

32.1g metal combines with 2.4g oxygen

1g metal will combine with 0.075g of oxygen

Experiment 2

Weight of the metal oxide = 119.5g

Weight of the metal = 103.9g

Weight of oxygen liberated = 15.6g

103.9g metal combines with 15.6g oxygen

1g metal will combine with 0.15g of oxygen

Therefore, different weights of oxygen that combine with a fixed mass of the metal (1g) are in the ratio of 0.150:0.075 = 2:1

The proportion by weight of oxygen is indicated by simple ratio. Thus, the law of multiple proportion is obeyed.

IV. The law of Reciprocal Proportions

This is also called the law of equivalent proportions or the law of permanent ratios. It states that the masses of several elements A, b and C which combines separately with a fixed mass of another element D are the same as the masses in which A, B and C themselves combine with one another. Simply, this law says that if we know the proportion of elements in compound AB and the proportion of elements in compound BC, we can determine the proportion of elements in compound AC.

For instance, methane(CH_4) and water(H_2O).

In methane(CH_4), the weight of carbon is 12 while the weight of H is 1. Since we have 4 atoms of hydrogen for every atom of carbon, the proportion is 12:4 or 3:1.

In water (H_2O), the proportion of elements is 2:16 or 1:8 (oxygen has an atomic weight of 16)

Methane and water both contain hydrogen and one other element (ie, carbon and oxygen). according to this law, if we combine carbon and oxygen it should be in a ratio of 3:8 or a simple multiple of that ratio.

When carbon and oxygen combine, they form carbon(IV) oxide, CO_2, which has a proportion of 12:32 or 3:8. This shows the law of reciprocal proportions.

Evaluation

1. Given a compound with a molar mass of 180 g/mol that contains 72 g of carbon, 12 g of hydrogen, and 96 g of oxygen, determine the percentage composition of carbon in the compound.
2. Describe the law of conservation of matter and provide an example that illustrates this law in a chemical reaction.

3. Explain the law of constant composition (or definite proportion) and give an example.
4. State the law of multiple proportions and illustrate it with an example involving two compounds composed of the same elements.
5. Appraise how the laws of chemical combination are crucial in understanding chemical reactions.

Week Nine

Topic: Chemical Equations

Specific Objectives: By the end of the lesson, the students should be able to;

1. Appraise chemical equation (definition),
2. Write balanced chemical equations,
3. Highlight the information given or not given by a chemical equation;
4. Discuss the mole and apply it in solving problems;
5. Identify the Avogadro's number (N_A)
6. Apply equation to calculate reacting masses of substances;

Chemical Equation

Chemical equations are symbolic representation of chemical reactions in which the reactants and the products are expressed in terms of their respective chemical formulae. This involves the use of symbols to represent factors of a chemical reaction such as the direction of the reaction, the number of moles of the reactants and products, and the physical states of the reacting entities.

Chemical reactions can be represented on paper with the help of chemical equations using chemical formulae and symbols.

The reactants are written on the left hand side while the products are written on the right hand side of the equation and separated by an arrow. Examples of chemical equations include

1. Reaction between hydrochloric acid and sodium hydroxide

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

2. Reaction between hydrogen and oxygen gases to form water

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

3. Reaction between calcium chloride and silvertrioxonitrate (V)

$CaCl_{2(aq)} + 2AgNO_{3(aq)} \rightarrow Ca(NO_3)_{2(aq)} + 2AgCl_{(s)}$

4. Reaction of methane and oxygen

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

Where a reaction is not possible, it is wrong to write a chemical equation to represent it. For instance, copper (Cu) does not react with hydrochloric acid, therefor it is wrong to write a chemical equation for it.

Parts of a chemical equation

$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$

1. The elements present in the reaction are represented by their chemical symbols
2. The number before the chemical formulae are called co-efficient. They represent the number of moles of the reactants or products formed.
3. The subscripts are part of the chemical formulae and cannot be changed.
4. The physical states of the reactants and products are represented using (g), (s), (l), and (aq) meaning gaseous, solid, liquid and aqueous states respectively.

Balancing of chemical equations

All chemical equations must be balanced in order to comply with the law of conservation of mass or indestructibility of matter. **The number of moles of each atom on the reactant side must be equal to that of the product side for a chemical equation to be balanced.**

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products because that would change the chemical identity of the species present. Equations can only be balanced by changing the coefficients.

The simplest and most generally accepted method for balancing chemical equation is by inspection, ie, trial and error.

Examples: balance the following chemical equations

1. $C_7H_{18(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
2. $Pb(NO_3)_{2(aq)} + NaCl_{(aq)} \rightarrow NaNO_{3(aq)} + PbCl_{2(s)}$
3. $H_2SO_{4(aq)} + K_2CO_{3(aq)} \rightarrow K_2SO_{4(aq)}) + H_2O_{(l)} + CO_{2(g)}$
4. $HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$
5. $C_8H_{18(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

<u>Solution</u>

1. $C_7H_{18(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

	left	right
C	7	1
H	16	2
O	2	3

The equation is not balanced

Since the total no of carbon(C) atoms on the left is 7, add 7 to CO_2, ie,

$C_7H_{16(l)} + O_{2(g)} \rightarrow 7CO_{2(g)} + H_2O_{(l)}$

Hydrogen on the left is 16. to balance the H atoms on the right, add 8 in front of H_2O, ie,

$C_7H_{16(l)} + O_{2(g)} \rightarrow 7CO_{2(g)} + 8H_2O_{(l)}$

The number of oxygen atoms on the right have changed to 22 (ie 14 in CO_2 and 8 in H_2O), therefore place 11 in front of O_2(ie 11 x 2 = 22)

$C_7H_{16(l)} + 11O_{2(g)} \rightarrow 7CO_{2(g)} + 8H_2O_{(l)}$

	left	right
C	7	7
H	16	16
O	22	22

The equation is now balanced.

Hint: always start from the most complex formula

2. $Pb(NO_3)_{2(aq)} + NaCl_{(aq)} \rightarrow NaNO_{3(aq)} + PbCl_{2(s)}$

	left	right
Pb	1	1
N	2	1
O	6	3
Na	1	1
Cl	1	2

The equation is not balanced

Place a 2 in front of $NaNO_3$ to balance the N atoms on the right

$Pb(NO_3)_{2(aq)} + NaCl_{(aq)} \rightarrow 2NaNO_{3(aq)} + PbCl_{2(s)}$

This has changed the O and Na atoms on the right to 6 and 2 respectively

Place a 2 in front of NaCl to balance the Na atoms on the left

$Pb(NO_3)_{2(aq)} + 2NaCl_{(aq)} \rightarrow 2NaNO_{3(aq)} + PbCl_{2(s)}$

	left	right
Pb	1	1
N	2	2
O	6	6
Na	2	2
Cl	2	2

Therefore the equation is now balanced

3. $H_2SO_{4(aq)} + K_2CO_{3(aq)} \rightarrow K_2SO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

	Left	Right
H	2	2

S	1	1
O	7	7
K	2	2
C	1	1

The equation of reaction is balanced

4. $HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

	Left	Right
H	1	2
Cl	1	1
Na	2	1
C	1	1
O	3	3

The equation is not balanced

Starting from Na_2CO_3, place a 2 in front of NaCl to balance Na atoms

$HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Cl on the right is now 2, so place a two in front of HCl to balance the number of Cl and H atoms

$2HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

	Left	Right
H	2	2
Cl	2	2
Na	2	2
C	1	1
O	3	3

The equation is now balanced

Equations of reaction can also be balance using fractions, for example

5. $C_8H_{18(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

	Left	Right
C	8	1
H	18	2
O	2	3

Starting from C_8H_{18}, place 8 in front of CO2 to balance the number of C atoms on the right, ie,

$C_8H_{18(l)} + O_{2(g)} \rightarrow 8CO_{2(g)} + H_2O_{(l)}$

Place 9 in front of H_2O to balance the H atoms, ie,

$C_8H_{18(l)} + O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(l)}$

O atoms is now 2 on the left and 25 on the right. To balance the O atoms on the left introduce a fraction 25/2. ie,

$C_8H_{18(l)} + \frac{25}{2}O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(l)}$

To remove the fraction, multiply the whole equation by the denominator of the fraction ie 2 to give the following

$2C_8H_{18(l)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(l)}$

	Left	Right
C	16	16
H	36	36
O	50	50

Therefore the equation is balanced

TRY THESE

Balance the following equations of reaction

1. $Cu_2S_{(s)} + O_{2(g)} \rightarrow Cu_2O_{(s)} + SO_{2(g)}$
2. $NH_{3(g)} + O_{2(g)} \rightarrow H_2O_{(g)} + N_{2(g)}$
3. $Pb(NO_3)_{2(aq)} + FeCl_{3(aq)} \rightarrow Fe(NO_3)_{3(aq)} + PbCl_{2(s)}$
4. $C_6H_{14(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

Information provided by a balanced chemical equation

1. The reactants and products
2. The elements and radicals involved in the reaction
3. The movement of elements and radicals during the reaction
4. The stoichiometry of the reaction, ie the relationship between the amounts of the reactants and products
5. The direction of the reaction, ie, whether the reaction is reversible or not
6. The physical states of matter in which the substances are present

Information not provided by a chemical equation

A balanced chemical equation does not tell us

1. The speed of the reaction
2. The heat changes during the reaction
3. The colors of the reactants or products
4. The concentration of the reactants or products

The mole

The mole is a unit similar to other units such as a pair, a dozen, etc. It provides a specific measure of the number of atoms or molecules in a bulk sample of matter.

A mole is defined as the amount of substance containing the same number of discrete entities (atoms, molecules or ions) as the number of atoms in a sample of 12g of carbon 12(C-12)

Particles such as atoms, molecules and ions are extremely small and are difficult to work with during a chemical reaction. Scientists therefore need a unit for measuring the amount of particles in a given mass of a substance and such a unit is called the mole.

The mole provides a link between an easily measured macroscopic property; mass, and other properties such as the number of atoms, molecules, ions, etc.

1 mole of any element contains the same number of atoms as 1mole of any other element. The masses of 1mole of different elements however are different since the masses of the individual atoms are different.

1 mole of a given substance is written in its formula form as one particle of the substance regardless of whether it is a while molecule, an atom or an ion, thus

K^+ means 1 mole of potassium ion

K means 1 mole of potassium atom

KOH means 1 mole of potassium hydroxide

O_2^- means 1 mole of oxygen ion

O_2 means 1 mole of oxygen molecules

$2O_2$ means 2 moles of oxygen molecules or 4 moles of oxygen atom

H_2O means 1 mole of water molecule or 2 moles of hydrogen and 1 mole of oxygen atoms

$2H_2O$ means 2 moles of water molecules or 4moles of hydrogen atoms and 2 moles of oxygen atoms.

Note: mole is not the same as molecule.

The Avogadro's Constant

The number of particles composing 1 mole has been experimentally determined to be

$$\mathbf{6.02 \times 10^{23}}$$

This number is called the Avogadro's constant (or number) in honor of the Italian scientist; Amedeo Avogadro.

Just as I dozen of eggs contains 12 eggs, 1 mole of eggs contains 6.02×10^{23} eggs. Similarly, 1 mole of sodium atoms contains 6.02×10^{23} atoms and 1 mole of hydrogen atoms contain 6.02×10^{23} Hydrogen atoms.

The relative atomic and molecular masses of a substance give the relationship between the mass of a substance and the mole. Thus;

Relative atomic mass of any element in grams ≡ 1 mole of an atom of that element ≡ 6.02×10^{23} atoms of the element. Also,

Relative molecular mass of an element or a compound in grams ≡ 1 mole of a molecule of that element or compound ≡ 6.02×10^{23} molecules of that element or compound

For example: 1 mole of magnesium = 24 g of magnesium = 6.02×10^{23} atoms of magnesium;

1 mole of calcium trioxocarbonate(IV) = 100 g of calcium trioxocarbonate(IV) = 6.02×10^{23} atoms of calcium trioxocarbonate(IV).

Calculations Involving Moles

While performing calculations involving moles, the following formula can be helpful

1. Number of moles = $\frac{\text{mass of the sample}}{molar\ mass}$

2. Number of moles of a substance = $\frac{\text{number of particles}}{6.02 \times 10^{23}}$

3. Number of particles(ie atoms, molecules or ions) = number of moles x $6.02 \times 10^{23.}$

Examples

1. How many moles of Iron (Fe) are present in a pure sample weighing 558.45grams?

(Fe = 55.845)

Solution

Given, mass of Fe = 558.45g, molar mass of Fe = 55.845g/mol

Number of moles = $\frac{\text{mass of the sample}}{molar\ mass} = \frac{558.45g}{55.845g/moil}$ = 10moles

Alternatively,

1mole of Fe = 55.845g

Xmole of Fe = 558.45g

X = $\frac{558.45g}{55.845g}$ = 10moles

2. How many molecules of water(H2O) are present in 36grams of water (H=1, O= 16)

Solution

Given, mass of H2O = 36g, molar mass of Fe = 18g/mol

Number of moles = $\frac{\text{mass of the sample}}{molar\ mass} = \frac{36g}{18g/moil}$ = 2moles

But Number of molecules = number of moles x $6.02 \times 10^{23.}$

= 2 x $6.02 \times 10^{23.}$

= 1.204×10^{24} H_2O Molecules.

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Calculations Involving Equations (Stoichiometry)

A balanced chemical equation gives information about the mole ratios of both the reactants and the products and hence the molecular mass ratios. It can be used to predict the amount of reactants used or products formed from given quantities of reactants. Stoichiometry is an important concept in chemistry that helps us use balanced chemical equations to calculate amounts of reactants and products.

Examples.

1. Determine the amount of barium tetraoxosulphate (VI) ($BaSO_4$) that will be formed by adding excess tetraoxosulphate (vI) acid to a solution containing 15.4g of barium chloride according to the following equation of reaction

$BaCl_{2(aq)} + H_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2HCl_{(aq)}$

(Ba = 137.3, Cl = 35.5, S = 32.0, O = 16.0, H = 1.0)

<u>Solution</u>

Given mass of $BaCl_2$ = 15.4g, molar mass of $BaCl_2$ = 208.3, mass of $BaSO_4$ = X, molar mass of $BaSO_4$ = 233.3

From the equation

1 mole of $BaCl_2$ → 1 mole of $BaSO_4$

208.3g of $BaCl_2$ → 233.3g of $BaSO_4$

15.4g of $BaCl_2$ → Xg of $BaSO_4$

$X = \frac{15.4g \text{ x } 233.3g}{208.3g}$ = 17.25g of $BaSO_4$

2. Propane (C_3H_8) burns in this reaction:

$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 4H_2O(l) + 3CO_{2(g)}$

If 200 g of propane is burned, how many g of H_2O is produced? (C = 12, H = 1, O = 16)

<u>Solution</u>

Given; mass of C_3H_8 = 200g, molar mass of C_3H_8 = 44g

Mass of H2O = ?, molar mass of H2O = 18g

From the equation,

1 mole of C_3H_8 →4 moles of H_2O

44g of C_3H_8 →4 x 18g of H_2O

200g of C_3H_8 →Xg of H_2O

$X = \frac{200 \, x \, 4 \, x \, 18}{44}$ = 327.27g of H2O

EVALUATION:

1. 10g of zinc was dissolved in excess tetraoxosulphate (VI) acid according to the following equation of reaction

$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(s)} + H_{2(g)}$

Calculate the mass of hydrogen gas that will be collected.(Zn=65, H =1, S= 23,O = 16)

2. What amount of silver trioxonitrate(V) would be required to precipitate 5.6g of silver chloride from sodium chloride solution?

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

(Ag = 108, N = 14, Na = 23, Cl = 35.5, O = 16)

3. Calculate the mass of 6 moles of ozone, O_3? (O = 16) What is the mass of 4 moles of sodium (Na) metal? (Na = 23)

4. How many atoms exist in 2g of magnesium, (Mg)? (1 mole of Mg = 6×10^{23}; Mg = 24).

5. How many moles are there is 20 g of $NaNO_3$? (take Na = 23, N = 14, O = 16)

6. Calculate the number of moles and number of atoms in 72 g of boron. (Boron = 10.8).

Week Ten

Topic: Chemical Combination/Bonding

Specific Objectives: By the end of the lesson, the students should be able to;

1. Discuss chemical bonding
2. Classify chemical bonds into Strong bonds (e.g. Electrovalent (ionic), Covalent, Co-ordinate covalent (dative) and Weak bonds (e.g. Metallic, Hydrogen bond and Vander waals force;)
3. Describe the types of bonds in various compounds
4. Enumerate the characteristics and differences between ionic and covalent compounds (in terms of volatility, solubility and electrical conductivity, Melting point and boiling point

Chemical Bonding

This refers to the formation of chemical bonds between two or more atoms, molecules or ions to form a chemical compound. A chemical bond is the force of attraction that joins two or more atoms together to form chemical compounds. Chemical bonds are what keep atoms or molecules together in a compound. Elements enter into a chemical combination in order to attain stability of the noble gases or rare gases(either the duplet or octet structure). this is called the octet rule.

The Octet Rule

The principle of atom of an element to attain the maximum of eight electrons in the valence shell of atoms is called the octet rule.

All atoms except noble gases have less than eight electrons in their valence shell. In other words, the valence shells of these atoms do not have stable configurations. Therefore, they combine with each other or with other atoms to attain stable electronic configurations.

The octet rule is the tendency of atoms of various elements to attain stable configuration of eight electrons in their valence shells in the cause of chemical combination.

Types of chemical bonds

There are two major types of chemical bonds; the strong bonds and the weak bonds

Examples of the strong bonds include the Electrovalent/ionic bond, Covalent bond and Metallic bonds

Examples of the weak bonds are Van Der Waals , Hydrogen bonds

Electrovalent/ionic bonding

This is a type of chemical bond which involves a transfer of electrons from one atom to the another, usually from a metallic atom to a non metallic atom.

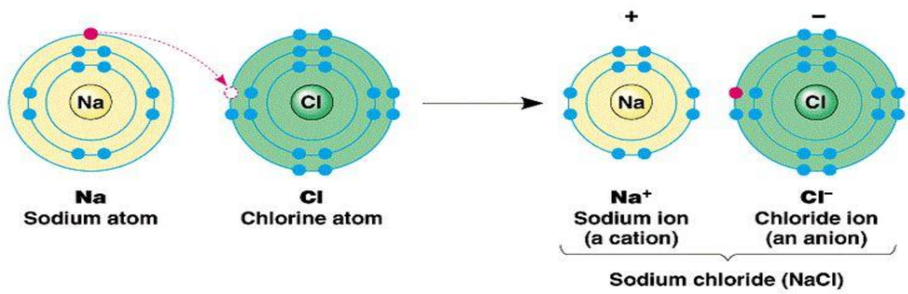
The metallic atom(donor) losses an electron to become a positively charged cation while the non metallic atom (acceptor) gains the electron to become a negatively charged anion.

Examples:

1. **Formation of sodium chloride (NaCl):** NaCl is an ionic compound that is formed by the formation of Na^+ ions and Cl^- ions. The electronic configuration of Sodium is (2,8,1) and it has 1 electron more than a stable noble gas configuration (2,8). So it easily forms Na+ ion by losing an electron. The electronic configuration of Chlorine is (2,8,7) and it has 1 electron short to achieve a stable noble gas configuration (2,8,8). So it easily forms Cl– ion by gaining an electron.

Now one electron from Na goes to Cl and they both form the Na^+ cation and Cl^- anion respectively, which are joined together by the electrostatic force of attraction, forming NaCl ionic compound.

The formation of the NaCl compound by the ionic bond is shown in the image below



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Equation of reaction

$Na \rightarrow Na^+ + e^-$

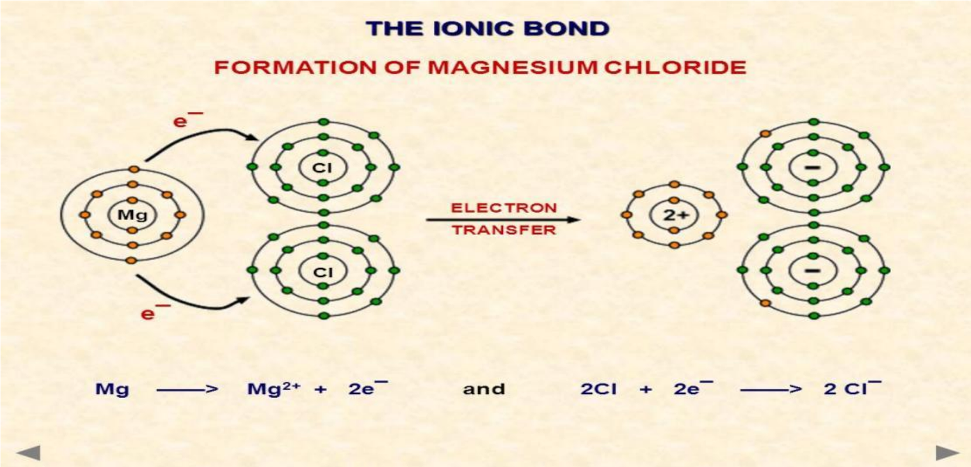
$Cl + e- \rightarrow Cl^-$

$Na + Cl \rightarrow Na^+Cl^-$

2. **Formation of Magnesium chloride ($MgCl_2$):** $MgCl_2$ is formed when oppositely charged magnesium and chloride ions attract each other in the following ways.

The atomic number of magnesium is 12 so its electronic configuration is 2,8,2 while the atomic number of chlorine is 17with an electronic configuration is 2,8,7.

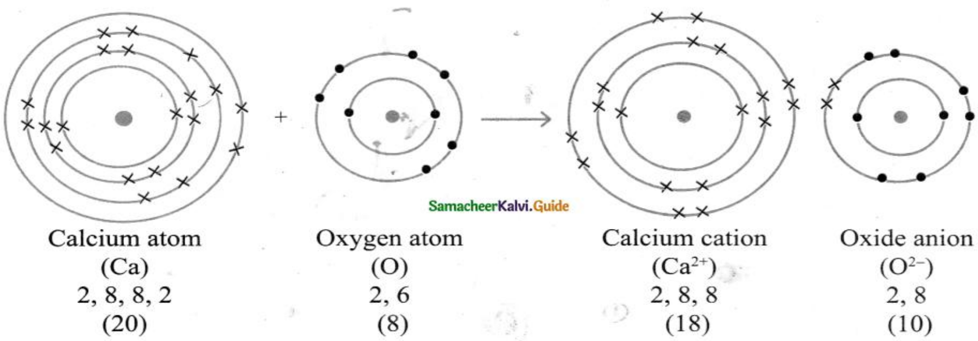
Chlorine needs 1electron to attain stable electronic configuration so magnesium donates its two electrons to two chlorine atoms and thus magnesium(Mg2+) and two chloride(Cl-) ions are formed. The formation of Magnesium chloride MgCl2 is shown below:



3. **Formation of calcium oxide (CaO):** In CaO, the electronic configuration of calcium atom is 2, 8,8,2. Calcium has two valence electrons in its outermost shell and can readily lose two electrons to attain nearest noble gas configuration.

The electronic configuration of oxygen atom in its ground state is 2,8,6. The oxygen atom requires two more electrons to attain the nearest noble gas configuration. The oxygen anion attains the electronic configuration of nearest noble gas i.e Neon and hence, is a stable anion.

The formation of Magnesium chloride CaO is shown below:



Equation of reaction;

Ca → Ca^{2+} + 2^{e-}

O + 2^{e}- → O^{2-}

Ca + O → $Ca^{2+}O^{2-}$

Characteristics of Electrovalent Compounds

Structure: Electrovalent compounds exist mainly as solids at room temperature. Crystals of electrovalent compounds are hard and brittle.

Melting and boiling points: Electrovalent compounds have high melting and boiling points because of the strong electrostatic atraction between the ions.

Solubility: Ionic compounds are generally soluble in water and other polar solvents like ethanol.

Electrolytes: electrolytes are substances which conduct electricity either in the molten form or when dissolved in water. Ionic compounds are good electrolytes. This is because they are composed of ions which are able to move about when the compound is in a liquid state . in the solid state, the ions are in fixed positions in the crystal lattice and so cannot conduct electricity.

Covalent Combination/Bonding

Covalent bonding involves the sharing of electrons between two or more atoms. There are two types of covalent combination; the ordinary covalent bonds and the coordinate covalent bonds

Ordinary covalent bonds: in ordinary covalent bonding, there is a sharing of electrons between the two reacting atoms so that the two atoms attain the stable octet structure. This pair of electrons is called the shared pair. The shared pair now extend around the nuclei of the atoms leading to the creation of a molecule. Conventionally, the shared pair is represented by a stroke between the atoms in association. One stroke represents two electrons.

Diatomic molecules such as H_2, O_2, Cl_2, Br_2, etc are formed by covalent combination. Organic compounds (ie compounds of Carbon) are also formed by this process.

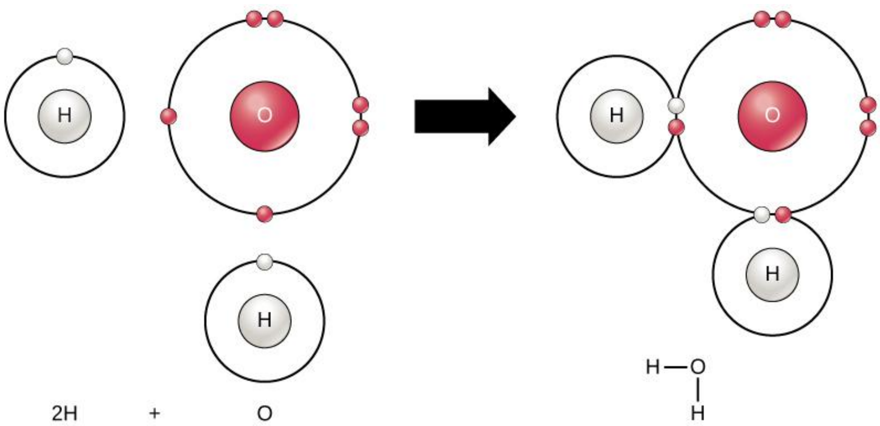
Types of bonds formed in covalent bonding

Depending upon the number of shared electron pairs, the covalent bond can be classified into: Single Covalent Bond, Double Covalent Bond and Triple Covalent Bond

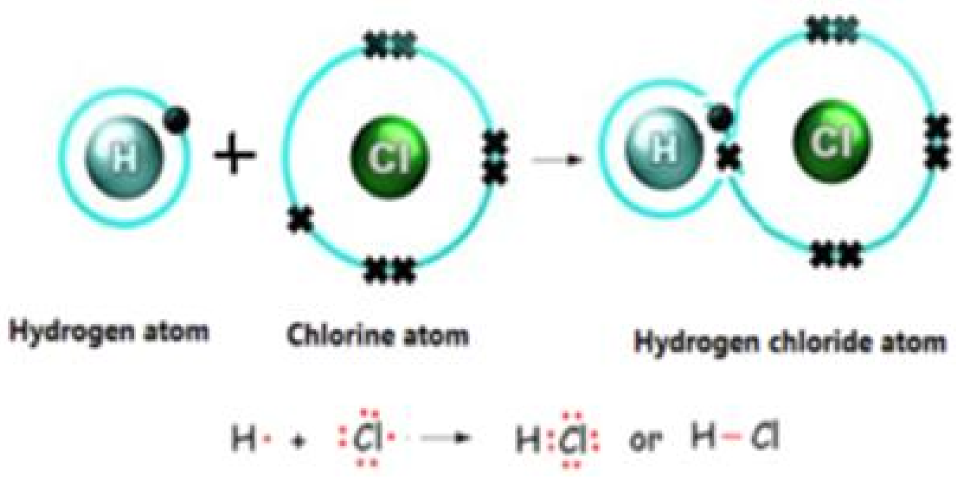
Single Bonds: A single bond is formed when only one pair of electrons is shared between the two participating atoms. It is represented by one dash (-). Although this form of covalent bond has a smaller density and is weaker than a double and triple bond, it is the most stable.

Examples of single covalent bonds

1. **Formation of water molecule H2O:** the oxygen atom, which has six electrons in its valence shell, completes its octet by sharing its two electrons with two hydrogen atoms to form a water molecule.



2. **Formation of hydrogen chloride HCl:** The HCL molecule has one hydrogen atom with one valence electron and one chlorine atom with seven valence electrons. In this case, a single bond is formed between hydrogen and chlorine by sharing one electron



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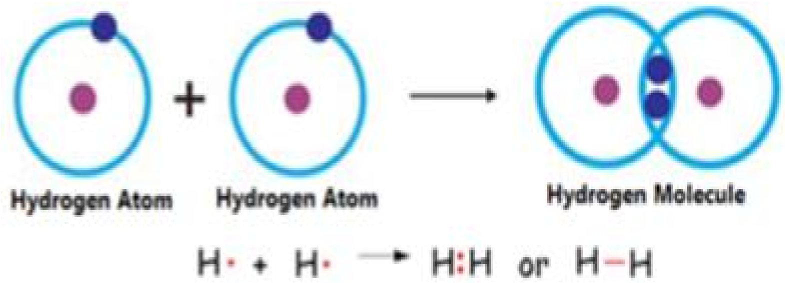
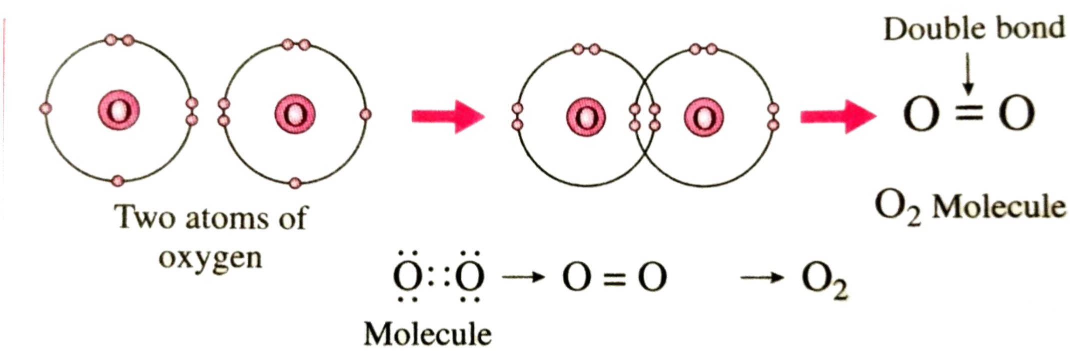
3. **Formation of Hydrogen molecule H_2:** in this case, two hydrogen atoms each share their valence electron in order to attain the duplet state of the nearest noble gas(helium). the formation of hydrogen molecule is shown below.



Double Bonds: A double bond is formed when two pairs of electrons are shared between the two participating atoms. It is represented by two dashes (=). Double covalent bonds are much stronger than single bonds, but they are less stable.

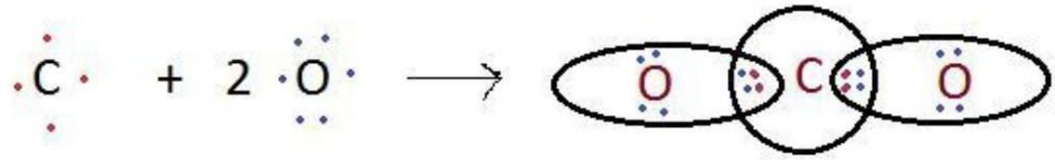
Example,

1. **Formation of Oxygen Molecule O_2:** In the formation of the oxygen molecule, each oxygen atom has six electrons in its valence shell. Each atom requires two more electrons to complete its octet. Therefore, the atoms share two electrons each to form the oxygen molecule. Since two electron pairs are shared, there is a double bond between the two oxygen atoms.



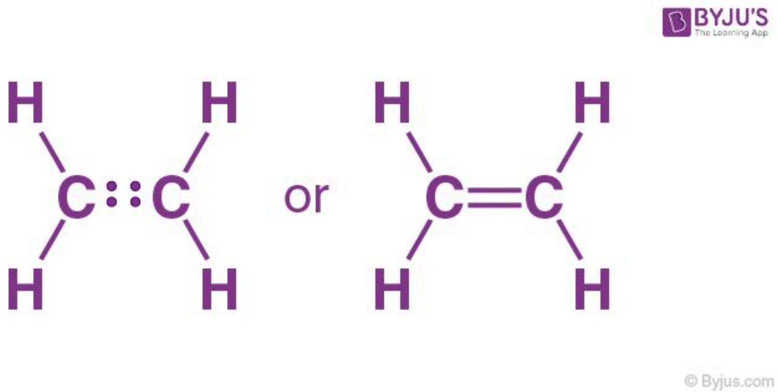
2. **Formation of Carbon (IV) oxide (CO_2):** a carbon (Iv) oxide molecule has one carbon atom with six valence electrons and two oxygen atoms with four valence electrons.

To complete its octet, carbon shares two of its valence electrons with one oxygen atom and two with another oxygen atom. Each oxygen atom shares its two electrons with carbon, and therefore there are two double bonds in CO2.



3.

Formation of Ethene Molecule C_2H_4: In ethylene, each carbon atom shares two of its valence electrons with two hydrogen atoms and the remaining two electrons with the other carbon atom. So, there is a double bond between the carbon atoms.



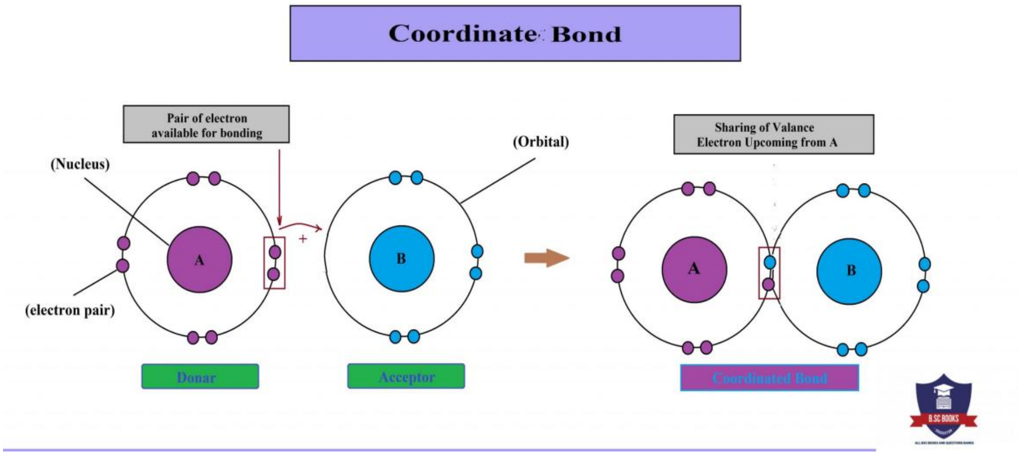
Triple Bond: A triple bond is formed when three pairs of electrons are shared between the two participating atoms. Triple covalent bonds are represented by three dashes (≡) and are the least stable type of covalent bonds.
For example,
formation of a nitrogen molecule: Each nitrogen atom having five valence electrons provides three electrons to form three electron pairs for sharing. Thus, a triple bond is formed between the two nitrogen atoms.

Other examples of molecules that contain ordinary covalent bonds include CO, NH_3, CH_4, etc.

Coordinate/Dative Covalent Bonding

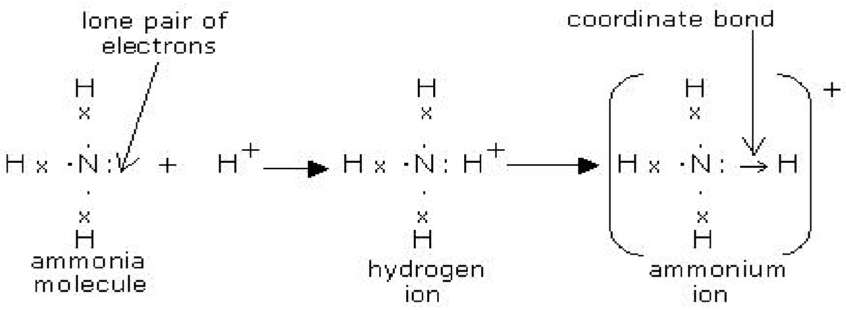
In this type of covalent bonding, there is also a sharing of electrons but the shared electrons (or lone pair of electrons) are contributed by only one of the participants. Thus, one of the reactants of a dative bond must have a lone pair of electrons. Examples of compounds that have the coordinate covalent bond the ammonium ion (NH_4^+), Oxonium ion (H_3O^+), hexacynaoferrate ion $[Fe(CN)6]^{3+}$, etc. In simple diagrams, a coordinate bond is shown by an arrow. The arrow points from the atom donating the lone pair to the atom accepting it.

The mechanism of coordinate bond is shown below



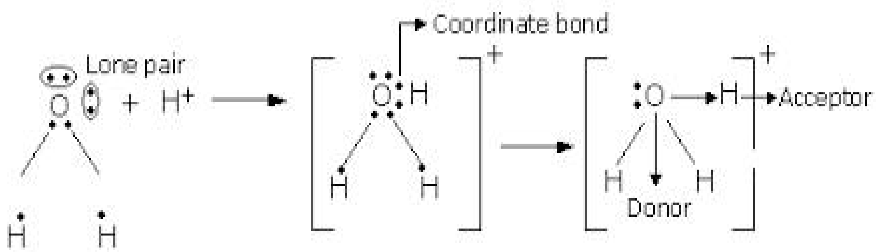
Examples

1. Formation of the ammonium ion: This involves combinationof hydrogen ion with ammonia. Hydrogen ion is electron deficient and requirestwo electrons to attain a duplet state while nitrogen in ammonia has a lone pair of electrons that can be shared with hydrogen ion.



As all the four hydrogen atoms in ammonium ion(NH_4^+)are not the same, the hydrogen ion which coordinated with the ammonia molecule is identified by an arrow which originates from the donor nitrogen atom and points to the coordinated hydrogen ion.

2. Formation of oxonium ion: the same principle as the one illustrated in the ammonium ion applies.

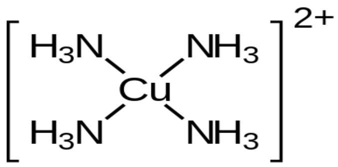


3. Formation of tetra amine copper (II) ion: the Cu^{2+} first coordinates with ammonia NH_3 according to the equation

$Cu^{2+} + NH_3 \rightarrow [Cu\text{-}NH_3]^{2+}$

In order for copper, Cu, to attain octet configuration, four molecules of NH_3 must coordinate with it as follows.

$Cu^{2+} + 4NH_3 \rightarrow [Cu(NH_3)_4]^{2+}$

ie 

Characteristics/properties of covalent compounds

Structure: Covalent compounds consist of molecules and they can exist in all three states of matter (solid, liquid or gas). This is because they are made up of molecules and thus the force of attraction between these molecules are weak and so they exist in all three states of matter.

Melting and boiling points: The melting and boiling points of covalent compounds are usually low. This is because they are made up of molecules that are held together by the weak force of attraction thus less heat is required to break the force of attraction between these molecules and so they have low melting and boiling points.

Solubility: Covalent compounds are not soluble in water but they are soluble in organic solvents like toluene and benzene. But this is the case with non-polar covalent compounds, polar covalent compounds are soluble in water.

Electrolytes: They are mostly non-conductor of electricity. This is because they are comprised of molecules, and, due to the absence of free mobile ions in these compounds electricity can not pass through them.

Non-polar covalent compounds do not ionize when dissolved into the water but polar covalent compounds like Hydrogen chloride, Ammonia etc. forms ions when dissolved into water and acts as electrolytes.

Differences between electrovalent and covalent compounds

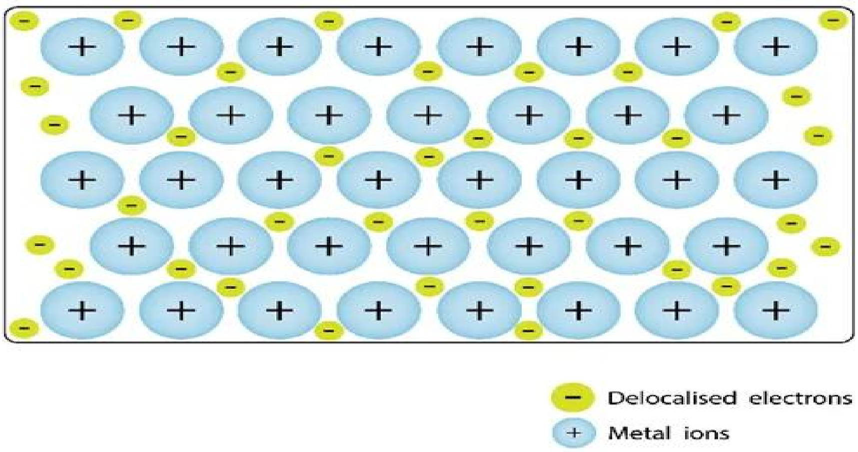
Table 5.4 Difference between Ionic and Covalent compounds

Ionic Compounds	Covalent Compounds
Formed by the transfer of electrons from a metal to a non-metal atom	Formed by sharing of electrons between non-metal atoms
Strong electrostatic force of attraction between cations and anions	Mutual sharing of electrons and so weak force of attraction between atoms
Solids at room temperature	Gases, liquids and soft solids
Conducts electricity in molten state or in solutions	Non-conductors of electricity
Have high melting and boiling points	Have low melting and boiling points
Soluble in polar solvents	Soluble in non-polar solvents
Hard and brittle	Soft and waxy
Undergo ionic reaction which are fast and instantaneous	Undergo molecular reactions which are slow

Other Binding Forces

Metallic bond: this is the force of attraction holding the atoms of a metal together in a crystal lattice. The electrons are responsible for binding the nuclei together to prevent repulsion. The higher the number of these sea of electrons per atom in the crystal lattice, the stronger is the metallic bond and hence the higher the melting point. In metals, the outer electrons are not attached to any individual metal atom. This is why they are referred to as 'delocalised', meaning they are free to move throughout the lattice.

Despite having positive metal ions and negative electrons, metals do not form ionic bonds. Instead, they form **metallic bonds**, which arise from the electrostatic force of attraction between the positive metal ions and the negative delocalised electrons. This bond holds the metal atoms together.

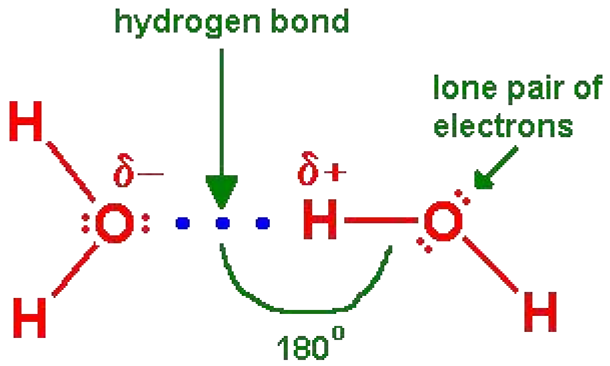


Properties of Metals

1. Metals are malleable, meaning they can easily be shaped. This is due to the ability of the metal's layers of atoms to slide over each other. This allows metals to be bent and hammered into various shapes.
2. Metallic bonds are very strong, so it takes a large amount of energy to break these bonds. As a result, metals have high melting and boiling points. All metals are solid at room temperature, except mercury, which is a liquid.
3. Delocalised electrons in metals can move freely, carrying electric charge and thermal energy. Therefore, metals are great conductors of electricity and heat.
4. Most metals do not dissolve in water. However, certain metals react with water, producing metal hydroxides and hydrogen gas.

Hydrogen bonds: this is the type of bond that occurs when hydrogen is covalently bonded with strongly electronegative elements such as oxygen, nitrogen, or fluorine. These highly electrongative elements have the tendency to attract the electrons to themselves more than hydrogen. Hence, compounds of the elements in which they are linked to hydrogen usually manifest unequal sharing of electron.

This gives rise to slight polarity or dipole between the atoms known as permanent dipole.



The hydrogen bond is normally dotted (ie …) to indicate that it is much weaker than the normal covalent bond. Hydrogen bonds are easily broken by thermal agitation.

The presence of hydrogen bonds in some molecules account for their high melting and boiling points

Van Der Waals forces: this is the weak attractive forces of attraction that exist between discrete molecules. They are important in the liquefaction of gases and in the formation of molecular lattices like iodine and naphthalene crystals.

Assignment

Use diagrams to illiustrate the types of bonds in the following; $Cacl_2$, Al_2O_3, CO, NO_2, CH_4, NH_3, Br_2, Cl_2.

Evaluation

1. Define chemical bonding.
2. Classify chemical bonds into strong and weak bonds, providing examples of each.
3. Describe the types of bonds found in various compounds.
4. What are the characteristics of ionic compounds?
5. What are the characteristics of covalent compounds?
6. Compare and contrast ionic and covalent compounds in terms of volatility.
7. Compare and contrast ionic and covalent compounds in terms of solubility and electrical conductivity.
8. Compare and contrast ionic and covalent compounds in terms of melting point and boiling point