CHEMISTRY NOTE FOR SS1 FIRST TERM 2026/2027 ACADEMIC SESSION

SCHEME OF WORK

LESSON ONE: INTRODUCTION TO CHEMISTRY

LESSON TWO: NATURE OF MATTER

LESSON THREE: SEPARATING TECHNIQUES

LESSON FOUR: PURIFICATION OF SUBSTANCES

LESSON FIVE: SYMBOLS, FORMULAS AND EQUATIONS

LESSON SIX: PARTICULATE NATURE OF MATTER

LESSON SEVEN: LAWS OF CHEMICAL COMBINATION

LESSON EIGHT: CHEMICAL COMBINATION

INTRODUCTION TO CHEMISTRY

Chemistry is a branch of science which deals with the study of the nature, composition, properties and uses of matter as well as the changes that matter undergo under different conditions. Studying chemistry gives us the required training in the scientific method which will help us become scientists.

Our environment is made up of matter and therefore explains activities that are of chemical changes. Examples are:

1. Lighting a match
2. Burning of firewood
3. Cooking
4. Palm wine turning sour after sometime
5. Rusting of nails
6. Rottening of leaves.
7. Slaking of lime (addition of water to calcium oxide/ quicklime)

The above listed points are examples of chemical reactions. Some things we use everyday are made by chemical process. e.g.

- Soap and detergent for cleaning
- Hair cream and perfume for grooming
- Oil and margarine
- Plastics

Chemistry may be divided into 3 branches namely:

(1) ***Inorganic Chemistry***: deals with the study of matter present in non-living things.
(2) ***Organic Chemistry***: deals with the study of matter found in both plants and animals (living things)
(3) ***Physical Chemistry:*** deals with the study of energy changes that accompanies transformation of matter.

THE SCIENTIFIC METHOD

Chemistry is an experimental science. Scientists are very alert and inquisitive. They use their senses to observe happenings around them giving rise to a set of ***observations*** which in turn form a certain ***pattern***. This often leads to a problem which they try to solve. They put up a reasonable explanation or ***hypothesis*** and they carry out ***experiments*** to test it. Then they carefully record the observations and the results of their experiment.

If the experiment supports the hypothesis, they carry out further investigations. When the hypothesis has been tested extensively and found to be correct within limits of available evidence, it becomes a ***theory***. A ***scientific law or principle*** is established after a theory has been extensively tested and proven to be true without any exception. If the experiments give negative results, the scientist goes back to his hypothesis and either modifies it or puts forward a new one. The method of studying a problem is known as the ***scientific method*** and it is the basis of all scientific research.

Flow chart of The Scientific Method

Hypothesis: This is a tentative, testable and reasonable explanation for a phenomenon or a problem based on observations and prior knowledge

Theory: This is a well substantiated explanation of some aspects of the natural world. A theory organizes well tested hypotheses into coherent and logical framework.

Scientific law/principle: This is a generalized rule that summarizes observations of consistent natural behavior. It is a concise descriptive statement or mathematical equation that details a consistent, observable pattern in nature.

APPLICATIONS OF CHEMISTRY TO LIFE

Chemistry has contributed greatly to the improvement of the quality of life and provision of our basic needs:

(1) **Food**: The manufacture of fertilizers and insecticides has helped to improve food production. Food preservation techniques have helped to store food for long periods and malnutrition is being eradicated by the manufacture of foods enriched with essential vitamins\nutrients.

(2) **Clothing**: Manufacture of cheap man-made textile fibres like artificial silk and rayon is possible.

(3) **Military and Space science**: The discovery of ammunition helps the military to attain specific goals. The launching of space rockets is made possible by chemistry

(4) **Housing**: Building materials like cement, concrete, steel, tiles, and bricks are produced by chemical industries even paints for beautifying our homes.

(5) **Medicine**: Drugs, antibodies produced by pharmaceutical industries help to improve the life of sick people.

(6) **Transportation**: Cars, trains, aircrafts are produced with the knowledge of chemistry, suitable fuels and structural materials.

(7) **Energy**: Alternative sources of energy has been discovered e.g biogas and nuclear energy

ADVERSE EFFECTS OF CHEMISTRY

Chemistry has also affected life in the following ways

(1) **Corrosion of iron**: rusting of iron means slow deterioration of iron.
(2) **Pollution**: Chemical processes have led to the pollution of our environment e.g. industrial wastes, nuclear wastes, oil spillage.
(3) **Drug abuse**: drugs like heroin, cocaine and morphine are addictive so people tend to abuse them.
(4) **War**: Scientific practices (chemical processes) has led to the production of weapons of destruction e.g gun powder, guns, nuclear bombs
(5) **Abortion**: This leads to death of an unborn child and sometimes that of the mother.
(6) **Euthanasia/ mercy killing**: This is the use of chemical products to kill patients with terminal disease without allowing natural death to take its course.

ROLE OF GOVERNMENT IN PREVENTING CHEMICAL DEGRADATION

1. Government should enact laws to curtail indiscriminate disposal of chemical wastes and pollutants, sale and use of illicit drugs or illegal chemical products.
2. They should establish regulatory bodies to ensure and enforce the law.
3. They should organize public enlightenment to educate the citizens on the adverse effects of chemical products.

CAREER OPPORTUNITIES IN CHEMISTRY

There are many career opportunities in chemistry. They are

1. **Teaching service**: chemistry teachers / lecturers and laboratory assistants.
2. **Health service**: pharmacists, biochemists, chemists, nutritionists, dietitians, doctors nurses, medical assistants, dispensaries and laboratory assistants.
3. **Food processing**: food technologists and research chemists
4. **Petroleum and petrochemical industries**: application chemists, research chemists, chemical engineers and laboratory assistants.
5. **Manufacturing industries**: research chemists and chemical engineers.
6. **Extractive industries**: geologists, chemists and mining engineers.
7. **Agriculture and Forestry**: agricultural scientists, chemists, physiologists and biochemists.

NATURE OF MATTER

Matter is a substance that has mass and occupies space. The S.I unit of mass is the mole but for practical purposes, mass is measured in kilogram with a weighing balance.

Substances can be identified by the characteristics they possess. These characteristics are called properties. Properties can either be physical properties or chemical properties which accompany physical change or chemical change respectively.

Matter exists in three states namely: **SOLID**, **LIQUID** and **GAS.**

Water is an example of a substance that can exist in the three states.

ICE WATER STEAM

PHYSICAL AND CHEMICAL CHANGES

Matter undergoes changes which may either be temporary or permanent.

Physical change is the type of change which is easily reversible and in which no new substance is formed. Examples include:

1. Changes in the state of matter(melting, evaporation, condensation, freezing).
2. Separation of mixtures by physical means (filtration, sublimation, crystallization e.t.c)
3. Magnetization and de-magnetization of iron.

Chemical change is the type of change which is not easily reversed and a new substance are formed. Examples include.

1. The dissolution of metals in water.
2. The rusting of iron
3. The addition of water to quick lime (slaking of lime).
4. Fermentation and decay of substances.
5. Changes in an electrochemical cell.

Comparison between Physical and Chemical Change

Physical change	**Chemical change**
It is easily reversible	It is not easily reversible
No new substance is formed	New substances are formed
There is no change in the mass of the substance	There is change in the mass of the substance
It does not involve any great heat changes	A considerable amount of heat is involved

ELEMENTS, COMPOUNDS AND MIXTURES

Matter can be classified as elements, compounds and mixtures.

ELEMENT: An element is a substance which cannot be split into simpler units by an ordinary chemical process. There are 119 known elements; ninety of them are occur in nature while the rest are made artificially. Elements are classified into metals (Iron, Aluminum, Zinc), non-metals (Oxygen , Hydrogen, Chlorine) and metalloids (semi-metals e.g Silicon, Germanium).

Student's exercise: Write 6 differences (3 physical and 3 chemical properties) between metals and non-metals

COMPOUNDS: A compound is a substance which contains two or more elements chemically combined together. A compound is a result of chemical change and has entirely different properties from the constituent elements e.g water is formed from hydrogen and oxygen which are gases but water is a liquid.

Compound	**Constituent elements**
Water (H_2O)	Hydrogen and Oxygen
Marble ($CaCO_3$)	Calcium, Carbon and Oxygen
Caustic potash (KOH)	Potassium, oxygen and Hydrogen
Sand (SiO_2)	Silicon and Oxygen
Glucose ($C_6H_{12}O_6$)	Carbon, Hydrogen and Oxygen
Ethanol (C_2H_5OH)	Carbon, Hydrogen and Oxygen
Ethanoic acid (CH_3COOH)	Carbon, Hydrogen and Oxygen
Calcium chloride ($CaCl_2$)	Calcium and Chlorine
Bleaching powder (CaOCl)	Calcium, Oxygen and Chlorine
Tetraoxosulphate(VI) acid (H_2SO_4)	Sulphur, Oxygen and Hydrogen
Table salt (NaCl)	Sodium and Chlorine

MIXTURE: A mixture is a substance that contains two or more constituents which can be separated by physical methods. The constituents of a mixture can be elements or compounds or both. The constituents of a mixture retain their individual identities because their physical and chemical properties are not changed by simple mixture.

Mixture	**Constituents**
Air	Oxygen, Carbon(IV) Oxide, Nitrogen, Rare gases, dust, Moisture
Soil	Sand, Clay, Humus, Water, Air, Mineral Salts
Urine	Urea, water, Mineral Salts
Palm wine	Water, Sugar, Alkanol, Mineral salt, Vitamin, yeast, Protein, Fat
Coca-cola	Water, Sugar, Carbon(IV)oxide, Coca-cola concentrate
Milk	Water, Sugar, Fat, Proteins, Mineral salts, Vitamins
Sea- Water	Water, Mineral salt, Bacteria, Remains of organic matter
Blood	Water, Proteins, Fat, Oil, Sugar, Mineral salts, vitamins, Hormones, Enzymes, Blood cells, Haemoglobin
Crude oil	Petrol, Heavy oil, Gas oil, kerosene, Naphtha, Bitumen, Gas e.t.c
Brass	Copper and Zinc
Bronze	Copper and Tin
Amalgam	Mercury and alkali metals

Comparison between Compounds and Mixtures

Compounds	Mixtures
The Constituents cannot be separated by physical means	The constituents can be separated by physical means.
It is always homogeneous.	It may be homogeneous or heterogeneous
The properties of a compound differ entirely from that of its component elements	The properties of a mixture are the sum of the properties of its individual constituents.
A compound can be represented by a chemical formula	A mixture cannot be represented by a chemical formula
Formation of a compound involves a remarkable amount of heat	Formation of mixture does not readily need a great amount of heat.
The composition is fixed	The composition is varied

SEPARATING TECHNIQUES

To study the nature and composition of a substance, the substance must be in a pure form i.e. it must not contain unwanted substances called impurities. The impurities are not chemically combined with the substance and hence, ***impure substances are regarded as mixtures***. Since the components of a mixture retain their physical and chemical properties, hence it is possible to separate the constituents of a mixture based on the differences in their individual properties.

The individual properties involved in the separation of a mixture include:

1. Size of a particle
2. Solubility of the substances in solvents.
3. Boiling point
4. Density
5. Miscibility
6. Magnetic property
7. Sublimability
8. Rate of movement of solute over an absorbent medium

Separating technique	Substances to be separated	Physical property
Hand picking	Mixture of solids	Colour and size
Sieving	Solids of different sizes	Size
Filtration	Insoluble solids from liquids	Solubility
Decantation, Centrifugation	Insoluble solids from liquids	Density
Evaporation	Soluble solids from solution	Solubility
Crystallization, Fractional crystallization	Soluble solids from solution	Solubility
Distillation	Solvent from solution	Boiling point
Fractional distillation	Mixture of miscible liquids	Boiling point
Use of separating funnel	Mixture of immiscible liquids	Miscibility and density
Chromatography	Solutes from solution	Rate of movement
Sublimation	Mixture of solids in which one can sublime	Sublimation
Magnetization	Magnetic solids from non-magnetic ones	Magnetic property

SIEVING: This is a technique used to separate a mixture of solid particles of different sizes. The mixture is placed on a mesh of a particular size, the smaller particles which are smaller than the mesh (sieve) size will pass through while the bigger particles remain on the sieve e.g separation of a mixture of sand and stones. The physical property employed is particle size. The technique is applied in mining and garri-making industries.

MAGNETIC SEPARATION: This is a technique used to separate magnetic substances from non-magnetic particles in a mixture by moving a bar magnet over the mixture. The magnetic substances are attracted to the magnet while the non-magnetic ones are not attracted e.g. separation of a mixture of iron fillings and sulphur. The physical property employed is magnetic property. This separation method is applied in the steel industry to remove magnetic impurities from tin ore. Magnetic substances include iron, nickel and steel.

SUBLIMATION: This method is used to separate a mixture of substances in which one can sublime. Sublimation is a process by which a substance changes directly from the solid state to the gaseous state without passing through the liquid state. Examples of such substances are Iodine (I_2), naphthalene and ammonium chloride (NH_4Cl), Iron (III) chloride ($FeCl_3$). These substances can be purified industrially by this method. The physical property employed is sublimability.

SEDIMENTATION AND DECANTATION: This method is used to separate a mixture containing liquid and solid particles which separate into two distinct layers on standing e.g. a mixture of sand and water. The denser substance settles at the bottom of the container (sedimentation) while the liquid portion stays on top. The clear liquid (upper layer) is carefully poured out or decanted into a second container using a glass rod. The physical property employed is density

FILTRATION: This is a method used to separate a mixture of insoluble solid from a liquid by means of a filter. The filter used in chemistry laboratories is a special type of paper known as the filter paper. e.g. a mixture of chalk dust in water or sand and water. The mixture is poured into a funnel containing a cone of the filter paper. The liquid known as the **filtrate** passes through the pores of the filter paper and drips into the flask below while the substances retained in the filter paper is called **residue**. The physical property employed is solubility. This method is applied in the purification of pipe-borne water and in the treatment of sewage.

CENTRIFUGATION: Centrifugation is a separating method used to separate insoluble solids from a liquid where the liquid particles are so small that they can pass through the pores of a filter paper. Centrifugation is carried out with a machine called centrifuge. It spins test tubes containing mixtures at high speed so that mixtures can be separated based on differences in their density. The solids settle while the liquid separates out as an upper layer which can easily be decanted. Centrifugation is used to separate blood cells from the plasma in diagnostic laboratories and hospitals.

EVAPORATION: This is a method used to separate a mixture of a soluble solid from its solvent by heating the mixture so that water escapes into the atmosphere. A mixture of a soluble solid in a solvent is called a solution. A mixture of table salt and water is called a salt solution. The salt solution is poured into an evaporating dish and heated over a sand bath or water bath/ steam bath to ensure a steady or uniform rate of evaporation. The solid is left behind in the dish while the solvent escapes into the atmosphere. The physical property employed is solubility. This separation method is employed in salt-making industries.

CRYSTALLISATION: This is a method used to obtain crystals from a saturated solution without evaporating the solvent completely. It is used to separate crystalline salts which decompose easily on heating from its solution. The salt crystals obtained in this form are pure and usually contain water of crystallization e.g Copper (II) tetraoxosulphate (VI) pentahydrate, $CuSO_4.5H_2O$; Copper (II) trioxonitrate (V) trihydrate $Cu(NO_3)_2.3H_2O$; Iron (II) tetraoxosulphate (VI) pentahydrate $FeSO_4.5H_2O$. steps in crystallization

a. The salt solution is first heated to evaporate some solvent, the solution then becomes saturated.
b. The saturated solution is then placed in a bowl of cold water. When the solution is cooled, crystals of the solute begin to form
c. The crystals formed is filtered off and dried between filter paper.

To induce crystal formation, we can either scratch the inside of the container or add crystals of the salt into the solution as seeds.

The physical property employed in crystallization is solubility. Crystallization is applied in drug and sugar production.

FRACTIONAL CRYSTALLIZATION: This method is used to separate two or more soluble solutes (solid particles) which are present in the same solution in roughly equal amounts. The solubilities of the different solutes in a given solution must differ at different temperatures. As the solution is being cooled, at a particular temperature, crystals of the relevant solute will come out of solution leaving behind the others which are still in their limits of solubility. Sometimes fractional crystallization is also known as **recrystallisation**

Recrystallisation is required to obtain a purer non volatile sample. The advantages of recrystallisation include high purity and low cost. The procedure for recrystallisation include

a. Dissolving the solute in a hot solvent
b. Perform a gravity filtration
c. Cooling the mixture to enhance crystal formation
d. Performing a vacuum filtration
e. Drying the crystals

PRECIPITATION: A precipitate is a solid formed as a result of a chemical reaction. This method is used to separate a solid with different solubility in two different but miscible solvents. One of the solvents is used to separate or precipitate out the solid when it is dissolved in them e.g Iron (II) tetraoxosulphate (VI)

is soluble in water but insoluble in ethanol. If ethanol which is soluble in water is added to a solution of Iron (II) tetraoxosulphate (VI) in water, the Iron (II) tetraoxosulphate (VI) will be precipitated out of the solution and can be easily separated by filtration.

DISTILLATION: This method is used to separate a mixture of two or more miscible solvents with large difference in their boiling points. It is used to recover a pure solvent from a solution. A thermometer is required here. The solution is first heated in a flask to vaporize the solvent, instead of escaping into the atmosphere, the vapour is cooled by passing it along a condenser. The vapour re-condenses back into a liquid called the **distillate** which is collected in a receiver (flask). Distillation can be defined as the process of vaporizing a liquid and re-condensing the vapour to liquid. A mixture of ethanol (Boiling point 78°C) and water (Boiling point 100°C) can be separated with this method. The solute and other impurities can be left behind in the flask. The physical property employed boiling point. This method is applied in gin distilleries and water distilleries for production of gin and distilled water respectively.

FRACTIONAL DISTILLATION: Fractional Distillation is used to separate two or more miscible liquids with close boiling points e.g boiling points of not more than 10°C between successive fractions. Fractional Distillation involves separation of mixture into components or fractions. They distill in ascending order of their boiling point starting with the lowest. The apparatus for fractional distillation is the same with that of simple distillation except that a fractionating column (a column packed with glass beads) is introduced which does the actual separation. The upper part is at a lower temperature than the lower part. Only the vapour with the same temperature as the upper part of the column passes into the condenser. Vapours with higher boiling condense as they enter the upper part of the fractionating column and flow back into the flask. The process is repeated until the fraction with lowest boiling point distills over completely and then the temperature of the fractionating column rises for the distillation of the next fraction with a higher boiling point. This continues until all the components have been separated. The physical property employed is boiling point. This method of separation is applied in the petroleum industry for the distillation of crude oil and also for production of nitrogen and oxygen(fractional distillation of liquid air).

SEPARATING FUNNEL: This method is used to separate two liquids which are immiscible (a polar and a non polar solvent) e.g a mixture of petrol or kerosene and water. The mixture is poured into separating funnel, the less dense liquid which is petrol /kerosene will float on top while more dense liquid (water) sinks to the bottom. The lower layer is tapped into a container leaving the upper layer in the funnel. The physical property employed is density and miscibility. This method is applied in the separation of slag from molten iron during the extraction of iron in the blast furnace

CHROMATOGRAPHY: This method is used to separate a mixture of solutes by taking advantage of their different rate of movement in a solvent over an adsorbent material .e.g paper. There are many types chromatography which includes Paper chromatography (used to separate dyes and mixtures of coloured substances), Thin Layer chromatography, Gas chromatography (used to separate complex mixtures of gases), column chromatography (used to separate liquid extract from leaves, flowers, dyes and mixtures of protein).

In paper chromatography, the adsorbent material is a strip of filter paper or chromatographic paper while the solvents can be ethyl ethanoate, chloroform or ethanol. A faint pencil line of about 5cm is drawn across the breadth of the filter paper. A black ink (the mixture) is spotted at the center of the line and allowed to dry. The paper is suspended in a closed air-tight container with the spotted end (but not the spot) dipping in the solvent. The solvent then moves up dragging the different solutes in the mixture at different speeds. The strip is removed from the jar when the solvent has moved three-quarters way up. It is dried and sprayed with appropriate reagent to locate the position of each solute in the mixture. Each solute can be identified by the distance it has travelled. The rate at which the different solutes move up depend on

1. The adsorption of the solutes by paper
2. The dissolution of the solute by the solvent

As the solvents ascend the paper, the solutes that are strongly adsorbed by the paper find it difficult to move up while the ones that are weakly adsorbed by the paper dissolves in the solvent and readily move up the column. Separation is based on differences in the degree of adsorption of the components by the stationary phase or the degree of retention of the components by the mobile phase.

In chromatography, 2 phases are involved. They are:

1. Stationary phase e.g paper on which solutes are adsorbed
2. A mobile phase e.g solvent which dissolves the solute and carries them

Thin layer chromatograms are analysed by calculating their retention factor, R_f values for each of the component. Each component can be identified by comparing its R_f values with known values recorded using same materials and conditions.

Retention factor = $\frac{\text{Distance moved by component}}{\text{Distance moved by the solvent front}}$

Chromatography can be used to separate the components of ink, dye, proteins e.t.c. Chromatography is applied in food industries and diagnostic laboratories.

Student's exercise:

1. Describe how you separate

a. mixture of NH_4Cl, NaCl and $BaSO_4$

b. sugar contaminated by sand

2. differentiate between adsorption and absorption

PURIFICATION OF SUBSTANCES

An impure substance is a mixture and must be separated using any of the separation techniques discussed earlier.

CRITERIA FOR DETERMINING/TESTING OF PURITY OF SUBSTANCES

When a substance has been separated, it is important to know if it is pure or not. The purity of substances can be determined by:

1. Fixed/constant boiling point of a liquid: The boiling point of a liquid is the temperature at which the maximum vapour pressure of a given liquid is equal to the external pressure. This temperature is affected by impurities causing the substances to boil over a temperature range. Impurities increase the boiling point of liquids.

2. Fixed/constant melting point of a solid substance: this is the constant temperature at which solids and liquids are in equilibrium at a given constant pressure. A pure solid will melt completely over a narrow/sharp temperature range. Impurity depresses/decreases melting point and usually causes gradual softening instead of sharp liquefaction.

3. A pure substance gives only one spot in a chromatogram. This is used to check the purity of coloured substances like dyes and pigments.

4. Density: The higher the density of substances, the greater the purity.

5. Refractive index: Pure substances have high refractive index than impure substances

DETERMINATION OF THE BOILING POINT OF A LIQUID

Aim: To determine the boiling point of water.
Method: Put some distilled water in a test tube and place a thermometer in the water. Heat the water till it boils and record the highest constant temperature.

In case of distilled water(Pure water), the boiling point will be 100 °C while that of tap water may be over 100°C because of impurities. For pure ethanol, the boiling point is 78°C. It is heated in a water bath. Ethanol with impurities will boil at a temperature higher than 78°C.

Result: The water boiled at 100°C

Conclusion: The boiling point of water is 100°C.

DETERMINING THE MELTING POINT OF A SOLID.

Aim: To determine the melting point of Naphthalene .

Method: The substance (naphthalene) is filled in a capillary tube to a height of about 5mm. Attach the tube to the lower part of the thermometer by a thin rubber band. Lower the thermometer into the heating medium (Paraffin oil). Heat gently until the solid melts. The constant temperature at which the solid changes to liquid is known as the Melting point.

A solid with impurity melts at a lower temperature. Pure naphthalene melts at 80°C but impure naphthalene may melt at temperatures below 80°C due to impurities.

Result: Naphthalene melted at 80°C

Conclusion: The melting point of naphthalene is 80°C.

Student's exercise:

1. Describe an experiment to determine the boiling point of ethanol (boiling point of ethanol is 78°C).

2. Describe an experiment to determine the melting point of sodium (melting point of sodium is 97°C).

ATOMIC THEORY

Matter is made up of discrete particles known as atoms, molecules and ions.

ATOM: An atom is the smallest particle of an element which can take part in a chemical reaction. It is also defined as the smallest part of an element that can ever exist and still retain the chemical properties of that element. John Dalton, an English Chemist, put forward a theory to describe the nature of an atom.

MOLECULE: A molecule is the smallest particle of a substance that can normally exist alone and still retain the chemical properties of that substance, be it an element or a compound. Most atoms cannot exist alone. They bond with other atoms to form molecules. Molecules may be made up of atoms of the same element or different elements.

ATOMICITY is the number of atoms present in one molecule of an element. Elements can be

a. **Monoatomic**: These elements contain only one atom in one molecule. e.g. most metals and rare gases like Neon Ne, Argon Ar, Helium He, Sodium Na
b. **Diatomic**: These elements contain two atoms in a molecule e.g. Hydrogen H_2, Oxygen O_2, Nitrogen N_2, Chlorine Cl_2, Bromine Br_2, Iodine I_2
c. **Triatomic**: These elements contain three atoms in a molecule e.g Ozone O_3
d. **Polyatomic**: these elements contain many atoms in one molecule e.g. Phosphorus P_4 (tetratomic), Sulphur S_8

Sometimes, the atoms in a molecule of a compound may be different elements e.g Hydrogen chloride contains one atom of hydrogen and one atom of chlorine. Most compounds exist as molecules.

ION: An ion is any atom or group of atoms which possess an electric charge due to loss or gain of electrons. There are two types of ions:

a. **Cations** or positively charged ions: are atoms or group of atoms which possess a positive charge e.g H^+, Na^+, K^+, Ca^{2+}, Al^{3+}, Cu^{2+}, Mg^{2+}
b. **Anions** or negatively charged ions: are atoms or group of atoms which possess a negative charge e.g Cl^-, F^-, Br^-, O^{2-}, S^{2-}, N^{3-}

Ionic compounds are compounds with equal number of positive and negative ions and so it is electrically neutral e.g Na^+Cl^-, $Mg^{2+}O^{2-}$, $Mg^{2+}S^{2-}$, $N^{3-}3H^+$

DALTON'S ATOMIC THEORY

In 1808, John Dalton proposed the atomic theory, though the theory was partially supported by experimental evidences such as Law of conservation of mass, Law of Definite proportions and Law of Multiple proportions. However, it could not explain electrolysis and some other phenomena. As a result, the theory has been modified due to new discoveries based on experimental evidences.
The theory, their limitations ant the modifications are summarized as follows:

S/No	DALTON'S POSTULATE	LIMITATION	MODIFICATION
1.	All elements are made up of small indivisible particles called atoms	The atom is not an indivisible piece as proven wrong by Rutherford's discovery.	Atoms are made up of 3 main types of sub-particles: electron, proton and neutron.
2.	Atoms can neither be created nor destroyed	This statement holds for ordinary chemical reactions but not for nuclear reactions	Atoms can neither be created nor destroyed during nuclear reactions in which atom of an element undergo changes to

			form atoms of other element with the release of energy
3.	Atoms of the same elements are alike in every aspect and differ from those of other elements	The discovery of isotopes makes this statement unacceptable.	Isotopes are atoms of the same element which differ in their neutron content and relative atomic masses but they have the same atomic number and chemical properties. e.g ^{35}Cl and ^{37}Cl
4.	When atoms combine with other atoms, they do so in simple ratios.	This statement is true for inorganic reactions involving a few atoms per molecule but cannot hold for complex organic reactions involving thousands of atoms per molecule	Complex organic reactions involving thousands of atoms per molecule does not occur in simple whole number ratio
5.	All chemical changes result from the combination or separation of atoms.	-	-

SYMBOLS OF ELEMENTS

In 1814, Berzelius suggested a simple system of representing elements with symbols. This system involved representing elements letters of the alphabets. Usually, it contains the first letter or the first two letters or the first letter and another letter or the letters from the Latin name of an element.

A chemical symbol is a sign made up of one or more letters used to represent an element.

Atomic number	Element	Latin name	Symbol
1	Hydrogen	-	H
2	Helium	-	He
3	Lithium	-	Li
4	Beryllium	-	Be
5	Boron	-	B
6	Carbon	-	C
7	Nitrogen	-	N
8	Oxygen	-	O
9	Fluorine	-	F
10	Neon	-	Ne
11	Sodium	Natrium	Na
12	Magnesium	-	Mg
13	Aluminium	-	Al
14	Silicon	-	Si
15	Phosphorus	-	P
16	Sulphur	-	S
17	Chlorine	-	Cl
18	Argon	-	Ar
19	Potassium	Kalium	K

20	Calcium	-	Ca
24	Chromium	-	Cr
25	Manganese	-	Mn
26	Iron	Ferrum	Fe
28	Copper	Cuprum	Cu
30	Zinc	-	Zn
47	Silver	Argentum	Ag
50	Tin	Stannum	Sn
51	Antimony	Stibium	St
79	Gold	Aurum	Au
80	Mercury	Hydragyrum	Hg
82	Lead	Plumbum	Pb

FORMULAE OF COMPOUNDS

A chemical formula is a representation of a molecule of a substance by the symbols of its constituent elements. Where the element exists as a molecule, a number representing its atomicity is written as a subscript after the symbol of that element. e.g Oxygen, Nitrogen, Hydrogen and Chlorine are diatomic and are written as O_2, N_2, H_2, and Cl_2 respectively. Ozone is triatomic and written as O_3, phosphorus is tetraatomic and written as P_4, sulphur is written as S_8.

A molecular formular is a representation of a molecule of a compound by the actual number of atoms of the different elements present in the compound. The symbols in a molecular formular are written together as a group and the number of atoms of each component element is written as a subscript after the symbol of that element. e.g
Calcium trioxocarbonate (IV) $CaCO_3$ contains 1 atom of Calcium, 1atom of Carbon and 3 atoms of Oxygen.
Hydrochloric acid HCl contains 1 atom of Hydrogen and 1 atom of Chlorine
Water H_2O contains 2 atoms of Hydrogen and 1 atom of Oxygen
Ammonia NH_3 contains 1 atom of Nitrogen and 3 atoms of Hydrogen
Tetrachloromethane CCl_4 contains 1 atom of Carbon and 4 atoms of Chlorine
Zinc Sulphide ZnS contains 1 atom of Zinc and 1 atom of Sulphur.

VALENCY: This is defined as the combining power of an element. Sometimes, it can be regarded as the oxidation number of an element. Metals have positive valencies while non-metals have negative valencies. The valencies of the elements in a compound play a very important role when writing the formula of that compound.
OXIDATION NUMBER: This is the charge on the valency of an element.
Oxidation state = valency + charge
The oxidation state of any compound is zero i.e when the valencies of the component elements are added together, it gives zero e.g in NaCl, the valency of Na is +1 while that of Cl is -1 resulting to zero. Na^+ Cl^- =+1 + (-1) = 0
Application of the oxidation number: The oxidation number is applied in the following cases

a. Writing and Balancing of equations of redox reactions
b. IUPAC Naming of compounds

RADICAL
A radical is a group of atoms which behave as a single unit. It is treated as a single unit in a molecular formular. e.g $-OH^-$, $-NO_3^-$, $-CO_3^{2-}$, $-PO_4^{3-}$, $-SO_4,^{2-}$, $-NH_4^+$. There are two types of radicals: the positive

radicals and the negative radicals. When more than one radical is present in a molecule of a compound, they are enclosed in parentheses with the relevant numerical subscript after the parentheses. e.g $Ca(NO_3)_2$ is not written as CaN_2O_6, $Fe(OH)_3$ is not written as FeO_3H_3 and $Al_2(SO_4)_3$ is not written as $Al_2S_3O_{12}$.

COMMON ELEMENTS AND RADICALS WITH THEIR VALENCIES AND OXIDATION NUMBER

Name	Formular of ion	Valency	Oxidation state
Hydrogen	H^+	1	+1 (or-1 in metal hydrides)
Potassium	K^+	1	+1
Sodium	Na^+	1	+1
Magnesium	Mg^{2+}	2	+2
Calcium	Ca^{2+}	2	+2
Aluminium	Al^{3+}	3	+3
Iron (II)	Fe^{2+}	2	+2
Iron (III)	Fe^{3+}	3	+3
Copper (I)	Cu^+	1	+1
Copper (II)	Cu^{2+}	2	+2
Zinc	Zn^{2+}	2	+2
Lead (II)	Pb^{2+}	2	+2
Lead (IV)	Pb^{4+}	4	+4
Chlorine	Cl^-	1	-1
Fluorine	F^-	1	-1
Bromine	Br^-	1	-1
Sulphur	S^{2-}	2, 4 or 6	-2, -4 or -6
Iodine	I^-	1	-1
Phosphorus	P^{3-}	3 or 5	-3 or-5
Oxygen	O^{2-}	2	-2 (or -1 in peroxides
Nitrogen	N^{3-}	3 or 5	-3 or -5
Carbon		2 or 4	+2 or +4

Name of Radical	Formular of radical	Valency	Oxidation state
Trioxonitrate (V) ion	NO_3^-	1	-1
Trioxochlorate (V) ion	ClO_3^-	1	-1
Tetraoxomanganate (VII) ion	MnO_4^-	1	-1
Hydroxide ion	OH^-	1	-1
Trioxocarbonate (IV) ion	CO_3^{2-}	2	-2
Hydrogen trioxocarbonate (IV) ion	HCO_3^{1-}	1	-1
Tetraoxophosphate (V) ion	PO_4^{3-}	3	-3
Hydrogentetraoxophosphate (V) ion	HPO_4^{2-}	2	-2
Heptaoxodichromate (VI) ion	$Cr_2O_7^{2-}$	2	-2
Tetraoxosulphate (VI) ion	SO_4^{2-}	2	-2
Hydrogen tetraoxosulphate (VI) ion	HSO_4^-	1	-1
Ammonium ion	NH_4^+	1	+1

CALCULATIONS ON OXIDATION NUMBER

Example 1: Calculate the oxidation number of Nitrogen in (a). N_2O (b). NH_3 (d). NO_2 (e). NO_3^- (f). HNO_3

Solution

a. N_2O The oxidation number is equal to zero

b. NO The oxidation number is equal to zero

c. NH_3 The oxidation number is equal to zero

d. NO_2 The oxidation number is equal to zero

e. NO_3^- The oxidation number is equal to -1 (the charge on the ion)

f. HNO_3 The oxidation number is equal to zero

Example 2: Calculate the oxidation number of sulphur in (a). $FeSO_4$ (b). SO_3^{2-} (c). H_2SO_4

a. $FeSO_4$ The oxidation number is equal to zero

b. SO_3^{2-} The oxidation number is equal to -2 (the charge on the ion)

c. H_2SO_4 The oxidation number is equal to zero

Students Exercise

1. Calculate the oxidation number of manganese in (a). MnO_2 (b). MnO_4^- (c). $KMnO_4$
2. Calculate the oxidation number of chromium in $K_2Cr_2O_7$
3. Calculate the oxidation number of sulphur in $Fe_2(SO_4)_3$ and SO_4^{2-}

WRITING OF FORMULA OF COMPOUNDS USING THE VALENCIES OF ELEMENTS OR RADICALS

1. Write the symbols of the elements or radicals in the compound
2. Write the valencies above and to the right of the symbols
3. Rewrite the symbols exchanging the valencies and write the number below and to the right of the symbols.

Note: The sum of the positive valencies must be equal to the sum of the negative valencies. The number of atoms of the component elements must be written as a numerical subscript after the element concerned. If it is a radical, it must be enclosed within a parenthesis.

Component element/ Radical	**Valency**	**Formula**
A	x	A_yB_x
B	y	

Examples

1. Sodium and oxygen

2. Iron and oxygen

3. Aluminium and tetraoxosulphate (VI) ion

4. Calcium and trioxonitrate (V) ion

Students exercise – Write the formular of the compounds formed between

a. Magnesium and chlorine
b. Carbon and oxygen
c. Hydrogen and oxygen
d. Copper and trioxocarbonate (IV) ion
e. Potassium and hydrogen tetraoxosulphate (VI) ion

IUPAC NOMENCLATURE OF INORGANIC COMPOUNDS

The IUPAC system uses oxidation number in naming compounds.

Binary compounds: These are compounds which contain two elements only. They are named as follows:

a. The name of the first element is written first
b. The name of the second element is written with the suffix –ide and attached to the first i.e oxygen – oxide, hydrogen – hydride, chlorine – chloride, phosphorus – phosphide, sulphur - sulphide, e.t.c
c. When the first element has more than one valency, the valency/oxidation number is written as a roman figure in bracket after the name of the first element

Examples:
- KCl - Potassium chloride
- FeO - Iron (II) oxide
- CO_2 - Carbon (IV) oxide
- CS_2 - Carbon (IV) sulphide
- NaH - Sodium hydride

Trinary compounds: These are compounds which contain 3 elements. e.g. the trinary acids. They are named as follows:

a. The name of the first element is written first
b. Count the number of last element (oxygen) and represent it with the latin prefix i.e one-mono, two-di, three-tri, four- tetra, five- penta, six-hexa, seven-hepta, eight-octa etc,
c. Change the oxygen to –oxo and attach it to the latin prefix e.g dioxo, tetraoxo
d. Change the name of the middle element to end in –ate and attach it to –oxo i.e sulphur-sulphate, manganese-manganate, chlorine-chlorate, carbon-carbonate
e. Determine the oxidation number of the middle element. Express the oxidation number as a roman figure in bracket and placed after the name ending in –ate.

Examples

1. $Ca(NO_3)_2$ -
2. $CuSO_4$ -
3. $Fe_2(SO_4)_3$ -
4. Na_2CO_3 -
5. $KClO_3$

Quaternary compounds : These are compounds that contain four elements

a. The names of the first two elements are written first
b. Count the number of last element (oxygen) and represent it with the latin prefix i.e one-mono, two-di, three-tri, four- tetra, five- penta, six-hexa, seven-hepta, eight-octa etc,
c. Change the oxygen to –oxo and attach it to the latin prefix e.g dioxo, tetraoxo
d. Change the name of the middle element to end in –ate and attach it to –oxo i.e sulphur-sulphate, manganese-manganate, chlorine-chlorate, carbon-carbonate
e. Determine the oxidation number of the middle element. Express the oxidation number as a roman figure in bracket and placed after the name ending in –ate

Examples:

1. $NaHSO_4$ -
2. NH_4HCO_3
3. KH_2PO_4

To name hydrated salts. The name ends in hydrate. The multiplying prefix is used depending on rthe number of molecules of water present in the compound

1. $Na_2SO_4 .10H_2O$

2. $CoCl_2.2H_2O$
3. $MgSO_4.7H_2O$
4. $ZnSO_4.7H_2O$

Students exercise: Name these compounds – Na_2SO_4, $KHCO_3$, $Cu(NO_3)_2$, $KClO_3$, $Na_2CO_3.10H_2O$

Naming acids

a. Count the number of last element (oxygen) and represent it with the latin prefix i.e one-mono, two-di, three-tri, four- tetra, five- penta, six-hexa, seven-hepta, eight-octa etc,
b. Change the oxygen to –oxo and attach it to the latin prefix e.g dioxo, tetraoxo
c. Change the name of the middle element to end in –ate and attach it to –oxo i.e sulphur-sulphate, manganese-manganate, chlorine-chlorate, carbon-carbonate
d. Determine the oxidation number of the middle element. Express the oxidation number as a roman figure in bracket and placed after the name ending in –ate
e. The name ends in acid (the hydrogen in front is not written first)

Examples:

1. $HClO_4$
2. H_2SO_4
3. H_2CO_3

Student's exercise: Name these acids; HNO_3, $H_2Cr_2O_7$, H_3PO_4

Naming ions and radicals

a. The first elements is not shown
b. Count the number of last element (oxygen) and represent it with the latin prefix i.e one-mono, two-di, three-tri, four- tetra, five- penta, six-hexa, seven-hepta, eight-octa etc,
c. Change the oxygen to –oxo and attach it to the latin prefix e.g dioxo, tetraoxo
d. Change the name of the middle element to end in –ate and attach it to –oxo i.e sulphur-sulphate, manganese-manganate, chlorine-chlorate, carbon-carbonate
e. Determine the oxidation number of the middle element. Express the oxidation number as a roman figure in bracket and placed after the name ending in –ate
f. The name ends in ion

1. SO_4^{2-}
2. CO_3^{2-}
3. HPO_4^{2-}

Name the ions: SO_3^{2-}, NO_3^-, ClO_3^-

RELATIVE ATOMIC MASS

The mass of as atom is very small to be used in calculations so early scientists expressed the mass of an atom as a ratio by comparing it with the mass of hydrogen (lightest element) and hence the term "relative". Recently, comparison was made with carbon-12 isotope. However, the mass of an atom can be accurately determined with the mass spectrometer introduced by Aston.

The relative atomic mass, A, of an element is the number of times the average mass of one atom of the element is heavier than one-twelfth the mass of one atom of carbon-12. **It is expressed as a ratio and has no unit.**

$$A = \frac{\text{Average mass of one atom of an element}}{{}^{1}/_{12} \text{ mass of 1 atom of carbon-12 isotope}}$$

The relative atomic mass of each element has been accurately verified with the mass spectrometer. The atomic mass unit is the amount of matter that has the mass of one-twelfth the mass of carbon-12. It is measured on a scale on which one carbon atom has a mass of 12 amu (12 units).

Carbon-12 scale: This is a scale that uses the carbon-12 isotope as a reference standard for measuring/comparing the atomic masses of elements

Atomic number	Element	Relative Atomic Mass, A_r
1	Hydrogen	1.008
2	Helium	4.0026
3	Lithium	6.939
4	Beryllium	9.012
5	Boron	10.81
6	Carbon	12.011
7	Nitrogen	14.0067
8	Oxygen	15.9994
9	Fluorine	18.9984
10	Neon	20.183
11	Sodium	22.9898
12	Magnesium	24.312
13	Aluminium	26.9812
14	Silicon	28.086
15	Phosphorus	30.9738
16	Sulphur	32.06
17	Chlorine	35.453
18	Argon	39.948
19	Potassium	39.102
20	Calcium	40.08

ATOMIC NUMBER AND MASS NUMBER OF AN ELEMENT

ATOMIC NUMBER: This is the number of protons present in the nucleus of an atom of an element. It is denoted by the letter Z. In a neutral atom, the number of protons is equal to the number of electrons. The atomic number of an element is a whole number. Atomic number of an element determines the nature of the atom and distinguishes it from atoms of other elements.

MASS NUMBER: This is the sum of protons and neutrons present in the nucleus of an atom of an element. It is denoted by letter A

(Mass number) A

(Atomic number) Z **X**

Examples: for Fluorine, F, Sodium, Na and Magnesium, Mg

1. $^{19}_{9}F$

2. $^{23}_{11}Na$

3. $^{24}_{12}Mg$

4. $^{14}_{7}N^{3-}$

5. $^{39}_{19}K^{+}$

Students exercise: Calculate the number of protons, electrons and neutrons present in these atoms

$^{40}_{18}Ar$

$^{40}_{20}Ca^{2+}$

$^{35}_{17}Cl^{-}$

$^{27}_{13}Al^{3+}$

$^{1}_{1}H$

ISOTOPY

Aston in 1920 discovered that atoms of some elements do not have exactly the same mass. Using the Aston mass spectrometer, he discovered that the difference is due to the number of neutrons. Isotopy is a **phenomenon** whereby atoms of an element have the same atomic number but different mass number. Isotopes are atoms of an element having the same atomic number but different mass number. Isotopes have same chemical properties but different physical properties because neutrons have no influence on the chemical properties of an atom but on the mass of the atom. It is electrons that determine the chemical properties of an atom. Examples include Hydrogen - $^{3}_{1}H$, $^{2}_{1}H$, $^{1}_{1}H$, Oxygen $^{16}_{8}O$, $^{17}_{8}O$, $^{18}_{8}O$, Chlorine – $^{35}_{17}Cl$, $^{37}_{17}Cl$, Boron – $^{10}_{5}B$, $^{11}_{5}B$, Carbon – $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$. The relative masses of isotopic elements are not whole number because their relative atomic mass is an average weight of the masses of the isotopes.

Examples:

1. Chlorine exists in two isotopic forms: $^{35}_{17}Cl$ and $^{37}_{17}Cl$. The relative abundance is 75% and 25% respectively. Calculate the relative atomic mass of Chlorine.
   Solution

Relative Atomic Mass = (% abundance x Mass of isotope 35) + ((% abundance Mass of isotope 37)

2. Two isotopes of Boron with mass number 10 and 11 exist in ratio 24:2. Calculate the Relative Atomic Mass of Boron.
   Solution

3. Given that the relative atomic mass of two isotopes of chorine is 35.5. determine the relative abundance of the two isotopes if the molar masses are 35 and 37 respectively

Student exercise: A sample of oxygen was found to contain three isotopes of masses 16, 17 and 18. The relative abundance are 99.76%, 0.04%, and 0.20% respectively. Calculate the relative atomic mass of oxygen.

RELATIVE MOLECULAR MASS M_r

This is the sum of the values of the relative atomic masses of the atoms of the elements present in the molecule. This includes relative atomic masses of ions, radicals and particles. It has no unit.

Examples:

Calculate the relative molecular mass of the following compounds.

a. $Fe_2(SO_4)_3$
b. Na_2CO_3
c. $(NH_4)_3PO_4$
d. $CaCl_2$ [Fe=56, S=32,O=16, Na=23, Ca=40, C=12, H=1, N=14, P=31]

Solution

a. $Fe_2(SO_4)_3$ has 2 atoms of Fe, 3 atoms of S and 12 atoms of oxygen
   Relative Molecular Mass = Relative atomic Masses of (Fe + S + O)

b. Na_2CO_3 has 2 atoms of Na, 1 atom of C and 3 atoms Of oxygen
   Relative Molecular Mass = Relative atomic Masses of (Na + C + O)

c. $(NH_4)_3PO_4$ has 3 atoms of N,12 atoms of H, 1 atom of P and 4 atoms of O
   Relative Molecular Mass = Relative atomic Masses of (N + H + P+O)

d. $CaCl_2$ has 1 atom of Ca and 2 atoms of Cl
   Relative Molecular Mass = Relative atomic Masses of (Ca + Cl)

Students exercise: Calculate the Relative Molecular Mass of these compounds (a). $Ca(OH)_2$ (b). $Al_2(SO_4)_3$ (c). AgCl (d). $H_2S_2O_7$

THE MOLE

The mole of a substance is the amount containing as many elementary units as the number of atoms in 12g of Carbon-12. It is the amount of a substance that contains same number of entities in exactly 12g of Carbon-12. These elementary entities can be atoms, molecules, ions, electrons etc. The mole is the **SI unit for measurement of mass**. Through experimental work, it has been established that one mole of any compound contains $6.02x10^{23}$ particles and this is known as the **AVOGADRO'S** NUMBER (N_A).

e.g one mole of Beryllium contains $6.02x10^{23}$ atoms of Be,
one mole of Helium contains $6.02x10^{23}$ atoms of He,
one mole of oxygen molecule contains $6.02x10^{23}$ atoms of O_2,
one mole of $CaCl_2$ contains one mole ($6.02x10^{23}$ ions) of Ca^{2+} and 2 moles (2x $6.02x10^{23}$ ions) of Cl^-
.

Number of moles = Number of Particles/ $6.02x10^{23}$

MOLAR MASS: This is the mass of one mole of a substance expressed in grams. The unit is g/mol

$$\text{Number of moles} = \frac{\text{Mass of a substance (in grams)}}{\text{molar mass of the substance (in g/mol)}}$$

CALCULATIONS ON MOLES

1. How many atoms of oxygen are there in 4.4g of carbon (IV) oxide?
   Solution
   1 mole of CO_2 contains 1 mole of carbon and 2 moles of oxygen

2. How many atoms are there in 6g of carbon? { I mole = $6.02x10^{23}$, C=12}
   Solution

3. What is the mass of 3 moles of oxygen, O_2?
   Solution

4. How many moles of gold ,Au, are present in 500g of Au_2O_3 {Au=197, O=16}
   Solution

5. How many moles are there in 30g of Calcium chloride, $CaCl_2$? {$CaCl_2$ =111}
   Solution

Students' exercise

1. How many moles are therein 20g of $KClO_3$ {K=39, Cl=35.5, O=16}
2. How many moles are there in 4g of NaOH {Na=23, O=16, H=1}
3. A sample of aluminium has a mass 5.4g. Calculate i. the number of moles ii. The number of atoms of Aluminium in the sample {Al=27}

PERCENTAGE COMPOSITION BY MASS OF ATOMS/ELEMENTS IN A COMPOUND

The formula of a compound shows the type and number of atoms present in a molecule of the compound. The formula mass can be calculated by adding the masses of the moles of the component elements. To calculate the percentage composition by mass of element in a a compound, we must

a. Write the correct formula of the compound
b. Calculate the molar mass(formula mass) of the compound
c. Divide the total mass of the atom/ element by the molar mass of the compound and multiply it by 100.

Examples:

1. Calculate the percentage by mass of all atoms in trioxonitrate(V) acid
   Solution
   Formular is HNO_3

2. Calculate the percentage by mass all the elements in calcium hydroxide, $Ca(OH)_2$
   Solution

3. Calculate the percentage by mass of all the component elements in Sodium tetraoxosulphate (VI) { Na=23, S=32, O=16)

Solution

Formular is Na_2SO_4

4. What is the percentage by mass of Iron (III) oxide {Fe=56, O=16}

Solution

Formular is Fe_2O_3

Student's exercise

Calculate the percentage by mass of the component elements in the following compounds

1. Calcium trioxonitrate (V)
2. Tetraoxosulphate (VI) acid
3. Water
4. Copper (II) tetraoxosulphate (VI)

EMPIRICAL FORMULAR AND MOLECULAR FORMULAR

Empirical formular is the simplest formular of a compound which shows the simplest ratio of the number of atoms in that compound. It shows the relative number of the different type of atoms present in a compound

Molecular formular is the formular which shows the actual number of atoms of different elements in one molecule of the compound.

In some cases empirical formular and molecular formular are the same but most times, it is a multiple of the molecular formular.

Relative molecular mass = (Empirical formular)n

Compound	**Empirical formular**	**Molecular formular**
Ethyne	CH	C_2H_2 n=2
Benzene	CH	C_6H_6 n=6
Hydrogen peroxide	HO	H_2O_2 n=2
Ethanedioic acid	CHO_2	$C_2H_2O_4$ n=2

DETERMINATION OF EMPIRICAL AND MOLECULAR FORMULAR

To determine the empirical and molecular formular of a compound,

1. Write out each element in the compound
2. Indicate under each element, the composition by mass
3. Divide by the molar mass of each element i.e mole ratio of each element.
4. Divide by the smallest value to obtain the ratio
5. Use the ratio to write the empirical formula
6. Solve for the value of n
7. Write the molecular formular with the value obtained

Examples

1. A compound was found by analysis to contain 64.45% carbon, 9.09% Hydrogen and 25.45% Nitrogen. Calculate its empirical formular? {C=12, H=1, N=14}

Solution

2. Find the empirical formular of a compound which on analysis gave the following as the reacting masses: Carbon= 2.0g, Hydrogen=0.34g, Oxygen=2.67g. From your results, find the molecular formular of the compound if the relative molecular mass is 60. (C=12, O=16, H=1)
   Solution

3. A sample of organic compound contains 0.624gof Carbon, 0.364g of Nitrogen and 0.208g of Oxygen (a) What is the empirical formula of the compound. (b). if the relative molecular mass of the compound is 194, what is its molecular formular (C=12, H=1, N=14, O=16) Caffeine

Solution

Students' exercise:

1. Calculate the empirical formular of an organic compound containing 81.8% Carbon and 18.2% Hydrogen.

2. A compound has the following percentage by weight C=40.0%, H=6.67%, O=53.3%. Calculate the empirical formula of the compound. If its molecular mass is 180, find its molecular formular. (C=12, H=1,O=16)
3. Determine the empirical formula of an oxide of Nitrogen containing 70% of oxygen. If the relative molecular mass of the oxide is 92, deduce its molecular formular.
4. An industrial raw material has the following composition by mass: iron=28.1%, chlorine=35.7%, water of crystallization= 36.2%. calculate the formula of the material l[H=1, O=16, Cl=35.5, Fe=56]

CHEMICAL EQUATIONS

Chemical equations are representations of a chemical reaction in form of symbols and formular of the elements and compounds involved. It is a short form of expressing the occurrence of a chemical reaction. The substances that participate in the reaction are called REACTANTS and are written on the left hand side. The new substances formed are called PRODUCTS and are written on the right hand side. Reactants and products are linked together by an arrow symbol pointing towards the direction of the products.

$$A + B \longrightarrow C + D$$

Reactants Products

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$

From the equation above,1 mole of magnesium reacts with 2 moles of HCl to give 1 mole of $MgCl_2$ and 1 mole of H_2

A chemical reaction must be balanced in order to comply with the law of conservation of matter.

To balance a chemical equation,

1. The formulae of the compounds involved must be correct. The formula must not be altered. Common gasses in the free state are diatomic e.g H_2, N_2, O_2, Cl_2, other elements are represented by their symbols e.g K, Na, Ca and radicals remain unchanged
2. The total number of atoms of each element on the Left hand side must be equal to the number of atoms of the same elements on the right hand side. This is achieved by multiplying the symbols and formula by appropriate whole numbers called co-efficient.
3. It is also necessary to indicate the state of the symbols of the substance involved in the equation: solid-s, liquid-l, aqueous-aq, and gas-g

There are three types of chemical reactions:

(i). **Word** equation : written using chemical name of the substances.

(ii). **Molecular** equation: written using the molecular formular

(iii). **Ionic** equation: written using formular of ions and it is deduced from the molecular equation.

Examples

Balance the following equations

1. $NH_{3(g)} + O_{2(g)} \longrightarrow NO_{(g)} + 6H_2O_{(g)}$
2. $Cu_{(s)} + HNO_{3(aq)} \longrightarrow Cu(NO_3)_{2(g)} + NO_{2(g)} + H_2O_{(l)}$
3. $Fe_{(s)} + O_{2(g)} \longrightarrow Fe_3O_{4(s)} + H_{2(l)}$

Solution

Students exercise

Write a balanced equation for the

(i) Reaction between Hydrogen and Oxygen to form steam.

(ii). Decomposition of calcium trioxocarbonate (IV) to yield calcium oxide and carbon (IV) oxide. What mass of calcium oxide is formed if 10.0g of calcium trioxocarbonate (IV) was decomposed?

IMPORTANCE OF CHEMICAL EQUATIONS

From chemical equations, we can determine

1. The stoichiometry of the reaction i.e the relationship between the mole ratio and the mass ratio of the reactants and the products
2. The reactants and products involved in the reaction.
3. The individual elements and the radicals involved in the reaction and their movement during the reaction.
4. The molar mass of each compound and their relative volumes (if gaseous)
5. The reactant in excess and by how much
6. The direction of the reaction whether reversible or irreversible reaction
7. The states of matter of the of each reactant and product

However, some information cannot be obtained from the chemical equation. They are

1. The colours of the reactants and products
2. The speed of the reaction
3. The heat changes that accompany the reaction

CALCULATIONS INVOLVING THE CHEMICAL EQUATIONS

1. Calculate the mass of sodium trioxocarbonate(IV) produced by the complete decomposition of 16.8g of sodium hydrogen trioxocarbonate (IV). (Na=23, C=12, O=16, H=1)
   Solution

2. (a) What volume of oxygen measured at stp will be produced on heating 24.5g of potassium trioxochlorate(V)
   (b) What is the mass of potassium chloride produced in the reaction (K=39, Cl=35.5, O=16)
   Solution
   $2KClO_{3(s)} \longrightarrow 2KCl_{(s)} + 3O_{2(g)}$

Students exercise

Chlorine is prepared by the oxidation of hydrochloric acid with Manganese (IV) oxide according to the equation:

$MnO_{2(s)} + 4HCl_{(aq)} \longrightarrow MnCl_{2(aq)} + Cl_{2(g)} + 2H_2O_{(l)}$

Calculate (a) the mass of hydrochloric acid required to react with 25.0g of Manganese (IV) Oxide (b). the mass of chlorine gas produced by the reaction?

LAWS OF CHEMICAL COMBINATION

When chemical changes take place, they do so according to some basic laws. There are four laws which descried the general features of a chemical change. They are:

1. **LAW OF CONSERVATION OF MASS (MATTER)**: This law states that matter is neither created nor destroyed during a chemical reaction but changes from one form to another. This is in support of Dalton's second atomic theory. In some reactions, matter may be lost in form of energy. Einstein derived an expression for the interconversion of matter and energy $E = mc^{2.}$ This law is also called the law of indestructibility of matter. The law was proposed by Antonie Lavoiser in 1789

EXPERIMENT TO VERIFY THE LAW OF CONSERVATION OF MASS

Reagents: Silver trioxonitrate (V) solution, sodium chloride solution.
Apparatus: conical flask with cork, test tube, string, weighing balance
Method: put some sodium chloride solution in a conical flask. Fill a small test tube with silver trioxonitrate (V) solution and by means of a string, suspend the test tube in the conical flask. Insert the stopper and place the apparatus on a balance. Record the mass of the whole system. Mix the two solutions by pulling the string attached to the bottom of the test tube.
Observation: When the two solutions mix, a white precipitate is formed indicating that a chemical reaction has taken place .
Result: The mass of the system is also recorded. The masses before and after the reaction was found to be the same.
Conclusion: Since there is no overall change in mass when the products were formed, we can infer that matter is neither created nor destroyed during a chemical reaction, hence the law of conservation of matter is verified
Note: Reactions involving gaseous substances are not suitable for this experiment due to pressure build up in the flask.

EXERCISE: Calculate the mass of silver chloride precipitate formed when 10.10g of silver trioxonitrate(V) reacts with 14.10g of sodium chloride to form 9.58g of sodium trioxonitrate (V)
Solution

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

Students exercise:
Calculate the mass of Barium chloride formed when 17.2g of Barium trioxonitrate (V) solution reacts with 14.8g sodium chloride solution to form 12.6g of sodium troxonitrate (V).

2. LAW OF DEFINITE PROPORTION/CONSTANT COMPOSITION

This law states that all pure samples of the same chemical compound contain the same elements combined in the same proportion by mass. According to Dalton's third atomic theory atoms of the same element are exactly alike and differ from atoms of other elements. For example, water can be obtained from different sources such as: sea, river, borehole, tap and well. It will be found that water from any of the sources contains only hydrogen and oxygen atoms in the same composition by mass. This law was proposed by Joseph Proust (1755 – 1826).

EXPERIMENT TO VERIFY THE LAW OF DEFINITE PROPORTIONS

This is done by analyzing samples of copper (II) oxide prepared by different methods; Methods A and B

Method A: Place some copper turnings in a crucible and add some concentrated trioxonitrate (V) acid, a little at a time so that all the copper dissolves completely. Evaporate the resulting green solution to dryness and continue heating until a black residue copper (II) is obtained. This is kept in a desiccator.

$$3Cu_{(s)} + 8HNO_{3(aq)} \longrightarrow 3Cu(NO_3)_{2(aq)} + 2NO_{2(g)} + 4H_2O_{(l)}$$

Method B: Place some copper (II) trioxocarbonate (IV) in a crucible and heat it strongly until it decomposes to give a black copper(II) oxide and carbon (IV) oxide. Keep the residue in a dessicator.

Procedure: Weigh two clean porcelain boats A and B and record their masses. Place some of the sample B in boat B. Weigh the boats in a inclined combustion tube and reduce the copper (II) oxide by passing a stream of hydrogen over the samples while healing it strongly. After some time, a reddish brown copper residue is left in each boat. Remove the flame but continue passing hydrogen as the copper cools down. This is to prevent the re-oxidation of hot copper residues by atmospheric oxygen. Any water formed during the reaction is removed by the fused Calcium chloride in the adjacent U-tube. Calculate the percentage by mass of copper in each sample

Result:

Sample	A	B
Mass of copper(II) oxide	3.55g	3.02g
Mass of copper residue	2.81g	2.42g
% copper present in copper (II) oxide	$\frac{2.81}{3.55} \times 100\%$	$\frac{2.42}{3.02} \times 100\%$
	79.2%	80.1%

The percentage of copper residue in the two samples is approximately 80% to the nearest tens irrespective of the method of preparation of the copper (II) oxide

Conclusion: In pure copper (II) oxide, copper and oxygen are always present in the ratio 4: 1 (80%: 20%) hence, verifying the law.

Students exercise

Two samples of Lead (II) sulphide were prepared by

(i) Heating a mixture of Lead and sulphur

(ii) Passing hydrogen sulphide gas into a solution of lead (II) chloride

Analysis of both samples (i) and (ii) shows that

i. 0.80g of Sulphur combined with 1.4g of Lead

ii. 5.50g of Lead (II) sulphide contain 2.0g of sulphur

Show that these results illustrate the Law of constant composition

3. **LAW OF MULTIPLE PROPORTION** states that if two elements A and B combine to form more than one chemical compound, then the various masses of one element A which combines separately with a fixed mass of the other element B are in a simple multiple ratio. The law was proposed by John Dalton in 1803.

Examples are

Copper and oxygen form black copper (II) oxide, CuO and red copper (I) oxide Cu_2O

Iron and oxygen form brown Iron (III) oxide, Fe_2O_3 and black Iron (II) oxde FeO

Iron and chlorine form brownish yellow Iron (III) chloride $FeCl_3$ and green Iron (II) chloride $FeCl_2$

EXPERIMENT TO VERIFY THE LAW OF MULTIPLE PROPORTIONS

Method: Weigh two metal boats and place some copper (I) oxide in one and copper (II) oxide in the other. Reweigh each boat and record their masses. Place the two boats in a hard glass combustion tube and reduce the oxides to copper by passing dry hydrogen gas over them while heating them. When the samples are cooled, weigh the copper residues obtained after the reduction and find the masses of two samples of copper

Result:

Samples	Copper (I) oxide Cu_2O	Copper (II) oxide CuO
Mass of sample (oxide)	3.04g	1.91g
Mass of copper residue	2.55g	1.38g
Mass of oxygen removed from oxide	0.49g	0.53g

Calculation:

Calculate the various masses of copper which would combine with a fixed mass of oxygen eg. 1g of oxygen

Copper (I) oxide Cu_2O **copper (II) oxide CuO**

Conclusion:
The masses of copper which combine separately with a fixed mass of oxygen in CuO and Cu_2O are in simple multiple number ratio of 1:2

Students exercise:
Copper forms two oxides by reacting with oxygen to form copper (I) oxide and copper (II) oxide. Analysis of the two oxides show that in copper (I) oxide, 11.00g of oxygen combine with 89.00g of copper while copper (II) oxide 20.00g of oxygen combine with 80.00g of copper. Show that the result of this reaction illustrates the law of multiple proportions.

4. **LAW OF RECIPROCAL PROPORTIONS** states that the several masses of several elements A, B, C which combine separately with a fixed mass of another element D are the same as (or simple multiples of) the masses in which A, B, C themselves combine with one another. The law was proposed by Jeremias Benjamin Richert in 1792

*(CO_2, SO_2, NO_2)

ATOMIC STRUCTURE
A series of experiments carried out led to the discovery of the fundamental sub-particles of the atom: - the electrons, the protons and the neutrons.
The characteristics of the particles are listed below

Sub-particle	Location	Relative charge	Symbol	Mass
Electron	Outside nucleus	-1	e^-	$^1/_{1840}$ amu
Proton	Nucleus	+1	P^+	1 amu
Neutron	Nucleus	0	n^o	1 amu

amu means atomic mass unit: This is the mass equivalent to thee mass of one-twelfth the mass of one atom of carbon-12 isotope

The following scientists contributed to the discovery of the structure of the atom

1. J.J Thompson discovered the electron and proton
2. R.A Millikan worked on charge to mass ratio using the oil drop experiment
3. Ernest Rutherford worked on the nuclear model of the atom
4. Henri Becquerel and Madam Curie – radioactivity which led to discovery of new elements
5. James Chadwick discovered the neutron
6. Neil Bohr- worked on the planetary model of the atom
7. Henry Moseley- assigned atomic number to elements
8. De Broglie & Erwin Schrodinger – wave mechanical model of the atom

DISCOVERY OF ELECTRONS

In 1897, J.J Thompson used cathode ray experiment to prove the existence of electrons. He applied potential difference of 5000V across a glass tube containing a gas at a very low pressure of 0.001 atm, the tube began to glow i. e the gas becomes conducting. When the potential difference was increased to 15,000V, a bright green glow appeared on the tube. He proved that the glow was due to some rays travelling in straight line from the cathode to the anode. He called the rays, cathode rays. He found out the following about the rays:

Characteristics of cathode rays

i. They are produced in all matter and therefore present in all matter
ii. Travel in straight in straight line at high velocity and cast shadow on opaque objects placed in their paths
iii. They are negatively charged particles and therefore deflected in an electric and magnetic field to the positive plate
iv. They are able to produce mechanical motion i.e they make paddle wheel placed in their path to rotate
v. They cause fluorescence of some objects e.g Zinc sulphide
vi. The specific charge (charge/mass) is $1.76 \times 10^{11} CKg^{-1}$

R.A Millikan in 1910 used the oil drop experiment to measure the charge to mass ratio of an electron to be $1.6 \times 10^{-19} C$ from the specific charge ratio, the mass of an electron is calculated to be $9.11 \times 10^{-31} Kg$.

Observations and deductions from the properties of the cathode ray experiment

Observation	**Deduction**
Rays are deflected by electric / magnetic fields	Rays consists of charged particles
Rays are deflected to the positive plate of an electric field	Rays are negatively charged
Rays rotate small wheel placed in their path	Particles have mass and possess momentum
Rays cast shadow of an object placed in their path	Rays are moving in straight line and have low penetrating power
Rays heat metal foil placed in their path	Rays conduct electricity

DISCOVERY OF PROTONS

J.J Thompson repeated the experiment but he used a perforated cathode. He noticed a reddish glow in the opposite direction of the green glow and he proved that this glow was due to positively charged rays. The rays have the following characteristics:

1. They travel in straight lines
2. They are positively charged
3. They need larger magnetic field to cause their deflection i.e they are heavier than the electrons
4. The mass of the positively charged particles depend on the nature of the gas in the tube.
5. The specific charge is $9.58x10^{7}CKg^{-1}$

R.A Millikan calculated the mass of a proton to be $1.67x10^{-27}Kg$

The proton was found to be 1840 times heavier than the electron.

In 1886, Eugene Goldstein discovered the evidence for the existence of a positively charged particle (he laid the ground work) but in 1911, Ernest Rutherford performed the gold foil experiment where he detected the presence of a heavy sub-particle in the center of the nucleus. He called it proton.

DISCOVERY OF NEUTRONS

James Chadwick in 1932 bombarded a thin sheet of Beryllium with alpha particles and discovered a new sub particle called neutron. The particles have the same mass with the proton but it has no charge. They are found to be contained in the nucleus of an atom.

GOLD FOIL EXPERIMENT OR ALPHA SCATTERING EXPERIMENT

Rutherford wanted to test the structure of the atom as proposed by J.J Thompson so, he directed a beam of alpha particles (positively charged helium ions) at a very thin plate of gold foil suspended in a vacuum. He expected that the alpha particles would just pass straight through the gold foil but surprisingly, he observed the following

Most of the alpha particles did pass straight through the foil

A small number of the particles were deflected through small angles or large angles

A very small number came straight back off the foil (rebound)

Conclusions from

1. The fact that most particles passed straight through is means that the atom is mostly an empty space
2. There is a dense, positively charged region at the center of the atom that he called nucleus. Deflection observed in few particles show that the nucleus is very small.

Gold was used because it is the only metal that could be rolled out to be very thin without cracking

The vacuum was important because any deflection of the particles would only be because of the collisions with the gold foil and not any other substance.

Limitations of Rutherford's alpha scattering experiment

1. His model could not explain the stability of the atom
2. The arrangement of electrons in a circular path was not defined
3. Any particle moving in a circular path would undergo acceleration and radiates energy (loose energy until it collapses)

Models of the atom

In 1898, .J Thompson proposed a model for the atom. He said that the atom is a sphere of positively charged matter in which negatively charged electrons are embedded.

In 1911, Rutherford proposed the nuclear theory of the atom. He said that the atom consists of a positively charged core called nucleus (where most of the mass of the atom is contained) and electrons move around the nucleus (planetary electrons).

BOHR'S MODEL OF THE ATOM

Neils Bohr in 1913 proposed a model for the atom based on the quantum theory. He said that electrons are restricted to certain energy levels and moves along circular orbits.

These energy levels are identified by means of the **Principal Quantum Numbe**r. He made the following assumptions:

i. Rutherford's model of the atom is correct
ii. Electrons can exist only in circular orbit of definite quantum energy
iii. An electron emits energy in the form of radiation when it moves from a higher to a lower permitted orbit. This produces a line atomic emission spectrum

The limitation of Bohr's model of the atom was that although it explained the spectra line of simple atoms like hydrogen but not for complex atoms

Bohr's model has now been replaced by the wave mechanics model which assumes that electrons show wave-like properties.

In 1885, Roetengen discovered x-rays (penetrating radiation) by colliding high energy electrons with the anode

In 1914, Moseley found out that he could assign a number (atomic number) based on the frequency of x-rays the element emits.

QUANTUM NUMBERS

In addition to the principal quantum number, wave mechanics uses three additional quantum numbers to describe an electron.

An **ORBITAL** is a region or space around the nucleus where there is a high possibility of finding an electron with a certain amount of energy. The lowest orbit with a minimum energy is known as the ground state. When an electron jumps to a higher orbit with higher energy level, we say the electron is in an excited state.

The four quantum numbers that define an electron within an atom are:

1. **Principal quantum number, n**, is the energy level or principal shells and has integral values of 1, 2,3, 4, …e.t.c. the maximum possible number of electrons in a shell is $2n^2$ where n is the total number of orbitals

   In the first shell, K n=1 maximum electrons is $2(1)^2 = 2$ electrons
   In the second shell, L n=2 maximum electrons is $2(2)^2 = 8$ electrons
   In the third shell, M n=3 maximum electrons is $2(3)^2 = 18$ electrons
   In the fourth shell, N n=4 maximum electrons is $2(4)^2 = 32$ electrons
   In the fifth shell, O n=5 maximum electrons is $2(5)^2 = 50$ electrons

2. **The subsidiary/ azimuthal quantum number, l**, shows the type of sub energy levels/ orbitalsof the atom and it has integral values of 0, 1, 2, 3 …(n-1). The electrons with subsidiary values of 0, 1, 2 and 3 are referred to as s, p, d and f electrons respectively which are initial letters of adjectives used to describe spectral lines - **s**harp, **p**rincipal, **d**iffuse and **f**undamental. l gives the shape of the orbital

3. **The magnetic quantum number, m,** has integral number -1, 0 …. +1 shows the number of spatial orientations or degeneracy of the orbitals in each subshell (sub-orbitals).
4. **The spin quantum number, s,** has values -½ and + ½ . (i.e only two directions of spin is possible-clockwise and anti-clockwise).

   An orbital can only hold two electrons which spin on its axis in opposite direction in an orbital.

In summary, n gives the relative distance of the electron from the nucleus, l gives the subshell and shape of the orbital of the electron, m shows the orientation of the orbital in space while s states the spin of the electrons.

Degenerate orbitals are orbitals which split in a magnetic field and possess the same energy despite different spatial orientations.

	L	**M**	**No. of orbitals**	**No. of maximum electrons**	**Name of orbital**
1	0 - 1s	0	1	2	s
2	0 - 2s	0	1	2	p
	1 - 2p	-1, 0, +1	3	6	
3	0 - 3s	0	0	2	d
	1 - 3p	-1, 0, +1	3	6	
	2 - 3d	-2, -1, 0, +1, +2	5	10	
4	0 - 4s	0	1	2	f
	1 - 4p	-1, 0, +1	3	6	
	2 - 4d	-2, -1, 0, +1, +2	5	10	
	3 - 4f	-3, -2, -1, 0, +1, +2, +3	7	14	

LAWS THAT GOVERN ELECTRON FILLING INTO ORBITALS

1. **Pauli's exclusion principle** states that no two electrons in the same orbital of an atom have same set of values for all the four quantum number(n, l, m and s). Two electrons may have the same values for n, l and m but s will definitely differ.

2. **Hund's rule of maximum multiplicity**: states that electrons occupy degenerate orbitals of sub-levels singly before pairing occurs in a e.g for carbon and oxygen

   Carbon $1s^2 2s^2 2p^2$ is like this and not like this

   Oxygen $1s^2 2s^2 2p^4$ is like this and not like this

3. **Aufbau's principle** states that in building up of atoms, electrons are filled in order of increasing energy levels. Electrons with lower energy (ground state) are filled first before that with higher energy.

   1s
   2s 2p
   3s 3p 3d
   4s 4p 4d 4f
   5s 5p 5d 5f
   6s 6p 6d 6f

SHAPES OF S AND P ORBITALS

The s orbital is spherical in shape. The orbital is spherically symmetrical and the electrons move in a spherical space around the nucleus to produce a spherical cloud

The p orbital is dumb bell in shape with a nodal plane passing through the nucleus. Each p-orbital can be orientated along the x, y and z axis which are at right angles to each other to give P_x ,P_y and P_z orbitals

ELECTRONIC CONFIGURATION AND THE PERIODIC TABLE

The electronic configuration is an arrangement of electrons in an atom in order of increasing energy levels.

Element	Symbol	Atomic number	Electronic configuration	Electronic configuration using spdf
Hydrogen	H	1	1	$1s^1$
Helium	He	2	2	$1s^2$

Lithium	L	3	2,1	$1s^2 2s^1$
Beryllium	Be	4	2,2	$1s^2 2s^2$
Boron	B	5	2,3	$1s^2 2s^2 2p^1$
Carbon	C	6	2,4	$1s^2 2s^2 2p^2$
Nitrogen	N	7	2,5	$1s^2 2s^2 2p^3$
Oxygen	O	8	2,6	$1s^2 2s^2 2p^4$
Fluorine	F	9	2,7	$1s^2 2s^2 2p^5$
Neon	Ne	10	2,8	$1s^2 2s^2 2p^6$
Sodium	Na	11	2,8,1	$1s^2 2s^2 2p^6 3s^1$
Magnesium	Mg	12	2,8,2	$1s^2 2s^2 2p^6 3s^2$
Aluminium	Al	13	2,8,3	$1s^2 2s^2 2p^6 3s^2 3p^1$
Silicon	Si	14	2,8,4	$1s^2 2s^2 2p^6 3s^2 3p^2$
Phosphorus	P	15	2,8,5	$1s^2 2s^2 2p^6 3s^2 3p^3$
Sulphur	S	16	2,8,6	$1s^2 2s^2 2p^6 3s^2 3p^4$
Chlorine	Cl	17	2,8,7	$1s^2 2s^2 2p^6 3s^2 3p^5$
Argon	Ar	18	2,8,2	$1s^2 2s^2 2p^6 3s^2 3p^6$
Potassium	K	19	2,8,8,1	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Calcium	Ca	20	2,8,8,2	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Scandium	Sc	21	2,8,9,2	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$
Zinc	Zn	30	2,8,18,2	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

Student's exercise

Write the electronic configuration of the following elements using the spdf notation

(i) Ni (ii) Iron { Ni= 28, Fe = 26}

The modern periodic table is an arrangement of elements into vertical columns called **groups** and horizontal rows called **periods**. This is based on the number of electrons in the outermost shell (valence electrons). It can be observed that there is a periodic reoccurrence of atoms of elements having the same numbers of electrons in the outermost shell.

	I	II	III	IV	V	VI	VII	0
1	$_1$H Hydrogen 1							$_2$He Helium 2
2	$_3$Li Lithium 2,1	$_4$Be Berylium 2,2	$_5$B Boron 2,3	$_6$C Carbon 2,4	$_7$N Nitrogen 2,5	$_8$O Oxygen 2,6	$_9$F Fluorine 2,7	$_{10}$Ne Neon 2,8
3	$_{11}$Na Sodium 2,8,1	$_{12}$Mg Magnesium 2,8,2	$_{13}$Al Aluminium 2,8,3	$_{14}$Si Silicon 2,8,4	$_{15}$P Phosphorus 2,8,5	$_{16}$S Sulphur 2,8,6	$_{17}$Cl Chlorine 2,8,7	$_{18}$Ar Argon 2,8,8
4	$_{19}$K Potassium 2,8,8,1	$_{20}$Ca Calcium 2,8,8,2						

The vertical columns are called **groups**. Elements in the same group have the same number of electrons in their outermost shell.

The horizontal rows are called **periods**. Elements in the same period have the same number of shells in their atoms.

GROUP 0 ELEMENTS

These elements have complete two electrons (duplet) or eight electrons (octet) in their outermost shell. They have stable configuration. They do not give out or accept electrons from other elements. They cannot react with other element. They are inert in nature and therefore called inert or rare gases

GROUPS I, II AND III ELEMENTS

Elements in groups I, II and III are called metals and have 1, 2 and 3 electrons in their outermost shells. They have the ability of losing the outer electrons in order to attain a stable configuration of the noble gas element preceding them. They become positively charged atoms called cations. They are electropositive elements.

e.g $Na - e^- \longrightarrow Na^+$
2,8,1 → 2,8
$11e^-, 11p^+$ → $10e^-, 11p^+$

$Mg - 2e^- \longrightarrow Mg^{2+}$
2,8,2 → 2,8
$12e^-, 12p^+$ → $10e^-, 12p^+$

$Al - 3e^- \longrightarrow Al^{3+}$
2,8,1 → 2,8
$13e^-, 13p^+$ → $10e^-, 13p^+$

GROUPS V, VI AND VII ELEMENTS

Elements in groups V, VI and VII are called non-metals and have 5, 6 and 7 electrons in their outermost shells. They have the ability of accepting 3, 2 and 1 electrons in order to attain a stable configuration of the noble gas element succeeding them. They become negatively charged atoms called anions. They are electronegative elements.

e.g $F + e^- \longrightarrow F^-$
2,7 → 2,8
$9e^-, 9p^+$ → $10e^-, 9p^+$

$O + 2e^- \longrightarrow O^{2-}$
2,6 → 2,8
$8e^-, 8p^+$ → $10e^-, 8p^+$

$N + 3e^- \longrightarrow N^{3-}$
2,5 → 2,8
$7e^-, 7p^+$ → $10e^-, 7p^+$

GROUP IV ELEMENTS

Elements in group IV have four electrons in their outermost shell. They can neither loose nor gain four electrons in order to attain octet configuration or become cation/anion. They are called metalloid.

PERIODICITY/ VARIATION OF PROPERTIES IN THE PERIODIC TABLE

Electropositivity decreases across the periodic table while electronegativity increases. If we consider the first twenty elements, Potassium is the most electropositive element while Fluorine is the most electronegative element

CHEMICAL COMBINATION/ BONDING

Atoms react or combine with one another in order to attain stable octet of electrons. These combinations are known as CHEMICAL BONDING. A chemical bond is a pair of electrons shared by two positively charged nuclei and attracted by both nuclei at the same time. There are two major categories of chemical bonds:

a. Intramolecular/interatomic bonds: these bonds join an atom with another atom to form molecules. These bonds determine the chemical properties of the substances. e.g ionic/electrovalent, pure covalent, co-ordinate covalent/dative and metallic bonds.
b. Intermolecular bonds are attractive forces between molecules or unit. These bonds determine the physical properties of substances like state of matter, boiling point, volatility, viscosity e.t.c. e.g van der waals forces (London dispersion forces, dipole-dipole forces, hydrogen bond and crystal lattice forces.

There are different types of combinations or chemical bonds. They are:

1. **ELECTROVALENT (IONIC) BOND**: This bonding involves the transfer of electrons from a metallic atom to a non-metallic atom (elements with large electronegativity difference >0.7). The metallic atom then becomes positively charged while the non-metallic atom becomes negatively charged i.e the atoms become oppositely charged ions. It is based on donor-acceptor principle in which there is complete transfer of electrons. Both ions have stable outer duplet or octet structure and are held together by **electrostatic forces of attraction between them**. E.g in sodium chloride sodium donates it outer electron to form Na^+ while chlorine accepts the atom to become negatively charged Cl^-

   Before combination After combination

Other examples are Calcium oxide, Calcium chloride, Magnesium oxide, Magnesium chloride

CHARACTERISTICS/PROPERTIES OF ELECTROVALENT COMPOUNDS

- They consist of aggregates of positive and negative ions arranged in an orderly pattern of three dimensional crystal lattices. They are crystalline in nature
- They are hard solids at room temperature and do not vaporize easily.
- They have high boiling point and melting point due to strong electrostatic bonds between the ions
- They dissolve readily in water and other polar solvents (like dissolves like)
- They do not dissolve in non polar solvents like toluene, benzene, ether, trichloromethane e.t.c
- They are good conductors of electricity molten or aqueous state (electrolytes) because the ions are free to move about but not in solid state since the ions are arranged into a crystal lattice and not free to about
- Reactions involving ionic compounds in aqueous forms are very fast

2. COVALENT BOND

This type of bonding involves sharing of electrons by two participating atoms in order to attain octet configuration. This pair of electrons forms an electron cloud which binds the atoms together and is known as the **shared pair (Covalent bond)**. Each of the electrons in the shared pair is contributed equally by each of the reacting atoms. Sharing of electrons occur between atoms of the same element or between atoms of very small electronegativities. A covalent bond is found in molecules and denoted by a horizontal bar in between the elements eg. H-H, O-O, HCl, N-N

In hydrogen

Other examples are CO_2, H_2O, O_2, CH_4, NH_3, C_2H_2

Covalent bonds can be polar or non-polar

Non-polar covalent bond is between homonuclear molecules like O_2, N_2, H_2, Cl_2 where there is equal sharing of electrons and there is no electrostatic charge between them

Polar covalent bond is between heteronuclear molecules like CO_2, CCl_4, SO_2, H_2O where there is unequal/assymetrical sharing of electrons resulting in the formation of dipole (electrostatic charge between them)

CHARACTERISTICS/PROPERTIES OF COVALENT COMPOUNDS

- They consist of molecules with a definite shape.
- They are gases or volatile liquids at room temperature and they vaporize easily.
- They have low boiling point and melting point due to weak intermolecular forces (hydrogen bonds and Van der waals forces) that bonds the molecules which can easily be broken by a small amount of heat energy
- They do not dissolve in water and other polar solvents
- They dissolve readily in non polar solvents like toluene, benzene, ether, carbon (IV) sulphide, trichloromethane e.t.c
- They do not conduct of electricity either in molten or aqueous state (non-electrolytes) because the molecules do not contain charged particles.

3. CO-ORDINATE COVALENT (DATIVE) BOND

This is a type of covalent bonding which combines the principle of both covalent and ionic bonding. The shared pair is contributed by only **one** of the participating atoms in order to attain octet stability. Such pair of electrons is called the **lone pair** of electrons. Lone pair of electrons are electron which did not participate in chemical bonding. The co-ordinate covalent bond is represented by an arrow pointing from the donor atom towards the acceptor atom

Others are hydrozonium ion H_3O^+, tetraamine copper (II) ion $[Cu(NH_3)_4]^{2+}$, hydrated copper (II) ion $[Cu(H_2O)_4]^{2+}$, phosphorus oxochlorate (III) $POCl_3$, hexacyanoferrate(II) ion $[Fe(CN)_6]^{2+}$
Co-ordinate covalent compounds are similar in properties to covalent compounds. They are both non-electrolytes but a dative bond makes a compound less volatile.

Students exercise:

a. In a tabular form, write 4 differences between electrovalent and covalent compounds
b. Name the bonds present in NH_4Cl

4. **METALLIC BOND:** This is a bond that holds atoms of metals together. Each metallic atom contributes its valence electron to the sea of mobile electron cloud, thus becoming positively charged. The mobile electrons exert an attractive force on the positive ions which holds the atoms together in a crystal lattice. This force of attraction is known as metallic bonds. The larger the number the number of electrons, the stronger the bonds. It is strong in some metals like Iron but weak in some like Potassium and Sodium. When a force is applied on a metal, the layers of metallic ion can slide over one another without shattering the crystal lattice (this accounts for malleability and ductility of metals) because there are no rigid directed bonds in a metal. The bonding agent is a moving electron cloud. The metallic bond cloud is responsible for some of the physical properties of metals like ductility, malleability, electrical and thermal conductivity, high melting point and boiling points.

Other types of bonding are:

1. **HYDROGEN BONDNG:** This is a dipole-dipole attraction formed when hydrogen is covalently linked with strongly electronegative elements like Nitrogen, Fluorine, or Oxygen. The electronegative elements have strong affinity for electrons and so they attract the shared pair to themselves and this result in the formation of a dipole leaving a partial positive charge on the hydrogen and a partial negative charge on the electronegative atom. The electrostatic attraction between the two dipoles is known as hydrogen bond. Although this bond is weak, it has important effects on the physical properties of compounds like hydrogen fluoride and water (it is responsible for the high boiling and melting point). It is also responsible for the solubility of alkanols and alkanoic acids. The strongest of the bonds is in hydrogen fluoride. Other hydrogen bonds break under the influence of heat but not hydrogen fluoride.

2. **VAN DER WAALS FORCES**: These are weak, short range attractive forces between uncharged atoms or molecules. They are intermolecular forces which arise from fluctuating dipoles in atom and molecules brought about by movement of electrons around the atomic nucleus. They are much weaker than ionic and covalent bons. They are important in liquefaction of gases and sublimation of iodine and naphthalene crystals. These forces increases as the number of electrons increases i.e they are stronger in graphite and iodine (solid) than bromine (liquid) and less in chlorine (gas). These forces were first described by J.D Van der waals.
3. **DIPOLE-DIPLOLE ATTRACTION**: these attraction occur between permanently polar molecules. When electrons are unevenly shared between atoms of different elements, the distribution of electron density around the atoms will be such that the more electronegative atom become electron rich (δ^-) (nucleophile) while the other becomes electron poor (δ^+)(electrophile). The separation of charges in a polar covalent bond is a dipole. Dipole-diplole (dipolar) forces operate between polar covalent molecules as a result of the weak electrostatic force of attraction between the positive end and the negative end of a neighbouring dipole in polar molecules such as HF, HCl, H_2O, NH_3 and H_2S at low temperatures

Student exercise: define electrophiles and nucleophiles. Give 4 examples of each

INTRODUCTION TO ANALYTICAL CHEMISTRY
VOLUMETRIC ANALYSIS

Chemists all over the world analyze substances to determine their composition, either quantitatively or qualitatively.

In **qualitative** analysis, substances are analyzed to identify the elements or ions present in a given sample but in **quantitative** analysis, the quantity or amount of the elements or the compound present in a given sample is determined. There are two methods of quantitative analysis: **Volumetric** and **Gravimetric**. Volumetric analysis is based on the volume measurements of solutions while the Gravimetric analysis is based on direct mass measurements of substances.

TERMS USED IN VOLUMETRIC ANALYSIS

1. **Titration**: This is a method in volumetric analysis where a solution from a graduated vessel (burette) is added to a known volume of a second solution until the chemical reaction between them is just completed. This is shown by a colour change of the indicator.
2. **Standard Solution**: This is a solution whose concentration is accurately known and used to react with a solution of an unknown concentration.
3. **Molar Solution**: This is a solution that contains one mole or one molar mass of a compound in one dm^3 of solution.
4. **Molar Concentration**: This is the number of moles of solute in one dm^3 or 1000 cm^3 of a solution. It is expressed in mole per dm^3. The unit is mol/ dm^3 or $moldm^{-3}$.
5. **Mass Concentration**: This is the amount of solute in grams present in one dm^3 or 1000 cm^3 of a solution. It is expressed in grams per dm^3. The unit is g/dm^3 or gdm^{-3}.
6. **Standardization**: This is determining the unknown concentration of a given solution by titrating it with another solution of known molar concentration.
7. **Standard**: these are materials containing a precisely known concentration of a substance for use in quantitative analysis.
8. **Titration Mixture**: This is the solution produced in the conical flask after each titration containing salt and water only.
9. **Acid-Base Indicator:** This is a dye which changes colour according to the pH of solution which is used to determine end point of Acid-Base titrations.
10. **End point**: This is a point in which all the Base in a particular titration volume has been neutralized by the Acid. This is accompanied by a sharp colour change of the indicator.
11. **Equivalence point:** This is the precise moment when the added titrant is stoichiometrically equivalent to the analyte.

A standard solution is a solution with an accurately known concentration.

A primary solution is a solution made from primary standard. It contains a definite number of moles of substance.

A primary standard is a reagent which is pure, weighed easily and a true representative of the number of moles of the substance contained. A primary standard has high purity, high stability, non-toxicity and low hygroscopy.

FORMULAS EMPLOYED IN QUANTITATIVE CALCULATIONS

1. Molar Concentration of solution (mol/dm^3) = $\frac{\text{Mass Concentration } (g/dm^3)}{\text{Molar mass of solution } (g/mol)}$

2. Mass Concentration of solution (g/dm^3) = $\frac{\text{Mass} \times 1000}{\text{Volume}}$

3. Number of moles = Molar Concentration of solution (mol/dm^3) x $\frac{\text{Volume}}{1000cm^3}$

   $n = CV$

4. Number of specified entities = Molar Concentration (mol/dm^3) x Avogadro's Number

5. $\frac{\text{Concentration of Acid} \times \text{Volume of Acid}}{\text{Concentration of Base} \times \text{Volume of Base}} = \frac{\text{Number of moles of Acid}}{\text{Number of moles of Base}}$

   $\frac{C_AV_A}{C_BV_B} = \frac{n_A}{n_B}$

6. Concentration of solution1 x Volume of solution1 = Concentration of solution2 x Volume of solution2

   $C_1V_1 = C_2V_2$

   Dilution factor = $\frac{Cs}{Cd}$ or $\frac{Vd}{Vs}$

MATERIALS USED IN ACID-BASE TITRATION

1. Weighing Bottle
2. Chemical Balance
3. Pipette
4. Burette
5. Retort stand with clamp
6. Filter paper
7. Funnel
8. White tile
9. Standard Volumetric flask
10. Conical flask

PRECAUTIONS IN QUANTITATIVE ANALYSIS

PIPETTE

1. Rinse the pipette with the distilled water and then with the solution it should be used to measure i.e the base.
2. Avoid air bubbles in the pipette.
3. Take readings at eye level.
4. Do not blow the last drop on the pipette.
5. Do not use greasy pipette for titrations

BURETTE

1. Rinse the burette with acid or allow it to drain after rinsing it with distilled water.
2. Make sure that the bubble jet is filled.
3. Make sure that the burette is not leaking.
4. Take readings at eye level at the lower meniscus to avoid error due to parallax.
5. Remove the funnel before taking your readings.

6. Avoid inconsistent burette readings
7. Read the lower meniscus on the burette.
8. Do not use greasy burette for titrations.

CONICAL FLASK

1. Rinse the conical flask with distilled water only and allow it to drain
2. Avoid acid overflowing into the sides of the conical flask from the burette.

RECORDING AND AVERAGING

1. Subtraction errors
2. Make sure you do not omit the unit of volume of solution.
3. Do not average values that are not concordant.

SHARP END POINTS

1. Do not use more than 3 drops of the indicator.
2. Shake the conical flask thoroughly during titrations.
3. Use a white tile.
4. Add acid drop by drop just before end point
5. Avoid running in the acid at a very fast rate.

PRACTICAL TEST I

1. You are provided with samples A, B, C and D. Carry out the following exercises on them and record your observations

Test	Observation	Inference
Sample A+ distilled water + litmus		
Sample B + distilled water + litmus		
Portion of C + phenolphthalein indicator		
Portion of C + methyl orange indicator		
Portion of D + phenolphthalein indicator		
Portion of D + methyl orange indicator		

2. Name the laboratory apparatus used in the lab to:

i. Convert liquid to vapour during distillation
ii. Determine the volumetric composition of water
iii. Produce an intermittent supply of any gas which can be evolved by the action of a liquid on a solid without heating
iv. Separate a mixture of water and benzene
v. Obtain pure water from brine

3. a. Draw a labeled sketch to illustrate the
   i. Separation of a mixture by sublimation
   ii. Separation of a mixture of kerosene and water.
   b. List three apparatus needed for the evaporation of Sodium chloride to dryness
   c. Explain how you would purify a mixture of sodium chloride contaminated by ammonium chloride.
4. a. What is a primary solution?
   b. Write four precautions taken during titration
   c. Write three sources of error during titration.

PRACTICAL TEST II

QUANTITATIVE ANALYSIS

1. A is 0.05 moldm^{-3} of acid H_yX. B is a solution of NaOH containing 0.025 moles per 250cm^3 solution.

a. Put A in burette and titrate it against 20.00 or 25.00cm^3 portions of B using methyl orange as indicator. Tabulate your readings and calculate the average volume of acid used.

b. From your results and information provided above, calculate
   i. Amount of acid in the average titre.
   ii. Amount of base in 20.00 or 25.00cm^3
   iii. Mole ratio of acid to base

c. Write a balanced equation for the reaction between the H_yX acid and the base NaOH

d. State the basicity of the acid

PRACTICAL TEST III (ALTERNATIVE TO PRACTICAL)

1. The following table shows the burette reading obtained during the titration of 26.00cm^3 of a solution of Sodium trioxocarbonate (IV) containing 1.3g in 250cm^3 of solution was titrated against a solution of Hydrochloric acid of unknown concentration using phenolphthalein as an indicator

Burette reading (cm^3)	Rough	1st	2nd
Final reading (cm^3)	23.50	44.20	25.20
Initial reading (cm^3)	2.00	23.50	4.30
Volume of Acid used(cm_3)			

   i. Copy and complete the table
   ii. Calculate the average volume of acid used
   iii. Write a balanced equation for the reaction
   iv. Calculate the concentration of the base in moldm^{-3}
   v. Calculate the concentration of the acid in moldm^{-3} and in g dm^{-3}
   (Na=23, C=12, O=16, H=1, Cl=35.5)

2. Consider the table

Solution	**R**	**S**	**T**
pH	12	7	3

Which of the solutions
   i. have a soapy feel?
   ii. Turns yellow on addition of methyl orange
   iii. Reacts with Magnesium to liberate hydrogen gas
   iv. Could be aqueous solution of Sodium tetraoxosulphate (VI)

3. i. Name the substance that would be used in the laboratory (i) to dry ammonia gas (ii) as the mobile phase for the separation of chlorophyll by paper chromatography
   ii. If you were provided with anhydrous Sodium trioxocarbonate (IV) salt, stirrer and spatula, List three other materials you would need to prepare standard Sodium trioxocarbonate (IV) solution.

iii. Describe how you would prepare a 250cm^3 0.1 molar standard Sodium trioxocarbonate (IV) solution.

PREPARATION OF STANDARD SOLUTION

A standard solution is a solution of known concentration.

A molar solution is a solution that contains one mole of a substance in $1dm^3$ of solution.

To prepare a standard solution, you need

1. A chemical balance
2. Beaker
3. Funnel
4. Distilled water
5. Standard volumetric flask

Procedure

1. Weigh the amount of solid required with the chemical balance using the weighing bottle
2. Transfer the solid to the beaker
3. Dissolve the weighed mass with distilled water in a beaker
4. Transfer the solution in the beaker into the required standard volumetric flask with a funnel and rinse the beaker with more distilled water and add to the flask
5. Add more distilled water to about half of the volumetric flask, cork and shake to dissolve the solid completely
6. Add more distilled water to the graduation mark on the volumetric flask and cork again, hold tightly and turn upside down to mix well

For example,

To make $1dm^3$ 1M NaOH solution, weigh 40g and dissolve in $1000cm^3$ of water

To make $1dm^3$ 0.5M NaOH solution, weigh 20g and dissolve in $1000cm^3$ of water

To make $1dm^3$ 0.1M NaOH solution, weigh 4g and dissolve in $1000cm^3$ of water

To make $1dm^3$ 0.05M NaOH solution, weigh 2g and dissolve in $1000cm^3$ of water

To make $1dm^3$ 2M NaOH solution, weigh 80g and dissolve in $1000cm^3$ of water

Preparation of dilute acid solution from concentrated stock acid solution(manufacturer's label)

The formula for the dilution is

$$C_o = \frac{10pd}{M}$$

where, V_o = volume of conc. Stock solution
M = molarity of acid
P = percentage purity
d = density of concentrated solution

DILUTION LAW

Dilution of any given standard solution to a required concentration is carried out by adding water to the given volume to obtain a solution of a lesser concentration.

The formular for the dilution law is stated as $C_1V_1 = C_2V_2$

Example 1

Calculate the volume of 2M H_2SO_4 needed out of $100cm^3$ stock solution to prepare $0.05moldm^{-3}$ $100cm^3$ dilute H_2SO_4?

Solution

Example 2

Calculate the volume of $2.5moldm^{-3}$ stock HCl required to prepare $500cm^3$ of $0.20moldm^{-3}$ dilute HCl?

Solution

DILUTION FACTOR

The formular is Dilution factor = $\frac{Cs}{Cd}$ or $\frac{Vd}{Vs}$

QUALITATIVE ANALYSIS

Qualitative analysis deals with the chemical identification of the component ions present in a sample of a given substance. Positively charged particles are called cations, metallic ions or basic radicals e.g Cu^{2+}, Zn^{2+}, Al^{3+},Fe^{2+}, Fe^{3+},Pb^{2+}, NH_4^+

Negatively charged ions are called anions or acid radicals e.g NO_3^-, HCO_3^{2-}, SO_3^{2-}, SO_4^{2-}, Cl^-, S^{2-}, HCO_3^-.

The preliminary (physical) tests carried out on a sample include:

1. Solubility test
2. Appearance test (colour and texture)
3. Odour test
4. Action of heat

Solubility test: A substance which when in water dissolves completely is said to be a soluble substance while a substance which does not dissolve when water is added is said to be an insoluble salt.

Rules for solubility of a substance in water

1. All trioxonitrate(V) salts are soluble in water
2. All chloride salts are soluble in water except Pb, Ag, Hg and Cu
3. All trioxocarbonates(IV) are insoluble except Na, K and NH_4^+
4. All sulphides are insoluble except Na, K and NH_4^+
5. All tetraoxosulphates(VI) are soluble except Ba, Pb, Ag and Hg

Appearance/colour test

Colour test helps to suspect a substance under test by its colour. For example, a colourless substance can indicate any salt of the following ions: Zn^{2+}**,** Al^{3+}**,** Pb^{2+}, NH_4^+, Ca^{2+}, Mg^{2+}, K^+**,** and Na^+ since their compounds are mainly colourless.

Similarly, a coloured substance can make the following to be suspected about a substance under test.

Colour	**Probable salt/compound**
Green	$FeSO_4$ (green crystal), $CuCO_3$(green powder), $CuCl_2$ (green crystal)
Blue	Copper II salts e.g $CuSO_4$, $Cu(NO_3)_2$
Yellow/brown	Iron III salts, PbO
Black	CuO, FeS, PbS, CuS

Note: Relating solubility and colour test together is very important in qualitative analysis. That is, as a sample dissolves in water, the colour of the solution formed must be stated.

Odour test: Odour test is more prominent in gases. Some gases have odour while some are odourless. Often, the colour and odour of a gas go together in suspecting a gas evolved in a chemical reaction.

Gas	**Odour**	**Chemical test for the identification of the gas**
Steam, $H_2O_{(g)}$	A colourless, odourless with a neutral steaming appearance. The gas condenses at the cooler surface of the boiling tube	Turns white anhydrous $CuSO_4$ blue Turns blue $CoCl_2$ paper pink
Oxygen, O_2	A colourless, odourless gas, neutral to litmus	Rekindles a glowing splint
Hydrogen, H_2	A colourless, odourless gas, neutral to litmus	Gas makes a pop sound with a lighted paper
Carbon(IV) oxide, CO_2	A colourless, odourless gas, turns moist blue litmus paper red	Gas turns lime water milky. The milkiness disappears when the gas is passed in excess
Hydrogen chloride, HCl	A colourless gas with a pungent/irritating smell, turns moist blue litmus paper red	Gas gives dense white fumes with ammonia gas. Gas gives white precipitate with $AgNO_3$ in HNO_3 acid

Chlorine gas, Cl_2	A greenish-yellow gas with an irritating smell, turns moist blue litmus paper red and then bleaches it	Gas gives white precipitate with $AgNO_3$ in HNO_3 acid
Ammonia gas, NH_3	A colourless gas with a choking smell, turns moist red litmus paper blue (the only alkaline gas)	Gas gives dense white fumes with conc. HCl
Bromine gas, Br_2	A reddish-brown gas with a pungent smell, turns moist blue litmus paper red and then bleaches it. Gas condenses at the cooler part as a brown liquid	Gas gives pale yellow precipitate with $AgNO_3$ in HNO_3 acid
Hydrogen sulphide, H_2S	A colourless gas with a rotten egg smell, turns moist blue litmus paper red	Gas turns moist lead(II) ethanoate paper black Gas gives a black precipitate with $Pb(NO_3)_2$ Gas decolorizes purple $KMnO_4$ solution with a deposit of sulphur
Nitrogen, N_2	A colourless, odourless gas, neutral to litmus	No specific test
Nitrogen (IV) oxide, NO_2	A reddish-brown gas with an iritating smell, turns moist blue litmus paper red but does bleach it.	Gas gives no precipitate with $AgNO_3$ in HNO_3 acid
Sulphur (IV) oxide, SO_2	A colourless gas with a choking smell, turns moist blue litmus paper red	Gas decolorizes purple $KMnO_4$ solution without a deposit of sulphur Gas decolorizes orange $K_2Cr_2O_7$ solution without a deposit of sulphur
Methane	A colourless, odourless gas, neutral to litmus	Gas fails to decolorize purple $KMnO_4$ solution

Action of heat on specimen: On heating, a substance may get decomposed and sometimes there may be a colour change accompanying such reaction. In reporting heating test, the report should be made in respect of colour change of the substance heated and gas evolved, if any that accompanies the reaction.

CHEMICAL TEST FOR IDENTIFICATION OF IONS IN SOLUTION

A chemical test involves the use of chemical reagent in solution. The action or reaction of the solution of a chemical reagent with the solution of a substance under test usually gives a specific precipitate with specific colour that helps in identifying the ion present. A ***precipitate*** is a solid formed due to a chemical reaction taking place between two solutions.

Note: A chemical reagent used in chemical identification test must be in solution and also the substance under test must be in solution.

Chemical Identification of a Cation in a Solution of a Given Substance

The two common chemical reagents that are required for cation or metallic ion identification are:

1. Sodium hydroxide solution, $NaOH_{(aq)}$
2. Aqueous ammonia solution, also known as dilute ammonium hydroxide, $NH_4OH_{(aq)}$

The quantity of solution of sample is about $2cm^3$. The use of a larger volume of sample may lead to wrong observation even when the right test is done.

What is Cation Testing?

Testing for cations is a test used in chemistry to identify metal or metal ions (cations) found in compounds. There are two types of tests used in chemistry to test for cations.

1. Flame Test.

The Flame test involves exposing the compound to a flame and identifying the compound by the flame color produced. When the compound is heated, the electrons move to energy levels that are higher. This movement makes the ions energetically unstable and they move back into the previous energy join. As they move back, the ions release light energy. on the type of ions, this light energy is different and produces differing colors of flame.

Procedure for Flame Test

A flame test can be used to identify a compound using a flame. The procedure is as follows:

1. Use a clean flame test wire.

2. Heat the Nichrome wire.

3. Dip it in the hydrochloric acid.

4. Dip the wire into the compound you are testing so a small blob is collected on the wire.

5. Put the tip of the wire into a flame and see what color it turns.

These are the colors you will see for different ions:

Ion	**Color**
Sodium (Na^+)	Golden-yellow
Potassium (K^+)	Lilac
Calcium (Ca^{2+})	Brick-red
Copper (Cu^{2+})	Green

2. Sodium Hydroxide Test for Cations

This test uses sodium hydroxide to test and identify metal ions by the precipitation formed. Sodium Hydroxide is added to the solution being tested and the color of precipitation formed allows for identification of the compound.

Add several drops of sodium hydroxide (NaOH) solution to the solution being tested. If a colored precipitate is formed then stop and find out what the cation is. If a precipitate forms then continue to add NaOH to it and observe whether the precipitate dissolves.

Chemical Behaviour of Cations in Sodium Hydroxide Solution, $NaOH_{(aq)}$

Test	**Observation**	**Inference**
$2cm^3$ of solution of sample + $NaOH_{(aq)}$ in drops + in excess	White gelatinous precipitate in drops, soluble in excess White chalky precipitate soluble in excess White chalky precipitate insoluble in excess	Al^{3+} or Zn^{2+} may be present Pb^{2+} present Ca^{2+} present
	Pale blue gelatinous precipitate in drops, insoluble in excess	Cu^{2+}
	Dirty green gelatinous precipitate in drops, insoluble in excess	Fe^{2+}
	Reddish brown gelatinous precipitate in drops, insoluble in excess	Fe^{3+}
+ heat	No visible reaction A gas with a choking smell like that of urine is given off which turns red litmus paper blue	Gas is NH_3 from NH_4^+

3. Aqueous ammonia Test for Cations

This test uses aqueous ammonia to test and identify metal ions by the precipitation formed. Aqueous Ammonia is added to the solution being tested and the color of precipitation formed allows for identification of the compound.

Add several drops of aqueous ammonia ($NH_3.H_2O$) solution to the solution being tested. If a colored precipitate is formed then stop and find out what the cation is. If a precipitate forms then continue to add the aqueous ammonia to it and observe whether the precipitate dissolves.

Chemical Behaviour of Cations in aqueous ammonia Solution, $NH_4OH_{(aq)}$

Test	**Observation**	**Inference**
$2cm^3$ of solution of sample + $NH_{3(aq)}$ in drops + in excess	White gelatinous precipitate in drops, soluble in excess to form a colourless solution precipitate is insoluble in excess	Al^{3+} or Zn^{2+} Zn^{2+} present Al^{3+} present

	White chalky precipitate in drops, precipitate is insoluble in excess	Pb^{2+}
	Pale blue gelatinous precipitate in drops, soluble in excess to form a deep blue solution	Cu^{2+}
	Dirty green gelatinous precipitate in drops, insoluble in excess	Fe^{2+}
	Reddish brown gelatinous precipitate in drops, insoluble in excess	Fe^{3+}
	No visible reaction	Ca^{2+}or NH_4^+

Confirmatory test for cations

To an aqueous solution of the sample,

S/No	Test	Observation	Inference
1.	**Zn^{2+}** Add $K_4Fe(CN)_6$ potassium hexacyanoferrate (II) **OR** Add NaOH + H_2S gas	White precipitate White precipitate	Zn^{2+} ions confirmed Zn^{2+} ions confirmed
2.	**Ca^{2+}** Add $NH_{3(aq)}$ + $(NH_4)_2CO_3$ solution	White precipitate insoluble in excess of reagent	Ca^{2+} ions confirmed
3.	**Pb^{2+}** Add drops of dil. HCl **OR** Add KI or K_2CrO_4 solution	White precipitate soluble when warmed and reappears on cooling Yellow precipitate	Pb^{2+} ions confirmed
4.	**Al^{3+}** Add $NH_{3(aq)}$ + NH_4Cl solution **OR** Add KI solution	White precipitate formed No visible reaction/ colorless solution	Al^{3+} ions confirmed Al^{3+}ions confirmed
5.	**Cu^{2+}** Add excess ammonia solution OR Add $K_4Fe(CN)_6$ potassium hexacyanoferrate (II) OR	A deep blue solution Brown precipitate	Cu^{2+} ions confirmed Cu^{2+} ions confirmed

	Add KI solution	Brown solution and dirty white precipitate	Cu^{2+} ions confirmed
6.	**Fe^{2+}** Add $K_4Fe(CN)_6$ potassium hexacyanoferrate (II) **OR** Add $K_3Fe(CN)_6$ potassium hexacyanoferrate (III) **OR** Add a few drops of conc. H_2SO_4 and boil	Light blue precipitate Deep blue precipitate Green solution turns brown due to oxidation of Fe^{2+} to Fe^{3+}	Fe^{2+} ions confirmed Fe^{2+} ions confirmed Fe^{2+} ions confirmed
7.	**Fe^{3+}** Add $K_4Fe(CN)_6$ potassium hexacyanoferrate (II) **OR** Add $K_3Fe(CN)_6$ potassium hexacyanoferrate (III) **OR** Add KCN potassium thiocyanate	Deep blue precipitate Brown solution Blood-red coloration	Fe^{3+} ions confirmed Fe^{3+} ions confirmed Fe^{3+} ions confirmed
8.	**NH_4^+** Add Add $NH_{3(aq)}$ + warm	Colourless, odourless gas with a choking smell. The gas turns red litmus paper blue	NH_4^+ ions confirmed

ANION IDENTIFICATION

Generally, anions are identified by the gas evolved when an acid is added to the sample under test. The anions for chemical identification are: Cl^-, SO_4^{2-}, SO_3^{2-}, CO_3^{2-}, HCO_3^-, NO_3^-

Test	**Observation**	**Inference**
Solution of sample + dilute $HNO_{3(aq)}$ + silver trioxonitrate (V) + aqueous ammonia in excess	No visible reaction White precipitate formed Precipitate dissolves to form a colourless solution	 Cl^- present Cl^- confirmed
	Pale yellow precipitate formed Precipitate dissolves partially	Br^- present Br^- confirmed

	yellow precipitate formed Precipitate remains insoluble	I^- present I^- confirmed

Note: The reason for adding dilute HNO_3 acid first when testing for chloride ion (halide ions) in solution is to dissolve and prevent heavy cation that will interfere with the reaction since aqueous ammonia solution is to be added later.

Test	**Observation**	**Inference**
SO_4^{2-} Solution of sample + $BaCl_{2(aq)}/Ba(NO_3)_{2(aq)}$ + excess dilute $HCl/HNO_{3(aq)}$	White precipitate formed Precipitate dissolves Precipitate insoluble	SO_4^{2-}, SO_3^{2-}, CO_3^{2-} or S^{2-} SO_3^{2-}or CO_3^{2-} SO_4^{2-} confirmed
SO_3^{2-} Solution of sample + few drops of dil. H_2SO_4 + $KMnO_4$ solution	Purple colour turns colourless	SO_3^{2-} confirmed
CO_3^{2-} Solution of sample + dilute $HCl/HNO_{3(aq)}$ + lime water	Effervescence occurs and a gas which turns blue litmus paper red is evolved Gas turns lime water milky	CO_3^{2-} confirmed
HCO_3^{2-} Solution of sample + $BaCl_2/MgSO_{4(aq)}$ + boil **OR** Solution of sample + few drops of phenolphthalein solution + boil	No visible reaction White precipitate formed No visible reaction Solution turns pink	HCO_3^{2-} confirmed HCO_3^{2-} confirmed
NO_3^- Solution of sample + freshly prepared $FeSO_4$ solution + conc. H_2SO_4 down the slanted side of test tube	Brown ring is formed at the junction of the two solutions	NO_3^- confirmed
S^{2-} Solution of sample + $Pb(NO_3)_{2(aq)}$ **OR** Solution of sample + moist $Pb(CH_3COO)_2$ paper	Black precipitate Paper turns black	S^{2-} confirmed

QUALITATIVE ORGANIC ANALYSIS

This aspect is concerned with identification of element and functional groups present in organic compounds. There are four essential steps which are

1. Preliminary test
2. Solubility test
3. Detection of elements
4. Test for functional groups

PRELIMINARY TEST: This include physical state, smell and action of heat

a. **Physical state**
   liquids: hydrocarbons, alkanoic acids, carboxylic acids, esters, amine, aldehyde and ketones

   Solids: aromatic acids, carbohydrates, proteins, salts of organic acids and bases

b. **Smell**: fruity (esters) , fishy (amines)
c. **Action of heat**: substances that burn with a non- smoky flame (aliphatic/ saturated), substances that burn with a sooty/smoky flame (unsaturated)

Test for Functional Groups in Organic Compounds

To the sample,

Test	**Observation**	**Inference**
Alkane (single bond)		
Add bromine water	No visible reaction	Saturated hydrocarbon
OR		
Add dil. H_2SO_4 + $KMnO_4$	No visible reaction	Saturated hydrocarbon
Alkene (double bond)		
Add bromine water	Brown colour of bromine water decolourized	Unsaturated hydrocarbon
OR		
Add dil. H_2SO_4 + $KMnO_4$	Purple colour of $KMnO_4$ decolourized	Unsaturated hydrocarbon
OR		
Add ammoniacal copper(I) chloride solution	No visible reaction	Double bond confirmed
Alkyne (triple bond)		
Add bromine water	Brown colour of bromine water decolourized	Unsaturated hydrocarbon

OR Add dil. H_2SO_4 + $KMnO_4$ **OR** Add ammoniacal copper(I) chloride solution	Purple colour of $KMnO_4$ decolourized Yellow precipitate	Unsaturated hydrocarbon Triple bond confirmed
Alkanol (OH) Add sodium metal **OR** Add conc. H_2SO_4 + ethanoic acid + heat	Effervescence occurs and a colorless, odorless gas, neutral to litmus and produces a pop sound with lighted splint Sweet fruity smell	Gas is hydrogen Alkanol confirmed
Alkanoic acid (COOH) Add conc. H_2SO_4 + alcohol + heat **OR** Add saturated solution of Na_2CO_3	Sweet fruity smell Effervescence occurs and a colorless, odorless gas, turns blue litmus red and turns lime water milky is evolved	Alkanoic acid confirmed Alkanoic acid confirmed
Alkanal (CHO) Add Fehling's solution **OR** Add Tollen's reagent	Brownish red precipitate Silver mirror formed at the side of the test tube	Alkanal confirmed Alkanal confirmed
Alkanone (CO) Add Fehling's /Tollen's reagent **OR (iodoform test)** Add $iodine_{(s)}$ + $NaOH_{(aq)}$	No reaction Yellow precipitate (iodoform)	 Alkanone confirmed
Ester (-COOR) Add $NaOH_{(aq)}$ + phenolphthalein + heat	Pink colour turns colorless	Ester confirmed
Phenol Add $FeCl_{3(aq)}$ **OR** Add bromine water	Purple/violet coloration White precipitate	Phenol confirmed

Amine ($-NH_2$) Add dil HCl + $NaNO_{2(aq)}$ **OR** Add bromine water	 Effervescence(N_2 evolved), Yellow liquid No visible reaction White precipitate	 Primary amine Secondary amine Tertiary amine Aromatic amine
Amide ($-CONH_2$) Add $NaOH_{(aq)}$ + heat	Colourless gas with a choking smell, turns moist litmus paper blue and forms dense white fumes with conc. HCl	Gas is NH_3 from an amide

Food test for fats and oils

Test	**Observation**	**Inference**
Add Sudan (III) solution OR Add sample on a piece of paper OR Add Osmic acid	Red coloration Transclucent spot on paper Fat droplets turns black	Fat or oil confirmed Fat or oil confirmed Fat or oil confirmed

Food test for carbohydrates

Test	Observation	Inference
Reducing sugar(glucose) Add Fehling's solution + heat OR Add Benedict's solution	 Brick-red precipitate Yellow precipitate	 Reducing sugar confirmed
Non-Reducing sugar (sucrose) Add dil. HCl/H_2SO_4 + heat + Fehling's solution	Yellowish-red precipitate	Non-reducing sugar confirmed
Starch Add iodine solution OR	Blue-black colouration	Starch confirmed

Add iodine solution + KI/KI_2 + heat + cool	Blue-black colouration disappears on warming and reappears on cooling	

Food test for proteins

Test	**Observation**	**Inference**
Add Millon's reagent + heat OR Add Biuret's reagent ($NaOH + CuSO_4$) OR Add conc. HNO_3 + heat + conc. $NH_{3(aq)}$ Xanthoproteic test	White precipitate appears first but on warming brick red precipitate appears Purple coloration White precipitate appears then turns yellow on boiling and later turns orange when conc. $NH_{3(aq)}$ is added	Protein confirmed Protein confirmed Protein confirmed