COURSE GUIDE

SED 122
GENERAL CHEMISTRY FOR INTEGRATED SCIENCE 1

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INTRODUCTION

Welcome to this chemistry course. We suppose that many of you are taking this course because you want to strengthen your-background in chemistry. Armed with this strong background in chemistry, you can then proceed to advanced programmes in chemistry. The training of biologists, geoscientists, medical doctors, nurses, soil scientists, food scientists, chemical engineers, petroleum engineers and many others requires that students in those areas have a good exposure in chemistry.

We have packaged this chemistry course in such a way that you will learn chemistry using a technique peculiar to the open learning system. This technique has been adopted by the National Open University of Nigeria in most of the courses offered to students.

You will here learn the content of this chemistry course at a reasonable pace. You will need to master the language chemists use to describe the world around us. The language is simple, interesting and specific to the subject, chemistry.

We wish you success with the course and hope that you will find it both interesting and useful.

WHAT YOU WILL LEARN IN THIS COURSE

This course is titled “General Chemistry for Integrated Science l" this present course, you will be presented information in chemistry in a structured way to make learning easier. All the units follow the same pattern and so after the first few units, the rest will become easy to follow.

LEARNING OUTCOMES-AIMS AND OBJECTIVES

The board aims of this foundation chemistry course can be summarized thus: The course aims to provide you with chemistry content that will be sufficient for you in the integrated science program. Thus you will have solid foundation in chemistry which enable you go into an advanced science course needing a background of chemistry.

The objectives of this course are set out below. On completion of the course, you should be able to:

- distinguish between chemistry and the other science subjects.
- discuss the role of chemistry in our everyday living.
- apply the language of chemistry in describing the world around you.

- carry out simple chemical calculations.

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Unit 3	Periodic Table II: Gradations of Atomic Properties
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Unit 4	Liquids
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MAIN COURSE

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MODULE 1

UNIT 1 ELEMENTARY UNITS IN CHEMICAL REACTIONS

CONTENTS

1.0 INTRODUCTION

Chemistry is the study of matter and chemists in their investigations, study the properties and transformations of matter. Many materials that we use everyday, directly or indirectly are products of chemical research and examples of useful products of chemical reactions are limitless.

What then is matter? **Matter is anything that has mass and occupies space**. Matter is classified into solid, liquid and gas.

- Solids have fixed shapes and sizes. Glass, sand and most metals are examples of solids.

- Liquids have fixed volumes but no fixed shapes. Liquids take the shape of the containing vessel. Water and kerosene are examples of liquids.
- Gases flow and fill the entire space available. Air and cooking gas are examples of gases.

The above classification is commonly referred to as physical classification. Matter can also be classified into elements, compounds and mixtures. This latter classification is referred to as chemical classification. It is remarkable that all these substances, solids, liquids, gases, elements, compounds and mixtures are built up from simple basic Units.

What is the basic unit of matter?
What are the building blocks of matter and what laws govern the interaction of matter? The above are some of the questions that will be answered in this unit.

2.0 OBJECTIVES

At the end of this unit, you should be able to:

- define atoms, molecules, elements and compounds.
- differentiate between atoms and molecules as well as elements and compounds.
- explain why matter is said to be electrical in nature.
- state the postulates of Dalton's atomic theory and explain modifications to it.
- list and give relative masses and charges of the subatomic particles.
- state the laws of chemical combinations.
- write chemical symbols and formulae for common elements and compounds.
- determine chemical formulae from experimental data.
- write and balance simple chemical equations.

3.0 MAIN CONTENT

3.1 Elements, Compounds and Mixtures

An element is a pure substance, which cannot be split up into simple substances by a chemical reaction. A pure substance that can be broken down into elements is called a compound. There are over 100 chemical elements. Some occur naturally as free elements, mixed with other elements or compounds. Some are very rare, while most occur in combined state in compounds.

The table 3.1 below gives a list of some elements and their total abundance either as free elements or in combined states.

Table 3.1 List of some elements and their abundance

Element	**Abundance in percentage by mass**
Oxygen	49.1
Silicon	26.1
Aluminum	7.5
Iron	4.7
Calcium	3.4
Sodium	2.6
Hydrogen	0.88
Chlorine	0.19
Carbon	0.09

Elements can be broadly classified into two groups, metals and non-metals. Examples of metals are copper and iron. The general characteristics of metals are lustre, good conductor of heat and electricity. Metals can be rolled and hammered into sheets and drawn into wires. They are used for roofing and electrical cables respectively. All metals are solids at mom temperature except mercury which is a liquid at room temperature. About 75 percent of the elements are metals. Unlike metals, non metals do not have characteristic lustre. Many are gases at room temperature and others are solids except bromine which is a red brown liquid at room temperature.

Non metals are non-conductors of heat and electricity. They cannot be rolled into sheets or drawn into wires like the metals. Oxygen, nitrogen, carbon and iodine are examples of non metals.

- A compound is formed when two or more elements combine chemically in fixed proportion by mass. The properties of compounds are different from the elements of which they are formed. A lot of energy is often required to split compounds into the constituent elements. There are limitless numbers of compounds. Sodium chloride, (common salt), water and calcium trioxocarbonate (iv), (marble) are examples of compounds.
- A mixture is a physical combination of elements or compounds. The composition of a mixture varies and the components are separated by physical methods. Such physical methods include heating, cooling, dissolution, filtration and distillation. Air and petroleum are examples of mixtures.

• 3.2 The Particulate Nature of Matter

3.2.1 The Concept of Atoms and Molecules

Experiments show that very small quantities of matter have the same chemical properties or characteristics as larger ones of the same type e.g. a small iron nail and a big iron rod have the same chemical characteristics. For example both will rust when exposed to air and moisture and both will conduct heat and electricity.

- The smallest unit of an element that has the properties and characteristics of the element is the atom. The atom is the smallest unit of an element that can take part in a chemical reaction.
- The molecule is the smallest unit of a compound that has the characteristics of the compound. It is the smallest unit of the compound that can take part in a chemical reaction. The atom is to the element as molecule is to the compound.
- The behaviour of matter is explained using the above concept and chemical reactions are explained as combinations and rearrangement of atoms. The basic unit of matter in chemical reactions is the atom. Atoms and molecules are the building blocks of matter.

3.2.2 The Atomic Theory and Chemical Reaction

A theory is a tested proposal to explain an observed statement of facts. A theory serves as a guide to new experiments. When proved incorrect or inadequate by experiment, a theory is discarded or modified so that new experimental facts can be accounted for. This latter statement is true of Dalton's atomic theory. Dalton's atomic theory forms the basis of theory of the atom. It has been modified in the light of new experimental facts about the atom. The postulates of Dalton's atomic theory are:

- All elements are made up of small, indivisible particles called atoms.
- Atoms can neither be created nor destroyed.
- Atoms of the same element are alike in every respect and differ from atoms of all other elements.
- When atoms combine they do so in simple ratios.
- All chemical changes result from the combination or the separation of atoms.

Some important modifications to the theory are:

- Matter is composed of more fundamental particles, some of which are electrically neutral, some carry positive charge and some negative charge. This implies, that the atom is divisible.
- All atoms of the same element are not identical. The existence of atoms of the same element having different masses have been proved with the use of an instrument called mass spectrometer. Atoms of the same element having different masses are called isotopes.
- The smallest particle of an element that can take part in a chemical change is still the atom and in a chemical change atoms are neither created nor destroyed. Irrespective of masses, atoms of the same element have same chemical properties.

3.2.3 The Constituents of the Atom

Matter is electrical in nature. Evidences for this assertion came from results of experiments of early scientists like Faraday, Thompson and Millikan. The negatively charged particle in matter is the electron, It has negligible mass. The proton is the positively charged particle. It carries the same magnitude of charge as the electron and is very much heavier than the electron. The third particle is the neutron, a neutral particle with a mass approximately equal to that of the proton. These three particles are constituents of the atom except the hydrogen atoms that do not contain neutrons.

The number of each particle present in the atom varies from one element to another. For the atom (matter) to be electrically neutral, the number of protons must equal the number of electrons. Atoms of the same element will have the same number of protons and electrons but may have different numbers of neutrons. Such atoms will have different masses and are called isotopes.

Particle	Relative Mass	Relative Charge
Electron	5.5×10^4	-1
Proton	1.00 727	+ 1
Neutron	1.00867	0

3.2.4 Relative Atomic and Molecular Masses

Chemistry is a quantitative science and it is always desired to know the relative masses of substances that react as pure elements or compounds. The measurement of masses of atoms is not possible because they are very small. Their masses can however be compared to give relative atomic and molecular masses of elements and compounds. The relative atomic mass of an element is the mass of one atom of the element compared to one-twelfth (1/12) of the mass of one atom of carbon - 12 isotope.

On this scale 1 atom of carbon-12 isotope is given a mass of 12 atomic mass units. With the use of the mass spectrometer it has been possible to determine fairly accurately the relative atomic masses of elements.

The relative molecular mass of a compound is the sum of the relative atomic masses of the elements present in the chemical formula of the compound.

3.2.5 Chemical Symbols and Formulae

Chemical symbols and formulae are abbreviations used to represent elements and compounds. A chemical equation uses these symbols and formulae to summarise a chemical reaction. Chemical symbols consist of the first one or two letters of the name of the element. Some symbols do not correspond with the elements names; these symbols are derived from the Latin names of the elements. It is important that you know the symbols for as many of the common elements as possible. Table I .3 gives some examples of chemical symbols of elements.

Element	Symbol	Relative atomic mass
Hydrogen	H	1.00797
Oxygen	O	15.9998
Chlorine	Cl	35.453
Sodium	Na	22.990
Iron	Fe	55.847
Magnesium	Mg	24.305
Zinc	Zn	65.38
Cobalt	Co	58.933
Copper	Cu (cuprium)	64.456
Lead	Pb (plumbuim)	207.19
Gold	Au(Atrium)	196.967

- Hydrogen, oxygen and chlorine elements exist in nature as diatomic molecules. i.e. units consisting of two atoms that are chemically bound together. The latin names for copper, lead and gold are in bracket. The formula of a compound gives the proportion of the different elements present in the compound by mass. By the law of constant composition the proportion by mass of the different elements present is fixed for a pure sample of the compound irrespective of the method of preparations.

Example 1

Analysis of 1.630g pure sample of a compound of calcium and sulphur gives 0.906g Ca and 0.724 sulphur.

$$\%\ \text{Ca} = \frac{\text{mass of Ca x 100}}{\text{mass of sample}} = \frac{.906\ \text{x}\ 100}{1.630} = 55.6\%$$

$$\%\ \text{S} = \frac{\text{mass of sulphur}}{\text{mass of sample}}\ x\ 100$$

$$= \frac{.724}{1.630} = 44.4\%$$

Relative mass of Ca and S in the compound

$$= \frac{55.6}{Relative\ atomic\ mass\ of\ Ca} : \frac{44.4}{relative\ atomic\ mass\ S.}$$

$$= \frac{55.6}{40} : \frac{44.4}{32}$$

= 1.39 : 1.39

i.e. 1 : 1

The formula of the compound is CaS.

Having determined the relative atomic mass of the elements that make up the compound, you can proceed to determine the empirical formula of the compound. Since the relative masses of the elements in a compound depend partly on the masses of the atoms and also on the relative number of each atom of each element involved in the combination to form the compound.

Example 2

2.83g of a compound of lead (Pb) and surphur gives 2.45g lead and 0.38g sulphur. Determine the formula of the compound.

$\%\ \text{Pb} = \frac{2.45}{2.83}$ x 100 = 86.6

$\%\ \text{S} = \frac{0.38}{2.83}$ x 100 = 13.4

Relative number of Pb: S

$$= \frac{86.6}{207} : \frac{3.4}{32}$$

= 0.418 : .419
i.e. 1 : 1

Formula of compound is PbS.

Now determine a formula of a compound of calcium, carbon and oxygen with 40 percent Ca, and 48 percent oxygen.

3.2.6 Other Laws of Chemical Combination

In the previous section, the law of constant composition (proportion) is stated. Two more laws will be discussed in this section.

(a) The law of conservation of matter

This states that matter is neither created nor destroyed in a chemical reaction. A consequence of this law is that a chemical equation must always be balanced to account for all atoms present on the reactant side, and on the product side of the reaction.

For Example

$2H_2 + 0_2 \longrightarrow 2H_20$

$2Na + CL_2 \longrightarrow 2NaC1$

$H_2 + CL_2 \longrightarrow 2HC1$

$2C0 + 0_2 \longrightarrow 2CO_2$

(b) The law of multiple proportions

This states that when two different compounds are formed from the same two elements e.g. (CO and CO2 or SO2 and SO3) the masses of one element which react with a fixed mass of the other are in a ratio of small whole numbers.

Example 3

Analysis of two compounds of carbon shows that 1.33g oxygen combines with 1.00g of carbon, while in the second 6.64g oxygen combines with 2.49g of carbon. Show that this is in agreement with the law of multiple proportion.

For the 1st compound

1.00g carbon combines with 1.33g oxygen

For the second compound

2.49g carbon combines with 6.64g oxygen

$\frac{2.49g}{2.49g}$ carbon will combine with $\frac{6.64}{2.49g}$ g oxygen

1.00g carbon will combine with 2.667g oxygen.

Ratio of O mass combining with 1.0gC

= 1.33: 2.667

= 1 : 2 simple ratio.

1.00g carbon will combine with 2.667g oxygen.

Ratio of 0 mass combining with 1.0gC
= 1.33: 2.667
= 1 : 2 simple ratio.

3.4 Chemical Reactions and Equations

A chemical reaction involves the reshuffling of atoms from one set of combinations to another. In the reshuffling, one compound is converted to another. The smallest units that can take part in chemical reactions are the atoms and molecules. A chemical equation is often used to summarise the reaction that has taken place. A chemical equation gives the reactants and products of the reaction and the quantities of the reactants and products in correct ratio in accordance with the law of conservation of matter. Sometimes chemical equations give the physical states of the reactants and products.

(s) for solid
(1) for liquid
(g) for gas
(aq) for aqueous (solution in water)

The equation does not give you the rate of the reaction, the energy effect and the conditions necessary for the reaction to occur.

4.0 CONCLUSION

A chemical reaction involves elements, compounds and mixtures and leads to rearrangement or reshuffling of atoms. The elementary particles that form the basic units of elements and compounds in chemical reactions are atoms and molecules. Though atoms are composed of more fundamental particles, they are not split in chemical reactions. Dalton's atomic theory is the basis of modem atomic theory and explains satisfactorily the laws that govern chemical reactions of elements and compounds. Atoms and molecules are the building blocks of matter. The atom has a substructure of its own. This is the subject of the next unit.

5.0 SUMMARY

- The concept of atoms and molecules is introduced and explained.
- Dalton's atomic theory is discussed and the modification in the light of new experimental evidences highlighted.
- The atom has a substructure of its own and it is electrical in nature.
- In chemical reactions, atoms are not split but are reshuffled. They are neither created nor destroyed.
- Pure compounds contain fixed proportions of elements by mass.
- A chemical equation is a summary of the reaction.
- Chemical symbols and formulae allow for summary of a reaction in short hand form. A chemical equation must be balanced for the law of conservation enemy to be satisfied.
- Chemistry is a quantitative science and as such, amounts of matter used in reactions must be known.
- The relative atomic and molecular masses allow us to see on the atomic level, the relative amounts of is to products in chemical reactions.

6.0 TUTOR-MARKED ASSIGNMENT

1. (i) Calculate the percentage by mass of hydrogen in the following compounds:

   (ii) water (H_20) (ii) Hydrogen sulphide (H_2S) (iii) hydrogen chloride (HCT) and(iv) methane (CH_4)

2. (a) What are isotopes?
   (b) Natural chlorine consists of 75 percent of chlorine-35 and 25 percent of chlorine-37 isotopes. Calculate the relative atomic mass of chlorine element.
   (c) Explain why matter is neutral though

7.0 REFERENCES/FURTHER READING

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UNIT 2 ELECTRONIC CONFIGURATION 1 – STATIC MODEL

CONTENTS

1.0 INTRODUCTION

In the last unit the concept of atoms and molecules was introduced and discussed. Dalton's theory proposed that atoms are indivisible units of matter. The atom is the smallest unit of matter that can take part in a chemical change. Dalton's atomic theory satisfactorily explained the laws of chemical combination but could not explain why substances react the way they do. Why is oxygen able to react with maximum of two atoms of hydrogen as in water? Why do some elements exist only as diatomic molecules? Why are some elements very reactive and some inert? Dalton's law could not explain electrolysis neither could it explain the different masses of atoms of the same element.

Today we believe that the atom has a substructure of its own. The atom consists of much smaller particles that we call protons, neutrons and electrons. The relative masses and charges of these particles are given in unit 1.

What are the evidences in support of this new picture of the atom? How many of these particles are present in the atom of elements? How are these particles arranged in the atom? These are some of the questions that will be answered in this unit.

2.0 OBJECTIVES

At the end of this unit, you should be able to:

- discuss the scientific evidences for the electrical nature of the atom.
- give the number of subatomic particles in atoms of given elements.
- recall the relative masses and charges of the subatomic particles.
- define isotope
- explain the nuclear model of the atom proposed by Rutherford.
- state the limitations of Rutherford’s theory.
- explain the origin of electronic energy levels in atoms.
- write electronic shell configuration for elements and ions.

3.0 MAIN CONTENT

3.1 The Electrical Nature of the Atom

3.1.1 Historical Evidences

A very early evidence for the electrical nature of the atom came from Faraday. The result of Faraday's experiment on electrolysis showed that chemical change could be caused by the passage of electricity through aqueous solutions of chemical compounds. This evidence was closely followed by the discharge tube experiment A heated metal cathode emitted negatively charged particles. This beam of particles is called cathode rays and the particles are the electrons.

J.J. Thompson worked on cathode rays and confirmed that they are negatively charged. Their charge mass ratio was determined and found to be -l.76 x 10 coulomb/g^{-1}. It A. Millikan determined the electronic charge in his famous oil drop experiment in 1908. The charge of the electron is -1.602 x 10^{19} coulomb. The mass of the election was calculated. It is 9.11 x 10-^{28}g.

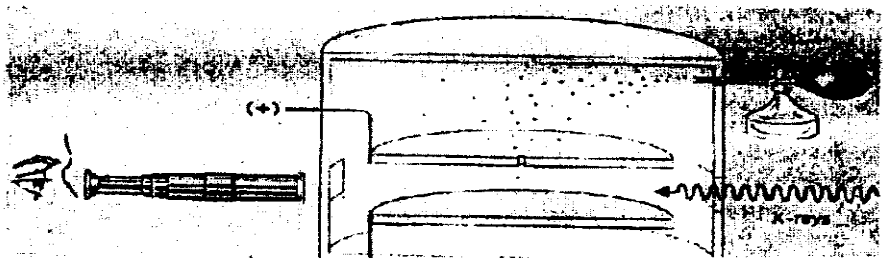


Fig. 2.1: Millikan oil drop experiment

The atom is electrically neutral, so it is reasonable to expect positively charged particles in the atom. The existence of positively charged particles was confirmed in a modified version of the set up used by J. J. Thompson by another scientist by name Goldstein. The charge/mass ratio of the positive particles are much smaller than for the electrons. The largest charge/mass was for hydrogen ion (IT) which now represents the fundamental particle of positive charge in the atom. The proton is about 1835 times as heavy as the electron and carry a charge equal but of opposite sign to that of the electron. Protons and electrons are present in all atoms.

The evidence of radioactivity by H. Becquerel further demonstrated the existence of subatomic particles. Becquerel observed that certain substances spontaneously emit radiations. The most important of these radiations are the alpha, beta and gamma radiations. Chadwick later confirmed the existence of a neutral particle in the atom and called this a neutron. This neutron has a mass approximately equal to that of the proton.

Table 2.1: Properties of subatomic particles

Particles	**Mass**	**Charge**		
	Grams	a.mu	Coulombs	electron charge
Proton	$1.67X10^{-24}$	1.007274	$+1.602X10^{-10}$	+ 1
Neutron	$1.68X10^{-24}$	1.008665	0	
Electron	$9.11X10^{-24}$	.000549	$-1.602X10^{-19}$	- 1

3.1.2 Atomic and Mass Numbers

Atoms of different elements have varying numbers of protons, electrons and neutrons. The atomic number is the number of protons in an atom of the element and for a neutral atom, the atomic number is also the number of electrons. The sum of the number of protons and neutrons is the mass number.

Isotopes are atoms of the same element with different mass numbers. We shall see in the subsequent units that the chemical property of an element depends on the number as well as the arrangement of the atomic electrons. This explains why atoms of the same element with different masses have the same chemical properties. $^{12}_{6}C$ and $^{14}_{6}C$ These are isotopes of carbon, mass number 12 and 14 and neutron numbers 6, and 8 respectively.

3.2 Atomic Models

3.2.1 The Nuclear Atom

The presence of subatomic particles posed challenge to early scientists. It was necessary to propose a model for the atom. J. J. Thompson proposed that the atom could be viewed as positive matter in which electrons are uniformly distributed to make it neutral at every point. This view was dropped because of the findings of two other scientists, Rutherford and Marsden. They bombarded a thin gold foil with fast moving alpha particles. They found that most of the alpha particles pass through the foil undeflected. Some were deflected as large angles while very few were sent back on their paths. Alpha particles are positively charged and many of them passing through the foil undeflected suggested that most of the gold foil was empty space.

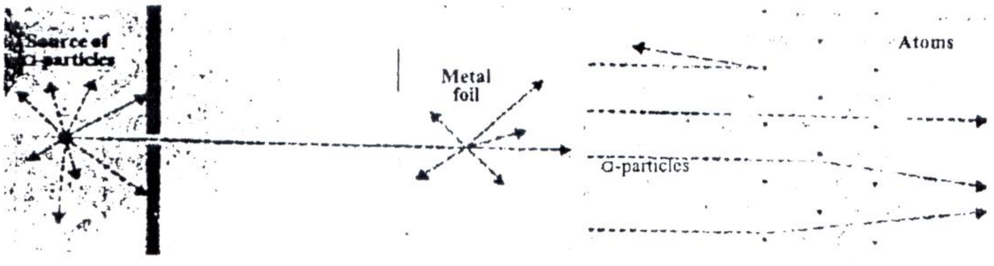
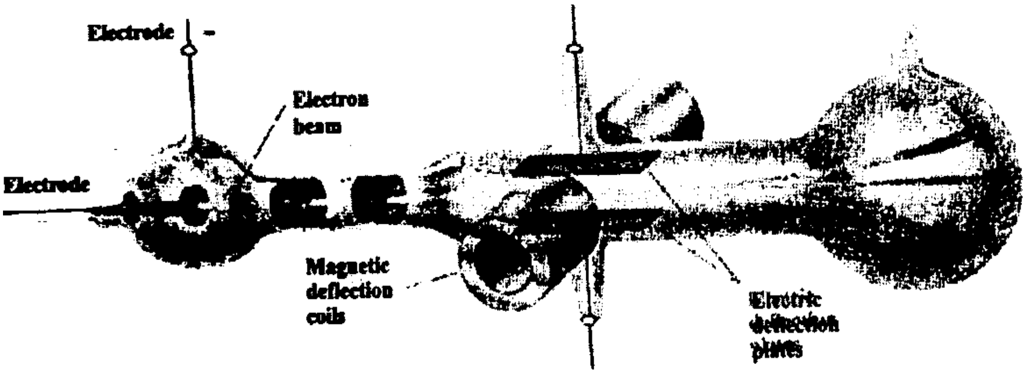


Fig. 3.2 Scattering of α-particles by a metal foil. (b) Deflection of α-particles by atomic nuclei

Rutherford, using the findings in the above experiment proposed a model for the atom. He proposed that the atom consisted of a tiny positively charged nucleus. The nucleus is centrally placed in the atom and the electrons surround it. The very small numbers of deflections of alpha particles suggest that the nucleus occupies a very small portion of the atom. For heavy particles such as the alpha particle to be so deflected suggests that the nucleus is a centre of heavy mass and positive charge. The protons and neutrons occupy the nucleus while the electrons we arranged around the nucleus and move in orbits around it, as planets around the sun. This is Rudierfords nuclear model of the atom.

By counting the number of alpha particles deflected in various directions, Rutherford was able to show that the diameter of a nucleus is about 1/100,000 times the diameter of the atom.



3.3 Schematic of a discharge tube. When a voltage is applied across electrodes that are sealed in a partially evacuated glass tube, the space between the electrodes glows.

3.2.2 Electronic Energy Levels

A major problem of Rutherford's model of the atom is that the electron (-vely charged) rotating around the nucleus (+vely charged) will lose energy continuously because of the electrostatic force of attraction of the nucleus on the electrons. This is not observed in practice. Infect energy absorption and emission by elements is discontinuous.

Have you ever seen the rainbow in the sky? The colours you see range from violet to red with no sharp line separating one colour from the other. This is a continuous spread of colours and is called a continuous spectrum. The different colours are component colours of light. In the laboratory the separation of light into its component colours also happens when light passes through a glass prism.

Light from the vapour of an element does not give a continuous spectrum. Each element has its own characteristic bright lines in particular positions. This is a line spectrum suggesting that light energy absorption or emissions by elements is only at particular energies characteristic of the element.

On the basis of the above observations Niel's Bohr proposed a model for the atom in which electrons move round the nucleus only in allowed orbits numbered serially. The orbit closest to the nucleus is assigned number 1 and is the orbit of lowest energy. Allowed transitions are transitions fern one orbit to another and will lead to emission or absorption of the energy difference between the orbits.

In the light of Bohr's model, there are electronic energy levels in atoms corresponding to different orbits of electron motion that are allowed. These energy levels are sometimes referred to as electron shells and designated K, I, M, N etc. corresponding to orbit numbers 1, 2, 3, 4, etc respectively.

Bohr's model gave satisfactory explanation of the hydrogen spectrum. The theory is limited however in its explanation for multi electron atoms. The wave mechanical treatment of the atom overcomes this limitation of Bohr's theory and is the subject of a subsequent unit.

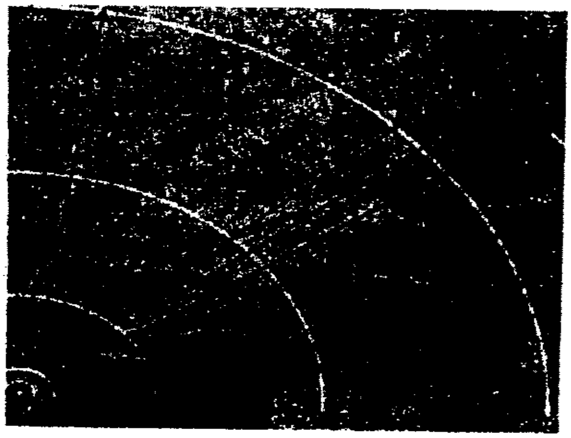


Fig. 2.4: The radii of Bohr's orbits.

3.2.3 Electronic Configuration

Electronic configuration gives the arrangement of electrons in the energy levels in the atom. In the ground state (most stable state of lowest energy) of the atom, electron assignment to energy levels is according to the following rules:

- the order of filling is K→L→M→N etc. The first shell is filled first before filling the second.
- the K shell can accommodate maximum of 2 electrons.
- the L shell can accommodate maximum of 8 electrons.
- the M shell accommodates 18 electrons.
- When 8 electrons are accommodated into the M shell however, there is extra stability and the next 2.
- electrons go into the N shell. Subsequently any electron goes into the M shell until it contains the maximum of 18 electrons.
- there are higher energy levels that accommodate a larger number of electrons than the K. L, M shells.

Table 3.2 Electron arrangement in some elements.

Element	symbol	Atomic number	Arrangement of electrons
Hydrogen	H	1	1
Helium	He	2	2
Lithium	Li	3	2, 1
Beryllium	Be	4	2, 2
Nitrogen	N	7	2, 5
Oxygen	O	8	2, 6
Fluorine	F	9	2, 7
Neon	Ne	10	2, 8
Sodium	Na	11	2, 8, 1
Magnesium	Mg	12	2, 8,2
Aluminum	Al	13	2, 8, 3
Silicon	Si	14	2, 8, 4
Sulphur	S	16	2, 8, 6
Chlorine	Cl	17	2, 8, 7
Argon	Ar	18	2, 8, 8
Calcium	Ca	20	2, 8, 8, 2

The electronic structure 2, 8,1 denotes 2 electrons in the K shell, 8 in the L shell and 1 in the M shell. Study the above table and note the following:

a) He and Ne; each has full shell of electrons and are both inert gases.
b) F and Cl; each has one electron short of a full electron arrangement.
c) H and Na; each has one electron outside a full shell electron arrangement. These pair of elements with similar electron arrangements also has similar chemical properties.

Now give the electron arrangement for each of the following:

Carbon (6) Phosphorus (15) and potassium (19)

Which of the elements in the table will have similar chemical properties with phosphorus?

Note

Isotopes have the same number of electrons and therefore have the same electron arrangement. Their chemical properties are the same.

3.2.4 Ion Formation

An ion is formed when an atom gains or loses electron.

e.g. H; F^- Na^+. The proton number is not affected by ion formation but the electron number increases or decreases.

Ion	Electron number	Electron arrangement
Li^+	2	2
F^-	10	2, 8
O^{2-}	10	2, 8
Mg^{2+}	10	2, 8
Cl^-	18	2, 8, 8
Ca^{2+}	18	2, 8,8

You will observe that these ions have closed or full shell electron arrangements and are stable ions. This observation and the MI shell arrangements for the inert gases, suggest that a full shell electron arrangement is a stable electron arrangement.

4.0 CONCLUSION

Matter is electrical in nature. This is supported by results of experimental investigations by early scientists. Three fundamental particles are present in the atom. The positive charge and mass are in a small centre of the atom. This centre is the nucleus and is surrounded by electrons moving round in allowed orbits of fixed energy. Electron assignments to these energy levels show that elements with similar properties have similar electron arrangement. The noble (inert) elements have closed shell electron arrangement suggesting that a closed shell arrangement of electrons is a stable configuration. This observation is used to explain chemical bonding.

5.0 SUMMARY

- The historical evidences for the electrical nature of the atom are many. Some of these evidences are revealed and discussed in this unit.
- Rutherford proposed a nuclear model for the atom. This model was modified by Bohr who proposed electronic energy levels, sometimes referred to as electronic shells around the nucleus.
- The electrons are accommodated into these shells starting with the shell of lowest energy.
- The rules for electron arrangement allow for elements with similar properties to have similar electronic configuration.
- The inert elements have closed shell electron arrangement

- Ions are formed when atoms gain or lose electrons. Stable ions also have closed shell electron arrangement.

6.0 TUTOR-MARKED ASSIGNMENTS

1. (a) State the contribution of Dalton and Rutherford to atomic theory.
   (b) Explain the limitations of Rutherford's model of the atom and show how Bohr's modification improved on Rutherford's model.
2. (a) Give the shell electron arrangement in the following elements. The atomic numbers are in brackets
   Be (4) Mg (12) K (19) Si (14) and CI (17)
   (b) Which of the listed elements in (a) above are likely to:

   i.) be metals.
   ii.) be non-metals
   iii.) have similar properties with carbon.
   iv.) form univalent ion of the type X.
   v.) have divalent ion of the type M^{2+}.

7.0 REFERENCES/FURTHER READING

Bah, S. T., Teibo, B. O., Onwu G. and Obikwere A. (2002). Senior Secondary Chemistry Textbook 2. Lagos. Longman Publishers.

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UNIT 3 THE NUCLEUS AND RADIOACTIVITY

CONTENTS

1.0 INTRODUCTION

The nucleus is the centre of positive charge and mass in an atom. The number of positive charges in the nucleus of a neutral atom determines the number of electrons that surround the nucleus. The arrangement of the electrons around the nucleus determines the chemical properties of the element.

Presently there is no theory that predicts the stability of the nucleus but a number of empirical observations suggest that the presence of neutrons account partly for nuclear stability. Except for the hydrogen atom, all nuclei contain neutrons. As the proton number increases, the neutron number also increases. For heavy elements the neutron number far exceeds the proton number in the nucleus.

e.g. ${}^{208}_{82}Pb$ has 82 protons and 126 neutrons.

Are all nuclei stable? Can nuclei react? What happens when a nucleus is unstable? Can a stable nucleus be made unstable? These are some of the questions to be answered in this unit.

2.0 OBJECTIVES

At the end of this unit you should be able to:

- state some empirical observations suggesting that neutrons are partly responsible for nuclei stability.
- define radioactivity.
- balance nuclear reaction equation.
- list the three types of radiations from a radioisotope giving their relative charges and masses.
- show the path of nuclear radiations in an electric field.
- define fusion and fission
- list some uses of radioactivity.
- explain why radioactive wastes are dangerous.

3.0 MAIN CONTENT

3.1 The Nucleus

The nucleus occupies a very small volume of the atom. According to Rutherford the nucleus is about 1/100,000 times the size of the atom. Except for hydrogen all elements have more than one proton in their nuclei. From your knowledge of elementary magnesium in Physics, you learnt that like charges repel each other whereas unlike charges attract each other.

How then can particles of the same charge stay so close in the nucleus without repelling. As mentioned in the introduction, no theory predicts nuclear stability. Empirical observations suggest that neutrons are partly responsible for nuclear stability. Listed below are these observations:

- All nuclei except hydrogen contain neutrons.
- As proton number increases neutron number also increases and far exceeds proton number for heavy elements.
- For some elements certain neutron-proton combination are unstable.
- Some neutron-proton combination are more abundant than others for the same element.

Table 3.1 gives some naturally occurring isotopes of carbon and oxygen and their abundance.

Table 3.1 Naturally occurring isotopes

Element	**Isotope**	**Natural abundance Carbon**
Carbon	$^{12}_{6}C$	98.89
	$^{13}_{6}C$	1.11
	$^{14}_{6}C$	trace
Oxygen	$^{16}_{8}O$	99.758
	$^{17}_{8}O$	0.038
	$^{18}_{8}O$	0.204

3.2 Nuclear Reactions

Nuclear reactions involve changes in the composition of nuclei leading to the conversion of one element to another. Contrary to the postulate of Dalton, an element is destroyed and a new one is created, when a radioactive element like uranium disintegrates.

Like in simple chemical reactions symbols are used to summarise the nuclear reaction in an equaiton. A nuclear reaction equation must be balanced to account for all particles and charge.

e.g. $^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$

Uranium 238, the uranium isotope having an atomic mass of 238 is radioactive. It breaks down emitting particles of Thorium and Helium.

3.2.1 Radioactivity

Some nuclei are unstable and will spontaneously emit radiations. This is called radioactivity and is an example of a nuclear reaction.

Very heavy metals with atomic number greater than 83 are radioactive. Some isotopes of light elements are also radioactive. A radioactive isotope is called a radioisotope Examples of radioisotopes are

$^{238}_{92}U$ and $^{14}_{6}C$

Radioisotopes decay at different rates. The half-life is a measure of the stability of a radioisotope.

Half life is the time it takes for half of the radioactive substance to decay $^{14}_{6}C$ and $^{238}_{92}U$

have half lives of 5570 and 4.47 x $\mathbf{10^9}$ years respectively. They are very stable compared to $^{23}_{11}Na$ and $^{214}_{84}P$ with half lives of 60 seconds and 10_1 microseconds respectively.

Some stable isotopes of elements can be made radioactive.

$$^{14}_{7}N + ^{1}_{0}n \rightarrow ^{14}_{6}C + ^{1}_{1}P$$

The above is a reaction that occurs instantly in the nature by the bombardment of cosmic neutrals upon nitrogen. It helps to keep $\mathbf{^{14}_{6}C}$ activity constant in the atmosphere.

3.2.2 Nuclear Radiations

A radioactive element in its decay emits radiations. These radiations are mainly of three types. They are called alpha (α); Beta (β) and gamma(γ) rays. The properties of cr, $ and y rays are summarised in the table 3.2. α, β and γ radiations are detected when they pass through an electric a magnetic field.

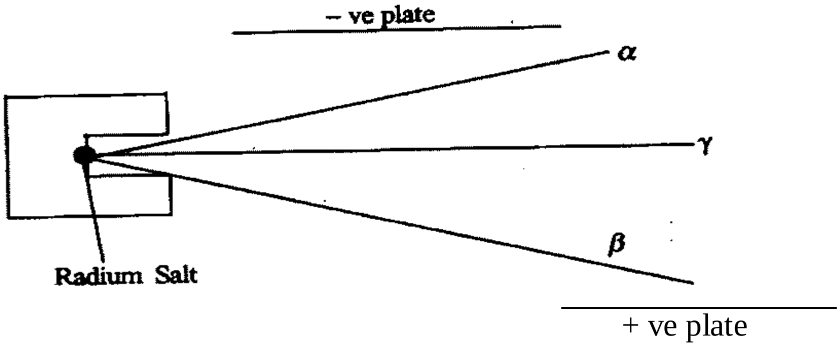

Fig. 3.2 Radiations

Table 3.2: Alpha (α), Beta (β)and gamma (γ) radiations

Alpha (α),	*Beta (β)*	*gamma (γ)*
i. positively charged	negatively charged	no charge
ii. they are helium nuclei $^{4}_{2}He$	they are electrons $^{0}_{-1}\beta$	electromagnetic radiation
iii. deflected towards a -ve2 pole	deflected towards a +ve pole	no effect of electric field.
iv. low penetrating power (absorbed by a thin sheet of paper).	higher penetrating power	very high penetrating power (absorbed by thick lead block).
v. high ionising power on gases	low ionising power on gases	very low ionising power on gases

The decay of heavy radioisotopes is usually accompanied by the emission of all the three types of radiations. Decay of light radio isotopes is usually accompanied by one or at most two of the three radiations.

3.2.3 Nuclear Fusion

This is a process in which two or more light nuclei combine to form a heavier nucleus with a release of energy. Nuclear fusion reactions require very high temperatures. Because of the very high temperatures in sun, fusion reactions are believed to occur and are responsible for the very high thermal energy that comes from the sun.

$$^{2}_{1}H + ^{2}_{1}H \rightarrow ^{4}_{2}He + \text{energy}$$

The above is suspected to be one of the reactions occurring in the sun. The energy from a fusion reaction initiates more reactions. If not controlled, can lead to explosion.

3.2.4 Nuclear Fission

Nuclear fission is a process in which the nucleus of a heavy element is spilt into two nuclei of nearly equal mass with a release of energy and radiation.

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{236}_{92}U \rightarrow ^{140}_{54}X_e + ^{93}_{38}S_\tau + 3^{1}_{0}n \text{ energy}$$

This fission process is initiated by the absorption of a neutron. The reaction releases three neutrons which can initiate more fission reactions. This is an example of a chain reaction. If the chain reaction is not controlled the fission reaction can become explosive. The energy from fission reactions are not as high as in fusion and the reactions in fission do not require very high temperatures to initiate. There are fission reactors used in generating electricity.

3.2.5 Other Examples of Nuclear Transformation Reactions

Nuclear transformation implies changing one element to another by the reactions of atomic nuclei. These reactions are many. Many isotopes of elements have been prepared by this method and used in chemical, biological and medical researches.

$$^{14}_{7}N \ + \ ^{4}_{2}He \rightarrow {}^{17}_{8}O \ + \ ^{1}_{1}H$$

$$^{27}_{13}Al \ + \ ^{4}_{2}He \rightarrow \ ^{30}_{15}P \ + \ ^{1}_{0}H$$

For heavy elements, neutron bombardment is more common because they are neutral and are not repelled by nuclei with the use of particle accelerators. Some heavy elements have been produced by the radioactive action of alpha particles, i.e. Helium atom.

$$^{238}_{92}U \ + \ ^{4}_{2}He \ \rightarrow \ ^{239}_{94}Pu \ + \ 3^{1}_{0}n$$

3.3 Uses of Radioactivity

Radioactivity has a large number of uses which include

- Treatment of cancer.
- Sterilising hospital equipment and instruments.
- Controlling the thickness of paper, metals and plastics.
- Controlling the filling of packets and containers.
- Tracing movement of a substance in a process.
- Sauce of energy as in fission.
- Radioactive dating e.g. carbon dating.

3.4 Hazards of Radioactivity

Gamma radiation is used to destroy cancerous cells. This is one use gamma radiation is put into. F-radiation also destroys healthy cells as well and too much exposure to it can do more harm than good. The extent of damage depends on the energy and type of radiation. The effect of radiation is also cumulative and small doses over a long period of time will also cause serious damage to biological systems.

Radioactive waste is very dangerous and must be disposed properly to avoid unnecessary exposure to its hazards.

4.0 CONCLUSION

The nucleus is sterilised in part by the presence of neutrals. Nuclear reactions unlike chemical reactions which involve valence electrons, nuclear reactions involve protons and neutrons. Nuclear reactions are much more exothermic than chemical reactions.

Many atomic nuclei are unstable. Some occur naturally and some are man-made. Unstable nuclei emit radiations with characteristic properties. The emitted radiations find application in various fields of human endeavour but also pose danger to users and non users alike. Radioactive waste must be properly disposed to avoid unwanted effects. Radioactive materials must always be handled with care.

5.0 SUMMARY

This unit discusses the properties and applications of the nucleus.

- When nuclei react they are transformed in the reactions.
- Unstable nuclei decay to give out three major types of radiations with characteristic properties. The radiations have properties that make them detectable.
- Reactions of atomic nuclei in fission and fission processes are important because of the high energies involved.
- Fission reactors are used to generate electricity.
- Radioactivity is used in medicine, agriculture, industry, engineering and chemical research laboratories.
- The emitted radiations can lead to unwanted effects in biological system.
- The environment must be protected by proper disposal of radioactive waste.

6.0 TUTOR-MARKED ASSIGNMENTS

1 (a) List the three types of radiations emitted by radioisotopes and show how they are differentiated in an electric field.

(b) Fill in the missing nuclei in the following nuclear equations.

i. $^{70}_{30}Z\text{n} \rightarrow {}^{0}_{-1}e + ?$

ii. $^{137}_{57}L\text{a} + {}^{4}_{2}He \rightarrow {}^{1}_{1}P + ?$

2. i. List far uses of radioactivity
   ii. Give some empirical observations suggesting that neutrons are partly responsible for nuclear stability.
   iii. How many protons, neutrons and electrons are present in the following?

   $^{60}_{27}C_0$, $^{43}_{19}K$ and $^{55}_{26}F_e$

7.0 REFERENCES/FURTHER READING

Bajah, S. T., Teibo, 8.0., Onwu, G. and Obikwere A., (2002). Senior Secondary Chemistry Textbook 2 Lagos Longman Publishers.

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UNIT 4 CHEMICAL BONDING 1: ELECTROVALENT, COVALENT AND CO-ORDINATE COVALENT

CONTENTS

1.0 INTRODUCTION

When elements combine, compounds are formed. The forces that hold atoms together in compounds are called chemical bonds. The combination of chemical elements to give a compound is a chemical reaction. Some elements are very reactive and exist in nature only in combined states, e.g. sodium in sodium chloride (common salt) and calcium in calcium trioxocarbonate (marble). Few *malts are relatively unreactive and exit rarely as free elements. They are called noble or rare elements, e.g. no compound of helium argon and neon exists. Most elements have intermediate reactivity and exist as free elements as well as in chemical compounds e.g. carbon occurs as graphile and diamond as well as in petroleum and many organic compounds. There are some non metallic elements that exist only as diatomic molecules in the free state. These dements also occur in combined states.

In the previous unit, you were shown that arrangement of electrons in atoms showed some correlation between electron arrangement and properties.

Li (2, 1), Na (2,8,1), K (2,8,8,1) all have similar configuration with one electron each in their outermost shell. They are metals. F (2,7) and C1(2,8,7) all need one electron to complete their outermost electron shell. They are non-metals. The inert or noble elements He (2), Ne (2,8), Ar (2,8,8) all have complete shell arrangement of electrons. The electron arrangement in stable ions of metals and non metals also show that complete shell of electrons is a stable configuration e.g. Na (2,8), P (2,8) and 0'- (2,8).

In chemical bonding therefore elements tend to Attain the noble or inert gas configuration. There are many types of bonds but in this unit you will faces on three types:

- electrovalent
- co-valent
- co-ordinate covalent

2.0 OBJECTIVES

At the end of this unit you should be able to:

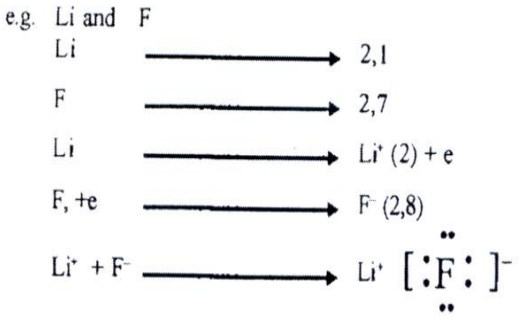
- explain deurovalent and covalent bonding.
- write electron dot formulae for compounds.
- list properties of electrovalent and covalent compounds.
- state the wet rule.
- explain the importance of noble gas configuration on chemical bonding.
- predict the type of bonds between atoms.
- explain the difference between covalent and co-ordinate covalent bonding.

3.0 MAIN CONTENT

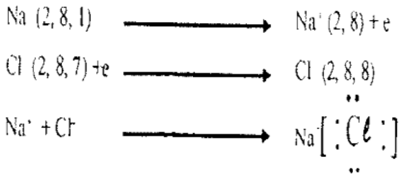
3.1 Electrovalent (Ionic) Bonding

When atoms interact for bond formation only the outermost parts of the atoms are in contact and so only the outer electrons in the outermost shell (valence electrons) are involved. The outermost shell electron arrangement is therefore very important in determining the type of bond. Electrovalent bonding involves electron transfer from the valence shell of one atom to the valence shell of the other.

One atom loses electrons to become positively charged and the other gains electrons to become negatively charged. The positively and negatively charged ions are called cations and anions respectively. The ionic bond results from the attraction between these oppositely charged ions. This type of bonding is usually between metals and non-metals.



The formula of LiF written as above is the electron dot formula (Lewis structure). The brackets around the fluorine are intended to show that all eight electrons are the exclusive property of the fluoride ion (F). Another example is the bond between sodium and chlorine.



As can be seen from the examples, the loss and gain of electrons result in inert gas configurations for the cation and anion. Except for helium, He (2) the inert gas configuration corresponds to eight electrons in the outershell. The electronic theory of valency as postulated by Kossel and Lewis 1716 was prompted by the remarkable stability of the rare gas elements. This stability is associated with the presence in the atoms of a group of eight electrons in the outer shell. This completeness appears to be the source of stability in rare gases.

The tendency for atoms to have eight electrons in their outermost shell (is explained by the octet rule). The octet rule states that:

- atoms tend to gain or lose electrons until there are eight electrons in their valence shell.

Note that the rule does not always hold. In cases like these, other stable configurations explain ion stability. The number of bonds to a particular atom depends on the number of electrons gained or lost to attain stable configurations for example.

$Ca\ (2, 8, 8, 2) \longrightarrow Ca^{2+}\ (2, 8, 8) + 2e$

$2C1\ (2, 8, 7) + 2e \longrightarrow 02\ CI \bullet (2, 8, 8)$

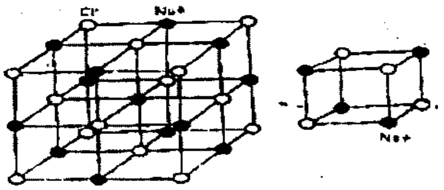
$Can + 2CI^{-} \longrightarrow \bullet \bullet [p.$

Now explain the bonding between sodium and oxygen. Write the fonnula of the compound.

3.1.1 Structure of Electrovalent Compounds

It is incorrect to speak of a molecule of an ionic compound like NaCI or LiF. Ionic compounds are usually solid consisting of regular arrangement of equal number of positive and negative charges. For Lif and NaCI there will be equal numbers of cations and anions.

This regular arrangement of cations and anions in the solid crystal is - called the lattice. The structure of sodium chloride is illustrated in the Fig 4.1



chlorine ion, Cl The basic unit of soditun ion, Na' the crystal Janice

Fig 4. l Lattice structure in sodium chloride

3.1.2 Properties of Electrovalent Compounds Electrovalent Compounds Are

i. usually crystalline solids, generally,
ii. soluble in water but generally insoluble in organic solvents like ether or kerosene,
iii. usually high melting point compounds,
iv. good conductors of electricity when molten or in aqueous solution but not when solid.

It is easier to explain the binding forces in the union between sodium ion and chloride ion, in the formation of sodium chloride since their opposite ionic charges attract each other. Its however difficult to comprehend the manner of bondage between non ionic or non polar atoms.

3.2 Covalent Bonding

It is easier to explain the binding forces in the union between sodium ion and chloride ion, in the formation of sodium chloride since their opposite ionic charges attract each other. Its however difficult to comprehend the manner of bonding between non ionic or non polar atoms.

Lewis, 1916, came up with a tenable explanation, suggesting that non ionic molecular compounds arise from the sharing of electrons among atoms, resulting in a form of bonding which was called the covalent bond. This type of bonding involves sharing of electron pairs rather than complete transfer. The shared electron pair/s Ware contributed by the two atoms involved in the bonding. The binding force results from the attraction of the shared electron pairs by the nuclei of the atoms involved in the bonding. This type of bonding is between non metals. Look at the following examples.

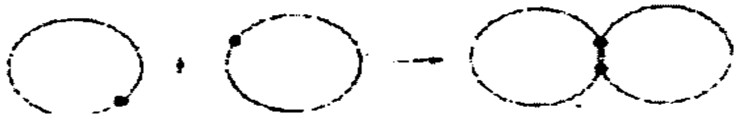


Fig 4.2: Formation of covalent hydrogen molecule

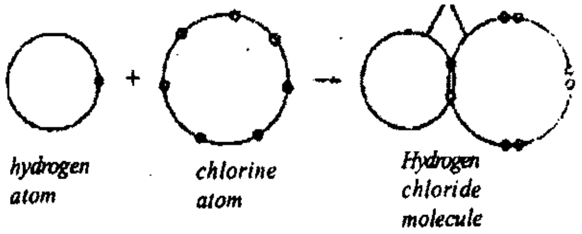


Fig 4,3 Shared electrons in hydrogen chloride

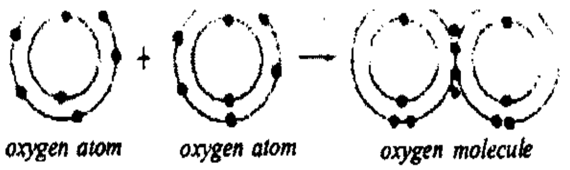


Fig 4.4 Formation of covalent oxygen molecule

An electron pair constitutes a bond and two pairs constitute a double bond as in the oxygen molecule. The number of electron pairs shared depend on the number of electrons each atom must share to attain an inert gas configuration. Covalent compounds form molecules and depending on the intermolecular forces between the molecules they may be gas (02, H2 HC1) or liquids (Br2 H20) or low boiling solids (candle wax).

3.2.1 Properties of Covalent Compounds

Covalent compounds are

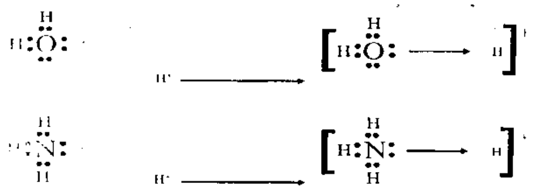
i. mostly liquids or gases,
ii. generally low melting when they are solids,
iii. generally not very soluble in water,
iv. generally soluble in organic solvents,
v. non conductors of heat and electricity unless they dissolve to form ions e.g. HC1

3.3 Co-ordinate (Dative) Covalent Bonding

In a covalent bond the shared electrons are donated and controlled by both atoms that are involved in the bonding. This is not the case with co-ordinate covalent bond. One atom donates the electron pairs but both atoms control the donated pair/s.

Once a coordinate covalent bond is formed it is not different from ordinary covalent bond. The electron "Pair is attracted by both nuclei of the bonded atoms. For this type of bond to be formed, one atom must have a lone paints of electrons. (Lone pairs are electron pairs that are not used in bonding to other atoms) The other atom must have a vacancy in its valence shell to accept the lone pair. The bond formation also results in inert gas configuration for both atoms.

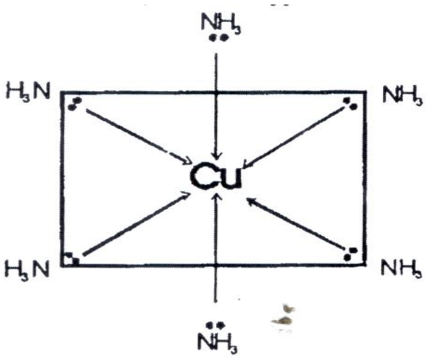
Let us look at the formation of the hydronium ion and the ammonium ion, H_3O+ and NH_4^+



Note that the positive charge is now controlled by the entire molecule. The dative bond cannot be differentiated once formed but sometime to indicate the presence of a dative bond in the molecule we indicate with arrow the direction of electron donation (→). The arrow points towards the acceptor atom.

Co-ordinate covalent bonding is common with metal complexes. The molecules donating the electron pairs are called ligands and the metal ion the central atom.

E.g. The formation of the ammonia complex with copper ions in solution. The octet rule



does not apply here. Why? It is a complex ion and has charges located on a group of atom.

4.0 CONCLUSION

In this unit we have discussed two major types of bonding, the electrovalent and covalent bonding. The properties of a compound depend on the type of bond between the constituent of atoms.

Electrovalent compounds consist of cations and anions in their solid structure. When an ionic solid dissolves in water or is melted, these ions become free. This explains why ionic compounds are good electrolytes when molten or in solution. Most covalent compounds are gases at room temperature because they consist of molecules held together by weak intermolecular forces. Dative bonding is an important type of bonding that helps to explain the structure and properties of additive compounds and complex ions.

5.0 SUMMARY

- An ionic bond is the electrostatic force that bonds opposite ions together.
- A covalent bond is as a result of attraction between shared electron pairs and nuclei of the atoms involved in the bonding.
- In bond formation elements attain that or stable configuration.
- Electrovalent bonding is between metals and non metals but covalent bonding is between non metals.
- A Lewis structure is very useful in chemical bonding because it shows the arrangement of the valence electrons of the atoms in a bond.
- For the formation of a dative bond, one atom must have lone pairs of electrons to donate, and the other atom must have available vacancy in the outer shell to accommodate the donated

electrons. Water and ammonia have lone electron pairs and take part in dative bonding with the hydrogen ion. Dative bonding is a special type of covalent bonding.

6.0 TUTOR-MARKED ASSIGNMENTS

1. (a) Define the following:
    - (i) cation, (ii) anion, (iii) ionic compound and (iv) valence electrons.
  - (b) Magnesium burns in oxygen to give a compound.
    - i. Explain the bonding between magnesium and oxygen.
    - ii. Write the formula of the compound.
    - iii. Calculate the percentage of oxygen in the compound.
2. a. State the octet rule.
  - b. Explain its significance in covalent bonding.
  - c. You are given the following : KCI, HBr, CI_3, NH_3, BF_3
    State with reasons in each, case, the compounds.
    - i. that will conduct electricity when melted
    - ii. that will dissolve in water to produce ions
    - iii. with lone pairs of electrons.
    - iv. that can accept a lone electric pair.

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UNIT 5 CHEMICAL BONDING II: METALLIC AND INTERMOLECULAR BONDING

CONTENTS

1.0 INTRODUCTION

There are three major types of bonds, electrovalent, covalent and metallic. Each bond type gives characteristic properties to the compounds that are formed.

In a previous unit we discussed electrovalent and covalent bonding. While electrovalent bonding is between metals and non metals, covalent bonding is between non-metals. In the formation of electrovalent and covalent bonds valence electrons play very important roles and each valence shell of the bonded Mans attain inert gas stable configuration. For metal- metal bond, the valence electrons are so few that electron sharing to attain electron octet is not possible. Electrovalent bonds cannot be formed as metals tend to lose electrons and not accept them. A negatively charged metal ion is not possible.

How then do we explain bonding in metallic solids? How can we explain the fact that metallic solids are good conductors of heat and electricity? What major differences are there in the structures of ionic and metallic solids? The above questions will be answered in this unit. We shall also explain the origin of intermolecular forces that hold covalent molecules together in the bulk sample and account for their special properties.

2.0 OBJECTIVES

At the end of this unit you should be able to:

- explain metallic bonding.
- account for metallic properties.
- list metallic properties.
- explain dipole-dipole interaction.
- account for properties of covalent compounds.
- explain hydrogen bonding.
- account for the high boiling point of water compared to hydrogen sulphide.

3.0 MAIN CONTENT

3.1 Metallic Bonding

Metals usually have 1, 2, or 3 electrons in their valence shells

Li (2, 1) or $1s^2\ 2s^1$
Na (2, 8, 1) or $1s^1\ 2s^2\ 2p^6\ 2s^1$
Mg (2, 8, 2) or $1s^2\ 2s^2\ 2p^4\ 3s^2$
Al (2, 8, 3) or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

In metallic bonding each metal atom contributes its valence electrons to form a 'cloud' or 'sea' of delocalised electrons. These electrons do not belong to any particular metal atom but will circulate freely through the metal lattice. The electrostatic attractions between the positive cores that form the metal lattice and the 'sea' or 'cloud' of electrons constitute the metallic bond. The above explanation of metallic bonding implies that the lattice forms a single large crystal. This accounts for the high strength of metals. There is no direction to metallic bond and so the metallic lattice can be distorted easily by hammering and drawing. Metals are malleable and ductile. The five moving electrons conduct heat and electricity by their movement.

The strength of the metallic bond depends on the attraction of the electron cloud to the positive cores in the metal lattice. The metallic bond strength increases with the number of valence electrons each metal contribute into the electron 'cloud'. Take the example of

Mg (2, 8, 2) $1s^2\ 2s^2\ 2p^6\ 3s^2$
Na (2, 8, 1) $1s^1\ 2s^2\ 2p^6\ 3s^1$

Sodium is a softer metal than magnesium because for sodium only one valence electron per atom but for magnesium two electrons are donated per atom to the electron cloud.

Following the above argument compare the strength of the metallic bonding in magnesium with that in aluminium (13).

For metals in the same group of the periodic table, metallic strength decreases down the group. The increase in atomic size down the group is not accompanied by any increase in electron cloud strength This listed properties of metals are explained by the metallic bonding just explained.

- High tensile strength
- Malleable and ductile.
- Conductor of heat and electricity.
- High density.
- Solid at room temperature except mercury that is a liquid at room temperature.
- Usually have shiny surface (lustre)

Now explain why metals have these three listed properties. (i) High strength (ii) ductility (iii) High density

3.1.1 Uses of Metals

The use of a metal for a particular application depends on a number of factors which include cost, availability, durability and performance. Table 5.1 gives the uses of some metals and the reasons for such uses.

Table 3.1 Uses of pure metals

Metal	Use	Reason for use
Copper	electrical cables	excellent conductor of electricity and very ductile.
Tin	coating tin cans	not poisonous
Aluminium	kitchen foil	very malleable

3.2 Intermolecular Bonding

The ionic and covalent bonds represent very strong interactions between atoms in a compound. In addition to these bonds there are other weaker attractive forces that exist between atoms and molecules. The existence of these weak attractive forces explains a number of physical properties of some compounds. Because these forces are usually between molecules they are called intermolecular forces. For example Van der Waal's forces, dipole-dipole attractions and hydrogen bonding.

3.2.1 Van der Waal's forces

Van der Waal's forces exist even between uncombined atoms and non polar molecules. A non polar molecule is one in which the electron pair for bonding is equally shared by the atoms involved in the bond formation. Examples of non polar molecules are N_2, $C1_2$, H_2, O_2 etc i.e. covalent bond between two identical atoms is a non polar bond. Non polar bond may also exist between unlike atoms if they have the same electronegativity. For example CO_2.

The movement of electrons around an atom can lead to a momentary shift of more electrons to one side of the molecule than the other. During this shift an imbalance in charge exists with one side of the molecule slightly positive and the other slightly negative. The positive end will attract the negative end of another molecule close to it. This attraction constitutes a bond. This attractive force may be strong but because it is for a short time its effect is generally very small.

The magnitude of this force increases with increasing number of electrons. This force is present between all molecules atoms and ions. Its effect can be very large when there are many electrons in the molecules or atoms. Take the case of the halogens (Group VII elements) fluorine, chlorine are gases, bromine is a liquid while iodine is a solid. Remember all of them exist as diatomic molecules and are only bonded together by van der weal forces, Van der Waal's forces are attractions between molecules which happen because of creation of temporary dipoles in all molecules. The very large number of electrons in bromine and Iodine allows for substantial cohesive force between bromine and iodine molecules making bromine liquid and iodine solid at room temperature. Van der Waal's force is sometimes called induced dipole- induced dipole attraction.

3.2.2 Dipole-Dipole Attractions

Covalent bonding between atoms of different elements will result in a polar bond. The shared electron pair will be more under the control of the more electronegative atom. Take the example of Ha. Chlorine is more electronegative than hydrogen. The shared pair of electron is controlled more by Chlorine. The chlorine end of the molecule will be slightly negative and the hydrogen end slightly positive e.g. HCl, H^{5+} - Cl^{5-}. The positive end of one hydrogen chloride molecule will attract the negative end of another molecule. This is dipole-dipole attraction.

Though dipole-dipole interactions are not as substantial as full ion-ion interactions, they are stronger than Van der Waal's forces.

The table 3.2 gives the normal boiling points of covalent compounds. Dipole interactions are only about one percent as strong as covalent and ionic bonds.

Compound	**Chemical formula**	**B.pt (K)** **ΔH_{vsp}**
Hydrogen	H_2	20
Nitrogen	N_2	77
Oxygen	O_2	90
Iodine	1_2	458
Bromine	Br_2	332
Water	H_20	373
Hydrogen chloride	HCl	188
Hydrogen bromide	HBr	206
Chlorine	Cl_2	239
Hydrogen sulphide	H_2S	223

The following are important observations from the table.

i. The boiling points of non polar compounds are very low compared to polar compounds.

ii. The boiling point increases with molecular mass for polar and non polar compounds.
$B_{p\tau}$ $H_2 < B_{pi}$ O_2 B_{pi} HCl $< B_{p\tau}$ HBr

iii. The boiling point of H_2O is abnormally high compared to those of other hydrides e.g. HC1, HBr, H_2S.

3.2.3 Hydrogen Bonding

Hydrogen bonding is a special type of dipole-dipole interaction which occurs when hydrogen is bonded to very small electronegative elements like N, F and Oxygen. In combination with these small electronegative elements, hydrogen carries a substantial positive charge. The attraction of this positive end with the negative end of another molecule will constitute a strong bond. This bond is the hydrogen bond. Hydrogen bond is about 5 to 10 times stronger than ordinary dipole-dipole interaction. It is not as strong as ordinary covalent bonds between atoms in a compound.

Hydrogen bonding is responsible for water being a liquid at room temperature rather than a gas. Hydrogen bonding explains the high boiling point of water compared to hydrogen sulphide (see table 5.2).

Hydrogen bonding explains why hydrofluoric acid is a weaker acid than hydrochloric acid.

4.0 CONCLUSION

From our discussion on chemical bonding, you can see that almost all atoms have the ability to combine with other atoms to form compounds. No wonder the numbers of compounds are limitless. In this unit the types of bonding discussed are interatomic and intermolecular bonding.

The metallic bonding is one of the major types of interatomic bonding and it explains very well the observed properties of metallic solids. Weak bonding exists between molecules, atoms and ions as a result of instantaneous shift in electron distribution around atoms in compounds. This weak bonding can be substantial leading to solid structure of covalent compounds at room temperature. Covalent bonding between unlike atoms will always lead to unequal share of bond electrons. Attraction between polar ends of molecules also account for the cohesive force between polar molecules, when the polar bond is between hydrogen and small electronegative elements. The cohesive energy of the dipole-dipole interaction can be very substantial. This may lead to abnormal behaviour of such compounds.

5.0 SUMMARY

- Metallic bonding is explained and used to account for properties of metals.
- Intermolecular faces are discussed and used to explain properties of polar and non polar compounds.
- Hydrogen bond is a very important dipole-dipole interaction that explains the behaviour of sane compounds of hydrogen with small electronegative elements.
- Hydrogen bonding is present in water. It explains why water is a liquid instead of a gas at room temperature.

6.0 TUTOR-MARKED ASSIGNMENT

1. (a) Give a description of the bonding between the atoms in a molecule of water. (b) Explain hydrogen bonding in water.
2 (a) Explain the type of bonding in magnesium metal.
   (b) Explain and account for the following properties of magnesium.
      i. High tensile strength and density.
      ii. Malleability and ductility
   (c) Compare and contrast magnesium with sodium chloride.

7.0 REFERENCES/FURTHER READING

Bajah, S. T., Teibo, B. 0., Onwu, G. and Obikwere A., (2002). Senior Secondary Chemistry Textbook 2 Lagos. Longman Publishers.

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MODULE 2

UNIT 1 PERIODIC TABLE I: CLASSIFICATION OF ELEMENTS

CONTENTS

1.0 INTRODUCTION

There are 103 known elements, some of which are familiar and others are rare. Recall that atoms are built of particles of three kinds: protons, neutrons and electrons. The nucleus of each atom is made of protons and neutrons. The number of protons (the atomic number) determines the electric charge of the nucleus, and the total number of protons and neutrons (the mass number) determines its mass. In a neutral atom the number of electrons surrounding the nucleus is equal to the atomic number.

The chemical and physical properties of an element are governed by the number and arrangement of the electrons. Several attempts have been made since 1817 to group elements together based on recurring properties such as atomic weight the most important step in the development of the periodic table was published in 1869 by Dmitri Mendelyeev, who made a thorough study of the relation between the atomic weights of the elements and their physical and chemical properties. The word periodic means recur at regular interval.

The initial arrangement has now been largely replaced following new knowledge about electronic structure of atoms. The present periodic table is based on the recurrence of characteristic properties when elements are arranged in order of increasing atomic number.

2.0 OBJECTIVES

At the end of this unit, you should be able to:

- identify the groups and periods of elements in the periodic table.
- list the first few members of each group.
- identify the blocks of elements.
- state general trends of some characteristics of elements.

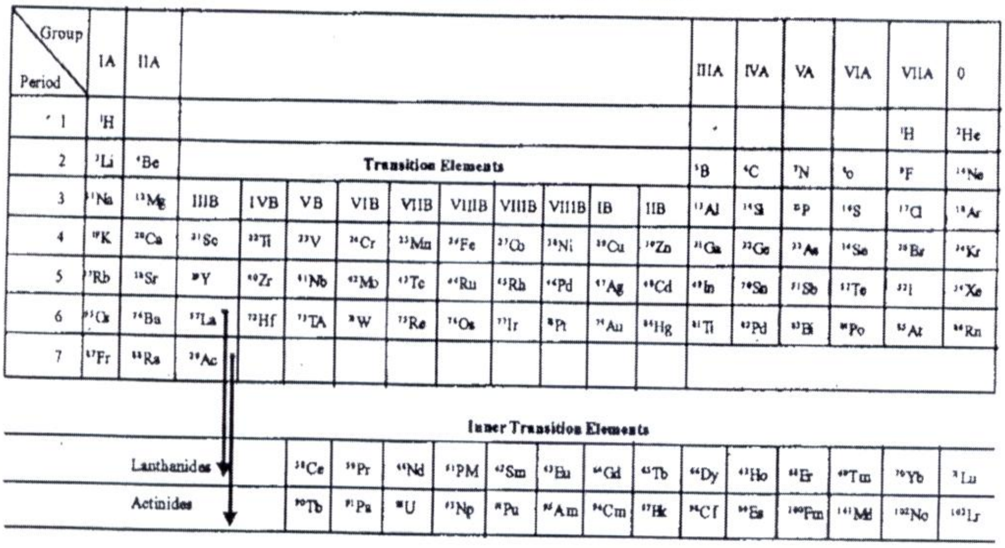
3.0 MAIN CONTENT

3.1 The Periodic Law and Periodic Table

The periodic law states that the properties of the chemical elements are not arbitrary, but depend upon the structure of the atom and vary with the atomic number in a systematic way. In other words, the properties of the elements are the periodic function of their atomic number. When elements are systematically arranged in order of increasing atomic number, certain characteristics recur at regular intervals.

The periodic recurrence of properties of the elements with increasing atomic number can be clearly shown by arranging the elements in a table, called the periodic table of the elements. The periodic table shows the arrangement of elements in seven horizontal rows and eight vertical columns as shown in table 3.1

Table 6.1 Periodic Table



Group / Period	IA	IIA											IIIA	IVA	VA	VIA	VIIA	0
1	^{1}H																^{1}H	^{2}He
2	^{3}Li	^{4}Be	Transition Elements										^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
3	^{11}Na	^{12}Mg	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	^{13}Al	^{14}Si	^{15}P	^{16}S	^{17}Cl	^{18}Ar
4	^{19}K	^{20}Ca	^{21}Sc	^{22}Ti	^{23}V	^{24}Cr	^{25}Mn	^{26}Fe	^{27}Co	^{28}Ni	^{29}Cu	^{30}Zn	^{31}Ga	^{32}Ge	^{33}As	^{34}Se	^{35}Br	^{36}Kr
5	^{37}Rb	^{38}Sr	^{39}Y	^{40}Zr	^{41}Nb	^{42}Mo	^{43}Tc	^{44}Ru	^{45}Rh	^{46}Pd	^{47}Ag	^{48}Cd	^{49}In	^{50}Sn	^{51}Sb	^{52}Te	^{53}I	^{54}Xe
6	^{55}Cs	^{56}Ba	^{57}La	^{72}Hf	^{73}TA	^{74}W	^{75}Re	^{76}Os	^{77}Ir	^{78}Pt	^{79}Au	^{80}Hg	^{81}Ti	^{82}Pd	^{83}Bi	^{84}Po	^{85}At	^{86}Rn
7	^{87}Fr	^{88}Ra	^{89}Ac															

Inner Transition Elements

Lanthanides	^{58}Ce	^{59}Pr	^{60}Nd	^{61}PM	^{62}Sm	^{63}Eu	^{64}Gd	^{65}Tb	^{66}Dy	^{67}Ho	^{68}Er	^{69}Tm	^{70}Yb	^{71}Lu
Actinides	^{90}Tb	^{91}Pa	^{92}U	^{93}Np	^{94}Pu	^{95}Am	^{96}Cm	^{97}Bk	^{98}Cf	^{99}Es	^{100}Fm	^{101}Md	^{102}No	^{103}Lr

3.2 Description of the Periodic Table

The periodic table contains rows of elements called periods and columns of elements called groups. The horizontal rows of the periodic table consist of a very short period (containing hydrogen and helium, atomic number 1 and 2), two short periods of 8 elements each, two long periods of 18 elements each, a very long period of 32 elements, and an incomplete period. The elements in the period have the same number of shells and the number of valence electrons increases progressively by one across the period from left to right. For example, members of period 2 are Li, Be, B, C, N, 0, F and Ne with 3, 4, 5, 6, 7, 8, 9, and 10 electrons respectively.

For all members of the period the additional electron is added to the second shell hence the name period 2. In general, every period starts with an element containing one electron in its outermost shell (e.g. Li, Na, K) and ends with an element whose outermost shell is completely filled (e.g. He, Ne, Ar - the noble or inert elements). The properties of elements change in a systematic way through a period. For example the first members of each period are all light metals that are reactive chemically, and this metallic character decrease across the periods which ends with unreactive inert gases.

The elements that appear in a vertical column belong to the same group or family. They have the same number of outer electrons or valence electrons and have closely related physical and chemical properties. The groups IA and ILA elements are usually placed at the left side of all the periods and MA, IVA, VA, VIA VILA and 0 (sometimes called VIIIA) elements are at the right side. The central elements of the long periods, called the representative elements have properties differing from those of the elements of the short periods. The transition metals are placed in groups 111B, IVB, VB, VIB,V11B, VIIIB, LB and JIB.

The actinides, which are in period 7, starting, with actinium (Ac) are known as artificial elements because they do not occur naturally but are formed during nuclear reactions. They are unstable and short-lived.

3.2.1 The Periods

The periods are numbered from! to 7 and the elements in the same period have the same number of electron shells Period 1 elements have one electron shell (K); period 2 elements have two electron shells (K,L); period 3 elements have three electron shells (K,L,M); etc. The number of valence electrons in the atoms of the elements in the same period increase progressively by one from left to right.

Across a given period, there is a progressive change in chemical properties. For example, metallic properties decrease across the period while non-metallic characteristics increases. The first three members of any period (Groups I to 3), except period I are metals while those of Group 4 to 7 and 0 are non-metallic in behaviour. Using period 3 as an illustration, sodium, magnesium and aluminium are metallic and form mainly ionic compounds and basic oxides. To the right of the period, phosphorus, sulphur and chlorine are non-metallic and form mainly covalent compounds and acidic oxides.

3.2.2 The Main Groups

The elements in groups IA to VILA and group 0 are generally referred to as the main Group elements.

Hydrogen is placed in group IA for convenience only because of the single electron but does not have similar characteristic with other members of the group.

a. Group IA, the alkali metals: Lithium (Li), Sodium (Na), Potassium (K), etc are light metals which are very reactive chemically e.g. reacts vigorously with cold water to liberate hydrogen gas. The atoms of each element have only one electron, which they readily donate and are strong reducing agent they react by losing this valence electron to form ionic or electrovalent bonds. The alkali metals are excellent conductors of electricity because the valence electrons are mobile. Na Na' + e-

The alkali metals are made by electrolysis of the molten hydroxides or chlorides. Because of their reactivity especially with water, the metals must be kept in an inert atmosphere of under oil.

Sodium metal catches fire when in contact with water, so avoid dropping it in the sink in the laboratory

The metals are useful chemical reagents in the laboratory, and they find industrial use in the manufacture of organic chemicals, dyestuffs and tetraethyl lead (the anti-knock agent in gasoline). Sodium is used in sodium - vapour lamps, and because of its high heat conductivity, in the stems of valves of airplane engines, to conduct heat away from the valve head.

b) Group IIA; the alkaline - earth metals: The metals of group HA are beryllium (Be) magnesium (Mg), calcium (Ca) etc and are much harder and less reactive than the alkali metals. They have two electrons in their outermost shell and react essentially by forming divalent ionic bonds. The compounds of all the alkaline - earth metals are similar in composition; they form oxides MO, hydroxides M $(OH)_3$, trioxocarbonates MCO_3, etc.

$Ca \longrightarrow Ca^{2+} + 2e^-$
$Ca + 2H_2O \longrightarrow Ca(OH)_1 + H_2$

The very long periods (6 and 7) are compressed into the table by removing fourteen elements each, called the inner transition or rare-earth metals (Z=58 to Z = 71 → lanthanides; and Z = 90 to Z l03 → actinides) and representing them separately below the table. The lanthanides, which are in period 6, starting from lanthanum (La) and ending with Lutetium (Lu) are rare metals that show a great resemblance to one another

(c) Group IIIA: Scroll (B) is a metalloid-intermediate properties between metals and non-metals; whereas aluminium (Al) and other members of the group are metals. Members of the group are trivalent since each of its atoms has three valence elections and forms electrovalent compounds.
$Al \longrightarrow Al^{3+} + 3e.$

Aluminium the only familiar element of the group reacts with steam at 600°C to give hydrogen gas. The oxide and hydroxide of aluminium are amphoteric - they have both acidic and basic properties.

$Al(OH)_3 + 3H_2SO_4 \longrightarrow AJ_2(SO_4)_3 + 3H_20$
an acid

$Al(OH)_3 \quad NaOH \longrightarrow NaAl(OH)_4$
a base

d) Group WA: Carbon (C), Silicon (Si), germanitun (Ge), Tin (Sn) and Lead (Pb) are members of this group. Their atoms each has four valence electrons and tend to form covalent compounds. Carbon is a non-metal, silicon and germanium are metalloids while tin and lead are metals showing a graduation from non-metallic to metallic character on going down the group.

Carbon and silicon form more stable +4 state compounds (CO_2 and S_1O_2 - Sand) while tin and lead form more stable +2 state compounds (SnO and PbO). The compounds of carbon and hydrogen called hydrocarbons form a large class of organic compounds used as fuels e.g. butane (C_4H_{10}), hexane (C_6H_{14}).

e) Group VA: Nitrogen (N) and phosphorus (p) are the familiar members of this group. Their atoms each has five valence electrons and show two common valence of 3 and 5. Both of them are non-metals. They are electron acceptors in their reactions and form several oxides e.g.

N_2O_3, N_2O_5, P_4O_6, P_4O_{10}

The oxides me acidic and they react with water to form acids

$N_20_5 + H_2O \longrightarrow 2HNO_3$ (trioxonitrate (v) acid)

Nitrogen and phosphorus react with hydrogen to form annnonia (NH) and phosphite (P113).

(f) Group VIA: Oxygen (o) and sulphur (s) are the familiar members of this group and are both non metals. They are electron acceptors and are oxidising agents e.g.

$2Mg + O_2 \longrightarrow 2MgO$
$2Na + S \longrightarrow Na_2 S$

The elements do not attack water but combine with hydrogen to give water and hydrogen sulphide

$2H_2 + O_2 \longrightarrow 2H_2O$
$H_2 + S \longrightarrow H_2S$

(9) Gram VIIA: Fluorine (F), Chlorine (CI), bromine (Br) and Iodine (I) are the members of this group. They are commonly called halogens. They are all non-metals and highly reactive. Their atom each has seven valence electrons and they are electron acceptors. The halogens show great similarity in their properties e.g.

i. all are non-metals gases.
ii. they exist as diatomic molecules e.g. l_2, Cl_2,
iii. they form ionic compounds when they react with metals but form covalent compounds with hydrogen.

$2Na + Cl_2 \longrightarrow 2NaCI$ ionic
$H_2 + Br_2 \longrightarrow 2HBr$ covalent

iv. their hydrides are soluble in water to give acids
$HBr(g) + H_2O \longrightarrow H_3O' + Br$
acidic

v. their reactivity decreases down the group e.g. while fluorine reacts explosively with hydrogen even in the dark, chlaine reacts slowly in diffuse light, bromine reacts slowly in bright light and iodine reacts incompletely even in bright light

Group O: Helium (He), Neon (Ne), Argon (Ar), are the familiar members of this group which are commonly referred to as rare gases or noble gases. They have no bonding electrons because the outermost shell is completely filled hence the group name zero. The rare gases are therefore unreactive and exist freely as monoatomic molecules in the atmosphere (about 1% of the composition of air).

Members of the group exhibit similar properties which are different from those of the halogens that come before them and alkali metals that come after them. This is a confirmation that the end of a period has been reached.

3.2.3 The Transition Elements

- This is a collection of elements with very similar behaviour and are usually placed between Groups II and III
- of the periodic table. All the transition elements have the following characteristics.
- They are metals with high tensile strength and high melting point.
- They exhibit variable valences.
- They form coloured ions.
- They have the ability to form complex ions. Several of them show catalytic element e.g. manganese and nickel

4.0 CONCLUSION

In this unit you have learned the periodic classification of elements. You should have teamed that when elements are arranged in order of increasing atomic number, certain properties recur at regular intervals. Furthermore, you should have learned that the periodic table of elements serve to justify the trend of behaviour exhibited by elements.

5.0 SUMMARY

What you have learned in this unit concerns the arrangement of elements in groups and periods based on their atomic number and properties. It has served to introduce you to the periodic Table. The units that follow shall use the atomic orbital model to further justify the classification and explain the gradation of properties of elements based on the periodic table.

6.0 TUTOR-MARKED ASSIGNMENTS

1. What is the atomic and mass number of an element?
2. Define the periodic law of elements.
3. From the elements carbon, oxygen, neon, sodium, magnesia aluminium, sulphur, chlorine and argon, name two which are

(i) in the same group (ii) in the same period (iii) reactive metals (iv) able to form acidic oxides (v) inert.

7.0 REFERENCES/FURTHER READING

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UNIT 2 ELECTRONIC CONFIGURATION II: ATOMIC ORBITAL MODEL

CONTENTS

1.0 INTRODUCTION

For a long time, it was believed that electrons were orbiting the nucleus in the same way the planets orbit the sun. You have learned in unit 2 about the contributions of Rutherford and Bohr to atomic structure in order to obtain a model of the atom. Their contributions went a long way to explain some of the observation about the atom. The Rutherford's model of an atom as consisting of a central positively charged nucleus and the negatively charged electrons some distance away from the nucleus is still acceptable. However, classical electromagnetic theory denies the possibility of any stable electron orbits around the nucleus.

In Bohr's model of the atom, the electron was restricted to being found in a definite region i.e. it had to be found in that orbit and nowhere else. In the Wave Mechanics Model, however, there is a slight chance that the electron may be located at distances other than in the restricted orbits. Despite this, we still accept Bolu's scheme for quantisation of energy in the atom and that the lowest energy level of the atom is the most stable state.

Although Bohr's contribution was remarkable, particularly his quantisation of energy, theory to explain the spectral lines for hydrogen atom; it has the following limitations.

- The Bohr model failed to account for the frequencies of the spectral lines for complex atoms other than hydrogen.
- The model also failed to satisfactorily explain why the Rutherford's atom did not collapse as predicted by the theory of electromagnetic radiation.

The present day picture of the atom is based on wave mechanical or quantum mechanical treatment. The treatment reflects on the wave-nature of the electron and the quantisation of energy in the atom. Although these treatments are fundamentally mathematical in nature, it describes the electron as point charge and that the density of the cloud at a specified point gives only the probability of finding electrons at that point.

We shall look at how this new thinking will help our understanding of the atom and the observed relation between electronic arrangement in atoms and the chemical behaviour of elements.

2.0 OBJECTIVES

At the end of this unit, you should be able to:

- describe the quantum theory of atomic orbitals.
- illustrate the shapes of the electron clouds corresponding to s - and p - orbitals.
- write electronic configurations of atoms based on the orbital model.
- relate the electron configuration of atoms to the classification of elements in the periodic table.

3.0 MAIN CONTENT

3.1 Quantum Theory of Atomic Orbitals

In the quantum model, the atom is still believed to consist of a central nucleus and orbiting electrons but the physical picture of the atom is by far different from the models proposed by Rutherford and Bohr. The quantum theory attempts to understand how electrons are arranged in the atom based on wave and quantum mechanics treatment. The electron is visualised as a point charge. The density of this point charge varies in different locations around the nucleus and gives a measure of the probability of finding the electron at a specified point.

The region or space, around the nucleus, in which an electron in a given energy level is most likely or probable to be found is defined as an orbital. So rather than describing a fixed Bohr orbit in which electrons are located, the modem theory gives a probability description of atomic orbitals. The results of the quantum mechanical treatment of the atom is summarised below.

3.1.1 The Principal Quantum Number (n)

You will recall that in the Bohr's model, each orbit is denoted by the principal quantum number. This designation is retained in the quantum model but to represent distinct energy levels and not shells or orbits. In other words, the quantum model recognises different quantised energy levels around the nucleus. Each principal quantum number (n) corresponds to a particular energy level and has integral values of 1, 2, 3, 4, etc. The principal quantum number n = I represent the lowest energy level that an electron can occupy. Electron with the largest ('n') value has the most energy and occupies the highest energy level; and therefore the most easily removable or ionisable electron. The maximum possible number of electrons in an energy level is given by $2n^2$. ($n \geq 1$).

3.1.2 The Subsidiary or Azimuthal Quantum Number, (l)

This subsidiary quantum number shows how many energy sub-levels are present in each electronic energy level. The subsidiary quantum number, 1, has integral values ranging from 0, 1, 2, ... (n-1). The sub-levels with 1 = 0, 1, 2, 3, are usually referred to s -, p-, d-and f-sub-levels respectively.

3.1.3 The Magnetic Quantum Number, (M)

The magnetic quantum number (m) gives the number of orbitals present in each sub-energy level and has integral values ranging from - l...0 ....+ l. For example, when l = 2, m has the values -2, -1, 0, 1, 2, i.e. the l = 2 sub-level (d-sublevel) has five orbitals; when l = 1, m has the values —I, 0 1 i.e. the 1= l sub-level (p -sub-level) has three orbitals (See Table 3.1)

3.1.4 The Spin Quantum Number (s)

The spin quantum number, (s) describes the fact that each orbital can only hold a maximum of two electrons with opposite spin. Their spin states are denoted by the spin quantum numbers -1/2 and +1/2.

Table 3.1 Atomic energy levels and sub-levels

Value of n	*Values of l*	*Number of sub-levels*	*Names of the sub4evels*
1	0	One	s
2	0 and I	Two	s and p
3	0, 1, and 2	Three	s, p, and d
4	0, 1, 2 and 3	Four	s, p, d,and f

3.2 The Shape of Atomic Orbitals

Contrary to Bohr's postulate, electrons do not move around the nucleus in definite paths, infact it is impossible to accurately determine the path of the electron. Rather, the location of electron is defined in terms of a probability which is described by the orbital. A region in space where there is a high probability of finding an electron in an atom is called an orbital. The density cloud of the electrons defines the shape of the orbital.

The electrons that move about to produce a spherical symmetrical cloud around the nucleus is an selection residing in an s-orbital. The p-electrons move about three axes, x, y and z that are at right angles to one another, producing a dumb-bell cloud around the nucleus along each axes. They are called the p-orbitals and are distinguished from each other by Px, Py and Pz in line with the direction of the electron cloud. Fig. 3.2 shows the shapes of the S, Px, Py and Pz orbitals. The geometrical representation of the d and f orbitals are more complex and beyond the scope of this programme.

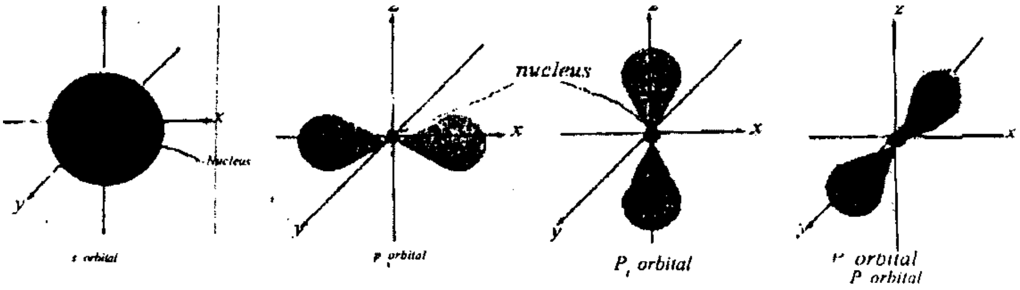


Fig ^{7}.2: "Seometrical representation of s- and p- orbitals

3.3 Electronic Configuration of Atoms — Orbital Model

The quantum treatment essentially deals with the arrangement of electrons in atoms. However, before we can apply the quantum numbers to express the electronic configuration of atoms, there are two important rules that you should be familiar with.

a. Pauli Exclusion Principle states that two electrons in the same atom cannot have the same values for all four, quantum numbers.

The principle simply means two electrons in an atom cannot behave in an identical manner.

b. ‘Hund's Rule; states that, in the filling of orbitals, electrons occupy each energy level singly before electron pairing takes place.

The way in which electrons are arranged in an atom is determined by the order of the sub-levels on a scale of increasing energy level. This is so because electrons are found in the lowest possible energy level, the ground state which is the most stable state of an atom. A simple representation of the orbitals on an energy scale is given in Fig. 3.3

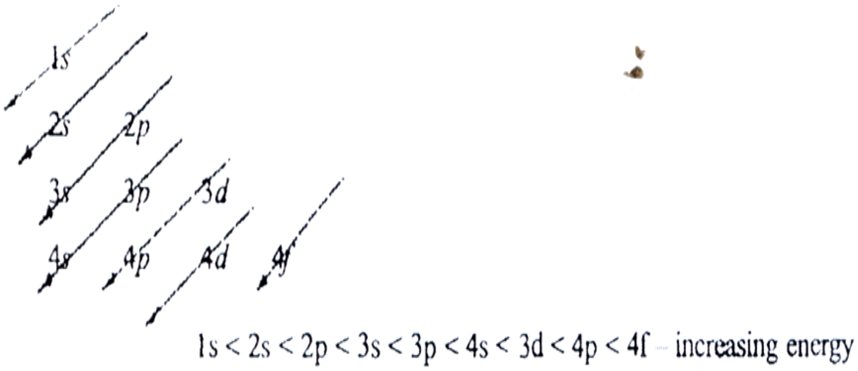


Fig 7.3; Atomic orbitals on an energy scale.

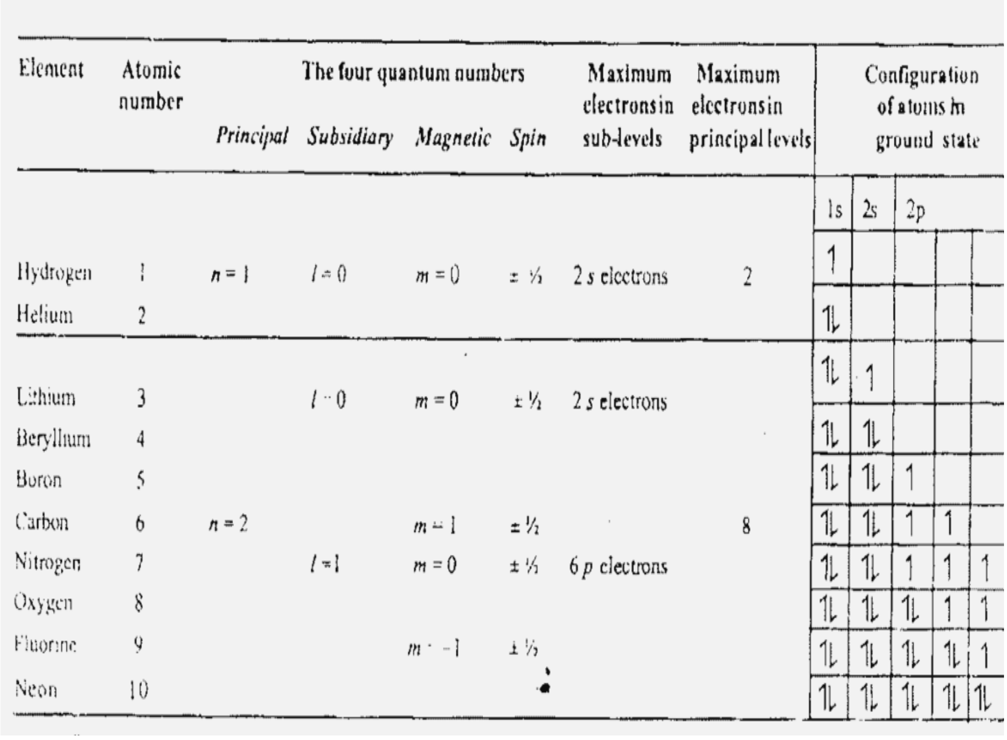
The maximum number of electrons in each sub-level is determined by the number of orbitals present in the sub-level bearing in mind that each orbital can take a maximum of two electrons with opposite spin. Table 3.2 gives the number of orbitals and maximum number of electrons in each sub-level.

Table 3.2 The number of orbitals and maximum electrons In the sub-level.

Value of l	Name of sub-level	Values of m	Number of values of m	Number of orbitals	Maximum number of electrons
o	5	0	One	One	2
I	p	1, o, - 1	Three	Three	6
2	d	2, 1, 0, -1, -2	Five	Five	10
3	f	3, 2, 1, 0, -1, -2, -3	Seven	Seven	14

Starting with hydrogen, which has an atomic number of!, the electron occupies the Is sub-level and this process of electronic occupation continues with increased atomic number according to the order of increasing orbital energy levels. To keep a check on the spin of the electron, arrows of opposite spins are used to distinguish two electrons in an orbital. Table 3.3 gives the electronic configuration of the first ten elements using the four quantum numbers.

Table 3.3 Electronic configurations of the atoms of elements using the four quantum numbers.



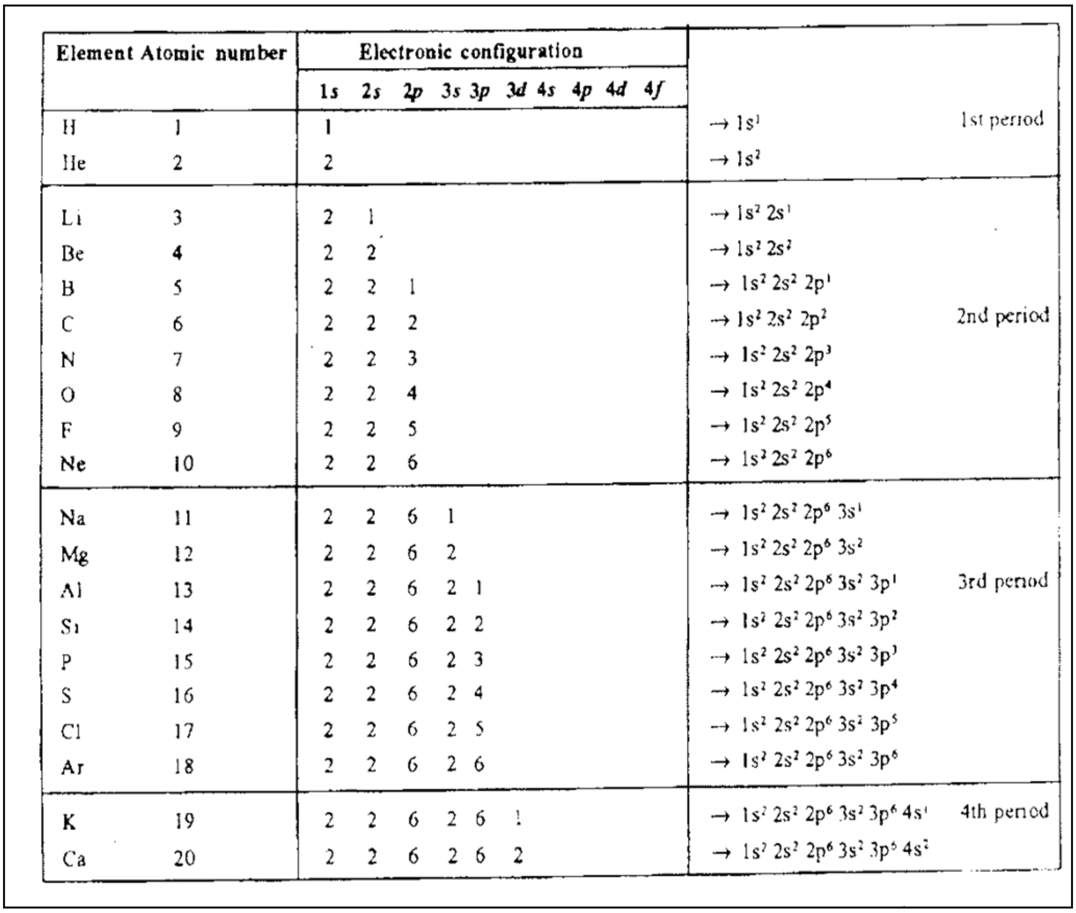
Element	Atomic number	The four quantum numbers				Maximum electrons in sub-levels	Maximum electrons in principal levels	Configuration of atoms in ground state				
		Principal	*Subsidiary*	*Magnetic*	*Spin*			1s	2s	2p		
Hydrogen	1	$n = 1$	$l = 0$	$m = 0$	$\pm \frac{1}{2}$	2 *s* electrons	2	↑				
Helium	2							↑↓				
Lithium	3		$l = 0$	$m = 0$	$\pm \frac{1}{2}$	2 *s* electrons		↑↓	↑			
Beryllium	4							↑↓	↑↓			
Boron	5							↑↓	↑↓	↑		
Carbon	6	$n = 2$		$m = 1$	$\pm \frac{1}{2}$		8	↑↓	↑↓	↑	↑	
Nitrogen	7		$l = 1$	$m = 0$	$\pm \frac{1}{2}$	6 *p* electrons		↑↓	↑↓	↑	↑	↑
Oxygen	8							↑↓	↑↓	↑↓	↑	↑
Fluorine	9			$m = -1$	$\pm \frac{1}{2}$			↑↓	↑↓	↑↓	↑↓	↑
Neon	10							↑↓	↑↓	↑↓	↑↓	↑↓

3.4 Electronic Configuration and Periodic Classification of Elements

One of the advantages of the electronic configuration of elements using quantum numbers is that it showed the basis for the periodic classification of element. In other words, the key to the periodicity of elements lies in the electronic configurations of their atoms.

The orbital arrangement of electrons clearly showed the great usefulness of the Period Table as it explains the groups and characteristic properties of elements. The correlations between electronic configuration and the physical and chemical behaviour of elements will be discussed in details in subsequent units. Table 7.4 shows the electronic configurations of the first twenty elements of the Periodic Table.

Table 3.4 The electronic configurations of the first twenty elements of the Periodic Table



Element	Atomic number	Electronic configuration											
		1s	2s	2p	3s	3p	3d	4s	4p	4d	4f		
H	1	1										→ $1s^1$	1st period
He	2	2										→ $1s^2$	
Li	3	2	1									→ $1s^2 2s^1$	
Be	4	2	2									→ $1s^2 2s^2$	
B	5	2	2	1								→ $1s^2 2s^2 2p^1$	
C	6	2	2	2								→ $1s^2 2s^2 2p^2$	2nd period
N	7	2	2	3								→ $1s^2 2s^2 2p^3$	
O	8	2	2	4								→ $1s^2 2s^2 2p^4$	
F	9	2	2	5								→ $1s^2 2s^2 2p^5$	
Ne	10	2	2	6								→ $1s^2 2s^2 2p^6$	
Na	11	2	2	6	1							→ $1s^2 2s^2 2p^6 3s^1$	
Mg	12	2	2	6	2							→ $1s^2 2s^2 2p^6 3s^2$	
Al	13	2	2	6	2	1						→ $1s^2 2s^2 2p^6 3s^2 3p^1$	3rd period
Si	14	2	2	6	2	2						→ $1s^2 2s^2 2p^6 3s^2 3p^2$	
P	15	2	2	6	2	3						→ $1s^2 2s^2 2p^6 3s^2 3p^3$	
S	16	2	2	6	2	4						→ $1s^2 2s^2 2p^6 3s^2 3p^4$	
Cl	17	2	2	6	2	5						→ $1s^2 2s^2 2p^6 3s^2 3p^5$	
Ar	18	2	2	6	2	6						→ $1s^2 2s^2 2p^6 3s^2 3p^6$	
K	19	2	2	6	2	6		1				→ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	4th period
Ca	20	2	2	6	2	6		2				→ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	

4.0 CONCLUSION

In this unit you have learned about the quantum mechanical treatment of the electronic structure of the atom. This is a follow up to what you learned about Rutherford and Bohr models of the atom. You should have also learned that the position of electrons can be defined only in tarns of the probability of finding it in a region in space referred to as orbitals. Furthermore, you learned about the four quantum numbers used for characterising an electron. You need to be aware of how to write the orbital electronic configurations of elements based on these four quantum numbers.

5.0 SUMMARY

What you have learned in this unit concerns the quantum treatment of electrons in atoms and how each of the four quantum numbers can be used to fully characterise an electron. It has served to introduce you to orbital electronic configuration. The unit on Period Table 11 shall build upon this treatment of the electrons in the atoms of elements.

6.0 TUTOR-MARKED ASSIGNMENTS

1. Write the electronic configuration of:
   (i) Aluminium (ii) Calcium (iii) Argon
   (iv) Oxygen (v) Carbon
2. Using electronic configuration account for the occurrence of periods and groups in the Periodic Table.
3. Define orbitals and sketch the s— and p— orbitals.

7.0 REFERENCES/FURTHER READING

Baja S. T., Teibo, B. 0., Onwu, G. and Obikwere, A. (2002). Senior Secondary Chemistry Textbook 3 Lagos. Longman Publishers.

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UNIT 3 PERIODIC TABLE II: GRADATIONS OF ATOMIC PROPERTIES

CONTENTS

1.0 INTRODUCTION

The periodic table consists of elements arranged in groups and periods based on their atomic number. The number of electrons in a neutral atom of an element gives the atomic number. The electrons occupy electronic shells. The elements in any period have the same number of electronic shells and the number of valence electrons increases progressively by one across the period from left to right. When elements are arranged in this way, it was observed that the properties of elements recur at regular intervals. That is, the properties of the elements are a periodic function of their atomic number. The periodic properties give rise to vertical columns of groups or families of elements with the same number of outer or valence elements. The elements in any group have closely related physical and chemical properties. One great advantage of this, is that it is only necessary to learn the properties of each group rather than the properties of each individual element.

Recall that the Bohr theory of shells was the basis for the arrangement of electrons in atoms. Recent discovery about the atoms have necessitated a revision of this idea. The electrons in atoms are now believed to occupy regions in space around the nucleus called orbitals rather than fixed shells. Orbitals are simply regions in space around the nucleus where the probability of finding an electron is high, they are usually denoted by s,p,d, and f- orbitals.

2.0 OBJECTIVES

At the end of this unit, you should be about to:

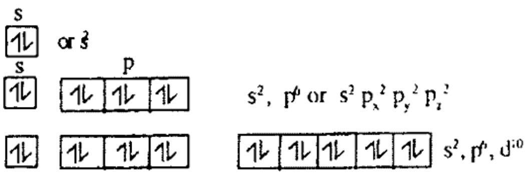
- explain the general trend in periodic properties down a group and across a period based on electronic configuration.
- define atomic radii, ionic radii, ionization energy electron affinity and electronegativity.
- explain the trend of the above properties within the periodic table.

3.0 MAIN CONTENT

3.1 Brief Revision of the Electronic Configuration

Electrons in atoms occupy various energy levels(n) and sub-energy levels (s,p,d,f - orbitals). Recall the Pauli's exclusion principle and Hund's rule in the arrangement of electrons in the energy levels. The summary below will refresh your memory.

Energy level	Number of sub-levels
1	s
2	s,p
3	s,p,d
etc	etc



For example, sodium with atomic number 11 will have the electronic structure ls^2 $2s^2$ $2p^6$ $3s^1$; magnesium with atomic number 12, ls^2 $2s^2$ $2p^6$ $3s^2$ and so on.

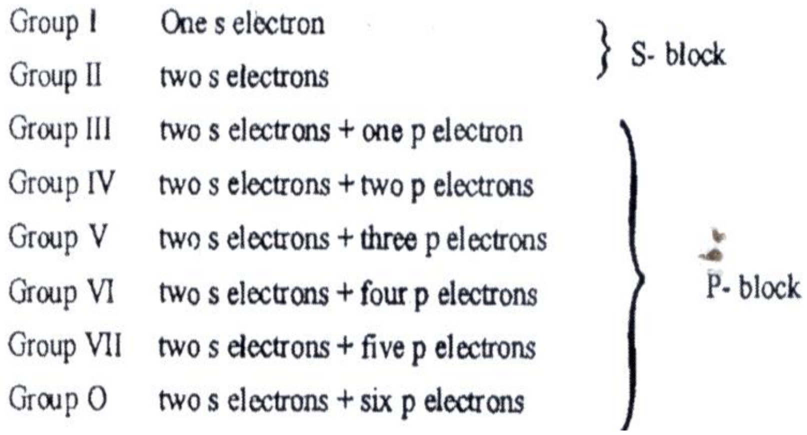
3.2 Electronic Configuration and the Periodic Table

The chemical and physical properties of an element are governed by the number and arrangement of the orbital electrons, that is, by the atomic number. The modem periodic classification (The Periodic Table) groups atoms of the element according to their electronic configurations.

Elements with one s electron in their outer shell are called Group I (the alkali metals) and elements with two s electrons in their outer shell are called Group II (the alkaline earth metals). These two groups are known as the s block elements, because their properties result from the presence of s electrons. Elements with three electrons in their outer shell (twos electrons and ones electron) are called Group 111, and similarly Group IV elements have four outer electrons, Group V elements have five outer electrons, Group VI elements have six outer electrons and Group VII elements have seven outer electrons. Group 0 elements have a full outer shell of electrons so that the next shell is empty; hence the group name. Groups III, IV, V, VI, VII and 0 all have p orbitals filled and because their properties are dependent on the presence of p electrons, they are called jointly the p block elements.

d block, or transition elements. In these, d electrons are being added to the penultimate shell (one shell .before the outer shell). Finally, elements what f orbitals are filling are called the f block, and here the f electrons are entering the antepenultimate (or second shell from the outer shell) shell.

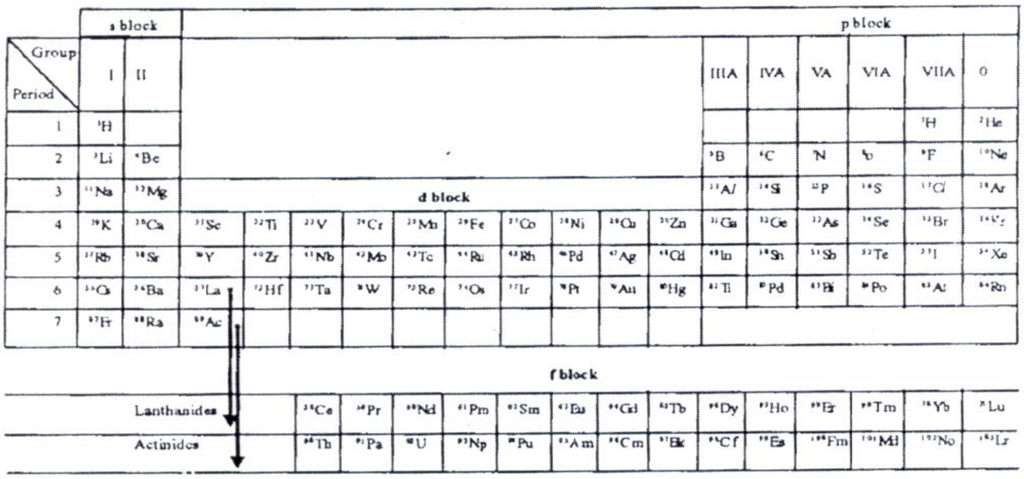
A summary of the block arrangement of elements based on the outermost energy levels for s- and p- block elements; and the orbitals being filled for d- and f- block elements.



Group I	One s electron	S- block
Group II	two s electrons	
Group III	two s electrons + one p electron	P- block
Group IV	two s electrons + two p electrons	
Group V	two s electrons + three p electrons	
Group VI	two s electrons + four p electrons	
Group VII	two s electrons + five p electrons	
Group O	two s electrons + six p electrons	

Transition Elements: d-orbitals being filled with electrons - d-block Inner Transition Elements: f-orbitals being filled with electrons - f-block (Lanthanides and Actinides)

Table 3.1 Periodic Table



	s block												p block					
Group / Period	I	II											IIIA	IVA	VA	VIA	VIIA	0
1	^{1}H																^{1}H	^{2}He
2	^{3}Li	^{4}Be											^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
3	^{11}Na	^{12}Mg	d block										^{13}Al	^{14}Si	^{15}P	^{16}S	^{17}Cl	^{18}Ar
4	^{19}K	^{20}Ca	^{21}Sc	^{22}Ti	^{23}V	^{24}Cr	^{25}Mn	^{26}Fe	^{27}Co	^{28}Ni	^{29}Cu	^{30}Zn	^{31}Ga	^{32}Ge	^{33}As	^{34}Se	^{35}Br	^{36}Kr
5	^{37}Rb	^{38}Sr	^{39}Y	^{40}Zr	^{41}Nb	^{42}Mo	^{43}Tc	^{44}Ru	^{45}Rh	^{46}Pd	^{47}Ag	^{48}Cd	^{49}In	^{50}Sn	^{51}Sb	^{52}Te	^{53}I	^{54}Xe
6	^{55}Cs	^{56}Ba	^{57}La	^{72}Hf	^{73}Ta	^{74}W	^{75}Re	^{76}Os	^{77}Ir	^{78}Pt	^{79}Au	^{80}Hg	^{81}Tl	^{82}Pd	^{83}Bi	^{84}Po	^{85}At	^{86}Rn
7	^{87}Fr	^{88}Ra	^{89}Ac															

f block

Lanthanides	^{58}Ce	^{59}Pr	^{60}Nd	^{61}Pm	^{62}Sm	^{63}Eu	^{64}Gd	^{65}Tb	^{66}Dy	^{67}Ho	^{68}Er	^{69}Tm	^{70}Yb	^{71}Lu
Actinides	^{90}Th	^{91}Pa	^{92}U	^{93}Np	^{94}Pu	^{95}Am	^{96}Cm	^{97}Bk	^{98}Cf	^{99}Es	^{100}Fm	^{101}Md	^{102}No	^{103}Lr

In the periodic table (Table 8.1 the elements are arranged in order of increasing atomic number, that is in order of increased nuclear charge, so that each element contains one more orbital electron than the preceding element. Instead of listing the 03 elements, the periodic table arranges them into several rows or periods, in such a way that each row begins with an alkali metal and ends with an inert gas. The sequence in which the various energy levels are flied determines the number of elements in each period, and the periodic table can be divided into four main regions according to whether the s, p, d or f levels are being filled.

1st period	ls			elements in this period	=	2
2nd period	2s		2p	elements in this period	=	8
3rd period	3s		3p	elements in this period	=	8
4th period	4s	3d	4p	elements in this period	=	18
5th period	5s	4d	5p	elements in this period	=	18
6th period	6s	5d	6p	elements in this period	=	32

This form of the periodic table has many advantages. The similarity of properties within a group and the relation between the group and the electron structure is emphasized. The d block elements are referred to as the transition elements since they are situated between the s and p blocks. Hydrogen and helium differ from the rest of the elements because there are no p orbitals in the first shell. Helium obviously belongs to Group 0, the inert gases, which are chemically inactive because their outer shell of electrons is full. Hydrogen is more difficult to place in a group. It could be included in Group I because it has one s electron in its outer shell, or in Group VII because it is one electron short of a complete shell. Hydrogen is included in both these groups in the periodic table, although it resembles neither the alkali metals nor the halogens very closely. The unique properties of hydrogen are largely due to the extremely small size of hydrogen atoms. Thus there is a case for placing hydrogen in a group on its own, or omitting it from the periodic table altogether.

3.3 The Periodic Table and Atomic Properties

We learned that physical and chemical properties of the elements vary periodically with the atomic number. The periodic table therefore provides an organized structure to the knowledge and understanding the chemistry of the elements. Apart from this there is also a variation of atomic properties of elements in the periodic table. Some of these properties are atomic and ionic sizes, ionization energy, electron affinity and electronegativity.

3.3.1 Atomic Size

The atom is a very small entity and so it is difficult to determine its size i.e. distance from the center of the nucleus to the outermost orbital. However with the aid of modern techniques such as X-ray and electron diffraction, it is possible to determine the distance between covalently bonded atoms. For example, the distance between the nuclei of oxygen atoms than oxygen molecule is 0.132nm, so die atomic radius of an oxygen atom is approximately 0.066nm. The atomic radius or sizes of any atom is taken to be one-half the distance of closest approach between the nuclei of atoms in the elemental substance.

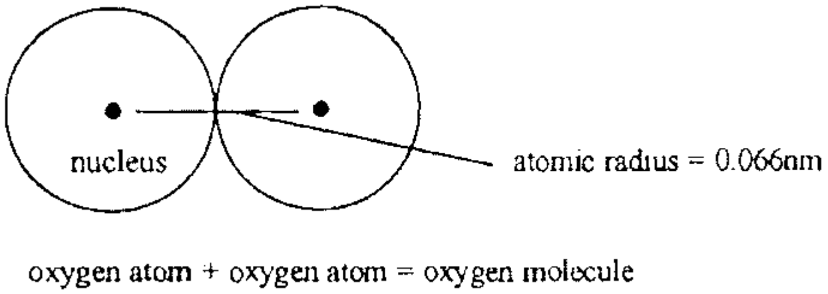
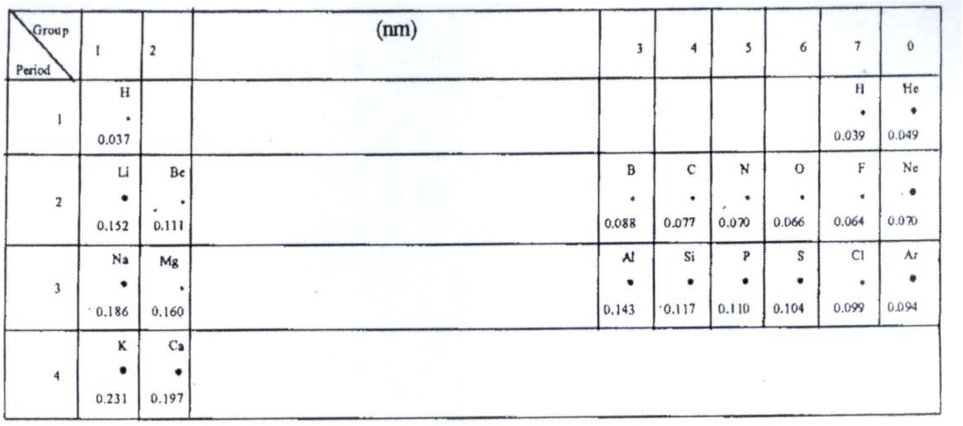


Fig 8.1 Atomic radius of oxygen

Atomic size decreases from left to right across a period in the periodic table. In other words, as the atomic number increases across any period, the size of the atom decreases. Recall that as we move across a period one electron is added increasingly from one element to the next and the electrons are being added to the same shell at about the same distance from the nucleus. At the same time, protons are also being added to the nucleus. Increase in the number of proton, increases the nuclear charge which progressively exert a stronger attraction upon the electrons around it and would pull them towards the nuclei. As the nuclear charge increases with atomic number across a period, the attractive force exerted by the nucleus on the outermost electrons of the atom increases hence the atomic radius or size decreases across a period. For example, on moving from Lithium to beryllium, the number of charges on the nucleus is increased by one, so that all the orbital electrons are pulled in closer to the nucleus. In a given period, the alkali, metal is the largest atom and the halogen the smallest. Table 8.2 shows the atomic radii (in nm) of the first twenty elements of the Periodic Table.



Group / Period	1	2	(nm)	3	4	5	6	7	0
1	H ● 0.037							H ● 0.039	He ● 0.049
2	Li ● 0.152	Be ● 0.111		B ● 0.088	C ● 0.077	N ● 0.070	O ● 0.066	F ● 0.064	Ne ● 0.070
3	Na ● 0.186	Mg ● 0.160		Al ● 0.143	Si ● 0.117	P ● 0.110	S ● 0.104	Cl ● 0.099	Ar ● 0.094
4	K ● 0.231	Ca ● 0.197							

On descending a group in the periodic table such as lithium, sodium, potassium etc, the atomic size increases due to the effect of extra shells of electrons being added; this outweighs the effect of increased nuclear charge. Recall the lithium in period 2 has two shells; sodium in period 3 has three shells while potassium in period 4 has four shells. In genera], as we go down the group, atomic size increases with atomic number (see Table 3.2).

3.3.2 Ionic Radius

Ions of elements are formed when atoms lose or gain electron. The size of an ion called ionic radii is different from atomic sizes. Ionic sizes are measured in electrovalent compounds.

The ionic radius of a given compound is the distance between the centre of one ton and the centre of its nearest neighbour of opposite charge.

A positive ion is formed by removing one or more electrons from an atom. When this happens, the number of positive nuclear charge is more than the number of negative electronic charge, hence the electrons are pulled in. A positive ion is therefore smaller than the corresponding atom and the more electrons removed '(that is, the greater the charge on the ion), the smaller it becomes e.g.

Atomic radius Na = 0.186nm Atomic radius Fe — 0.139nm
Ionic radius Na' = 0.1 lm Ionic radius Fe$^{2'}$ 0.090nm
Ionic radius Fe$^{3'}$ — 0.076nm

When a negative ion is formed, one or more electrons are added to an atom. The number of positive nuclear charge is now less than the number of negative electronic charge hence the pull on the electrons is reduced. In general, ionic radii of negative ions are greater than the corresponding atomic radii i.e. negative ions are bigger than the corresponding atom e.g.

Atomic radius Cl = 0.099
ionic radius Cr = 0.128

3.3.3 Ionization Energy

If energy is supplied to an atom, electrons may be promoted to a higher energy level. If sufficient energy is supplied, the electron may be completely removed, giving a positive ion. The energy required to remove the most loosely bound electron from each atom in a mole of gaseous atom, producing one mole of gaseous ion is called ionization energy (KJ mol^{1}). Since it is possible to remove one, two or three.... electrons from most atoms there is a first second or third.., ionization energy (I.E.).

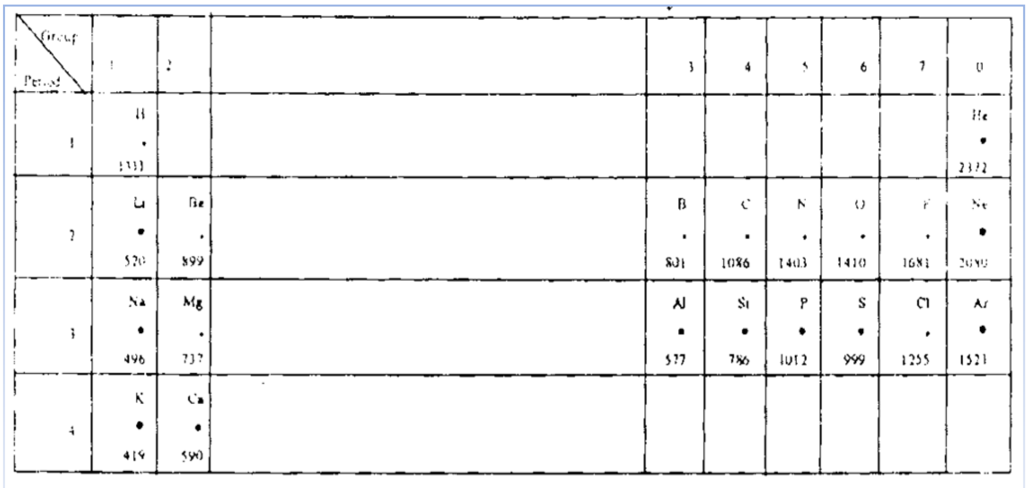
$$M \xrightarrow{1^{st}\ \text{l.E}} M^{+} \xrightarrow{2^{nd}\ \text{l.E}} M^{2+} \xrightarrow{3^{rd}\ \text{l.E}} M^{3+} \longrightarrow$$

The factors influencing the ionization energy are the:

a. size of the atom which is a measure of the outermost electron from the nucleus.
b. charge on the nucleus - as nuclear charge increases, the attraction for outermost electron increases, and so more energy is required to ionise.
c. screening effect of inner electron shells - the outermost electrons are repelled by all the other inner shell electrons and prevented from experiencing the full attraction of the positively charged nucleus.

Across a period, the first ionization energy increases as atomic number increases since the atomic radius decreases. As the distance decreases, the attraction of the positive nucleus for the electron will increase, hence more energy is required to remove the outermost electron hence the ionization energy will increase. Note that the screening effect remain almost the same across a period since electrons are added to the same shell. Table 3 shows the first ionization energies of the first twenty elements.

Table 3.3 First Ionization Energies of the Elements (KJ mot^1)
Table 3.2: Atomic radii of the first twenty elements



Group / Period	1	2		3	4	5	6	7	0
1	H • 1311								He • 2372
2	Li • 520	Be • 899		B • 801	C • 1086	N • 1403	O • 1410	F • 1681	Ne • 2080
3	Na • 496	Mg • 737		Al • 577	Si • 786	P • 1012	S • 999	Cl • 1255	Ar • 1521
4	K • 419	Ca • 590							

The first ionization energies of the elements in the first two short periods are shown in Fig. 3.2. These show a general upward trend from Li to Ne and from Na to Ar. The values for Ne and Ar are the highest in their periods because it requires a great deal of energy to break a stable filled shell of electrons. There are several irregularities. The high values for Be and Mg are attributed to the stability of a filled s level. The high values for N and P indicate that a half-filled p level is also particularly stable. The values for B and /A// are lower because removal of one electron leaves a stable filled s shell, and similarly with 0 and S a stable half-filled p shell is left.

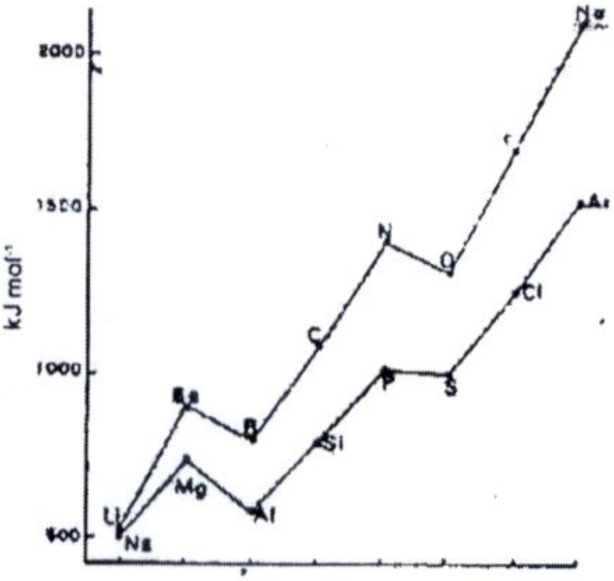
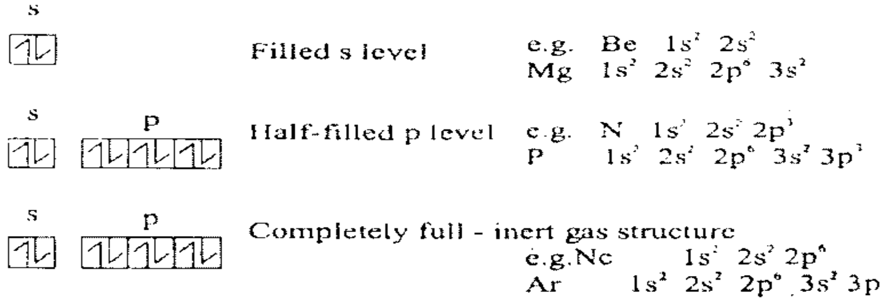


Fig. 3.2 First ionization energies of the elements in the first two short periods.



In general, the ionization energy decreases on descending a group because an extra shell of electron is added which also help to screen the outer election from the nucleus. This trend is shown with the alkali, metals from Li(520) to Na(496) to K(419).

3.3.4 Electron Affinity

Electron affinity has to do with the gain of electrons to fain negative ions while ionization deals with loss of electrons to form positive ions. The energy released when an extra electron is added to a neutral gaseous atom to form a univalent negative ton is termed the electron affinity. Since energy is given off in the process, electron affinity has a negative value. Electron affinities depend on the size and effective nuclear charge of the atom. Moving from left to right across a period, electron affinities decreases (i.e. increase in negative value), expect for the noble gases which have positive values. Down a group of the periodic table, electron affinities increase (i.e. decrease in negative value). The reason for the observed trend is that atoms with smaller atomic radii tend to have a stronger attraction for electrons and thus form negative ions more readily.

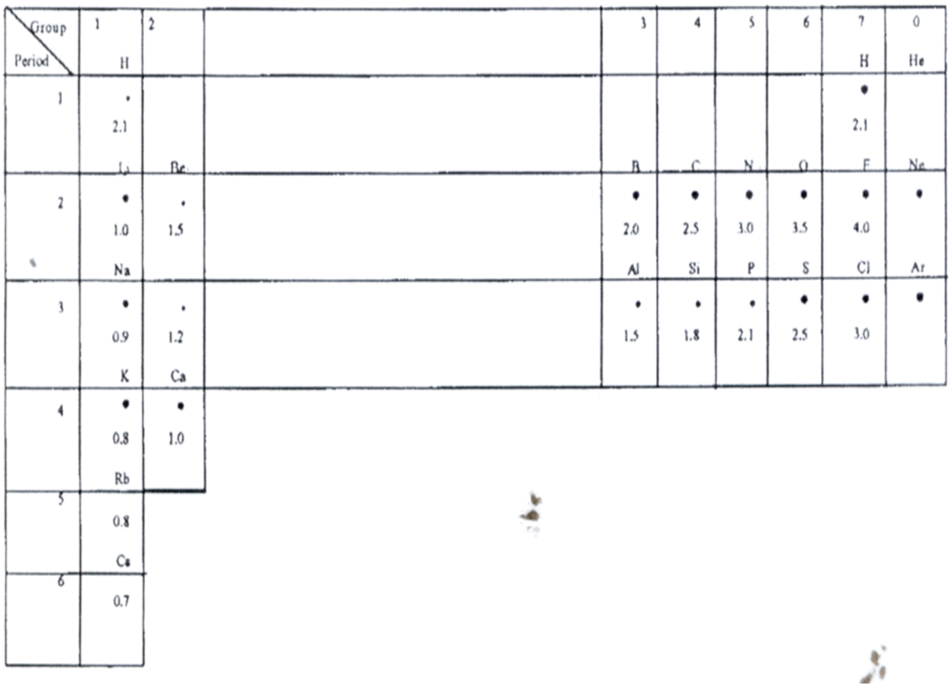
Element			**Energy (kJ mot^{-1} negative value)**
Li	⟶	Li^-	57
Na	⟶	Na^-	21
Al	⟶	Al^-	26
P	⟶	P^-	60
S	⟶	S^-	200
Cl	⟶	Cl^-	348

3.3.5 Electronegativity

The tendency of an atom in a molecule to attract bonded electrons to itself is termed the electronegativity of the atom. Generally, small atoms attract electrons (due to closeness of the nucleus) more than large ones and are therefore more electronegative. Atoms with nearly filled shells of electrons will tend to have higher electronegativity (because of the desire to have a stable filled shell) than those with sparsely occupied shells.

The electronegativites of elements decrease down a group and increase across a period. The reason for the trend is that down the group, atomic size increases and effective nuclear charge decreases hence electron attracting power (electronegativity) of the atom decreases. From left to right of a period, the opposite effect is observed, atomic size decreases and effective nuclear charge increases, these combine to increase electronegativity. The most electronegative elements are the reactive non-metals (e.g. Fluorine) at the top right-hand corner of the periodic table while the least electronegative elements are the reactive metals (e.g. caesium) at the bottom left-hand comer of the period table. (See Table 3.4)

Table 3.4 Electronegativity volume



Group Period	1 H	2		3	4	5	6	7 H	0 He
1	• 2.1 Li	Be		B	C	N	O	• 2.1 F	Ne
2	• 1.0 Na	• 1.5		• 2.0 Al	• 2.5 Si	• 3.0 P	• 3.5 S	• 4.0 Cl	• Ar
3	• 0.9 K	• 1.2 Ca		• 1.5	• 1.8	• 2.1	• 2.5	• 3.0	•
4	• 0.8 Rb	• 1.0							
5	0.8 Cs								
6	0.7								

4.0 CONCLUSION

In this unit you have learned a number of periodic properties that is evident when elements are arranged as a periodic function of atomic number. You should have also observed how this periodic properties vary down a group and across the period of a Periodic Table. You need to be aware of the reasoning behind S observed tend.

5.0 SUMMARY

What you have learned in the unit concerns the classification of elements in the Periodic Taft based on electronic configuration of atoms. It has saved to introduce you to the variation of atomic properties - atomic size and radius, ionization energy, electron affinity -and electronegativity, in the Palo& Table.

6.0 TUTOR-MARKED ASSIGNMENTS

l.(a) Use the electronic configuration to explain the occurrence of periods and groups in the Periodic Table. (b) Define ionization energy and electronegativity and explain their trends in the Periodic Table.

7.0 REFERENCES/FURTHER READING

I (Osei Yaw Ababio (2002), New School Chemistry Onitsha. Africana-Fep Publishers.

UNIT 4 MOLE CONCEPT 1

CONTENTS

1.0 INTRODUCTION

Though it is possible to compare the masses of atoms, they are too small to be weighed. A chemical reaction takes place always between large numbers of reactant particles. The products that are formed also contain a lame number of product particles. Chemists therefore use a large number of particles as a base unit when comparing amounts of different substances reacting or are formed in chemical reactions. This basic unit is the mole and the mole concept is one of the most important concepts in Chemistry. The mole concept is applicable to all chemical processes. In this unit the mole will be defined and the concept applied to chemical calculations involving masses and volumes of chemical substances.

2.0 OBJECTIVES

At the end of this unit, you should be able to

- define the mole.
- recall avogadro number.
- calculate molar mass tan formula mass.
- calculate percentage of elements in compounds with given formulae.
- establish mole ratios of reactants and products in reaction.

- calculate yield and percentage yield.
- recall molar volume of gas at s.t.p.
- use the above in calculations involving gases.

3.0 MAIN CONTENT

3.1 The Mole Defined

The mole is the amount of any substance that contains as many elementary particles as there are atoms in 0.012Icg of carbon-12. Now 0.012kg of carbon - 12 contains the Avogadro number of atoms, so one mole of a substance is that amount which contains Avogadro number of particles. The elementary panicles may be molecules, atoms, ions, electron etc and must be specified. Avogadro number has been determined experimentally and is 6.02 x 10 23. A very large number indeed. A mole of a substance therefore provides a quantity of material that can be measured for use in the laboratory.

1 mole ≡ 6.02×10^{23} particles

3.2 The Mole and the Molar Mass

The atomic mass of an element is the mass of one atom of the element relative to the mass of one atom of Carbon-12 which by convention is given a mass of 12.00 a.m.u.

6.02×10^{23} atoms of C-12 ≡ 12g

1 atom of C-12 ≡ $\frac{12}{6.02 \, x \, 10^{23}}$ g

take helium (relative atomic mass 4) as our first example

mass of 1 atom of He = $\frac{4}{12}$ x $\frac{12}{6.02 \, x \, 10^{23}}$ g

= 4g

Take oxygen (relative atomic mass 15.99)

mass of 1 atom of O = $\frac{15.999 \, x \, 12}{12 \, x \, 6.02 \, x \, 10^{23}}$ g

mass of 1 mole O atoms = $\frac{15.999 \, x \, 12}{12 \, x \, 6.02 \, x \, 10^{23}}$ x 10^{23} x 12 g

= 15.999g

It follows from the above calculation that the mole can also be defined as the quantity of a substance whose mass in grams is numerically equal to the atomic, molecular or formula mass of the substance. The molar mass of a compound is the number of grams of the compound needed to make up one mole of the compound i.e. contains 6.02 x 1023 molecules of the compound.

With this new definition of the mole you can calculate the number of

i. moles
ii. particles
iii. atoms etc in a given mass of a substance of known formula.

Table 9.1 Some substances and their molar masses

Substance	Formular	Formular mass	Molar mass (g)
Hydrogen gas	H_2	2.0	2.0 g
Magnesium	Mg	24.0	24.0g
Methane	CH_4	18.0	18.0g
Water	H^20	18.0	18.0g
Common salt	NaCI	58.5	58.5g

Example 1

Calculate the following for 4.5g of water

i. number of moles.
ii. number of water molecules.
iii. mole and number of oxygen atoms.
iv. mole and number of hydrogen atoms.

Solution

i. Formula of water is H_2O
formula mass = 2 (1) + 16 = 18g
∴ molar mass of water = 18g
18g of water = l mole

$\therefore$ 4.5g of water = $\frac{1}{18}$ x 4.5mol

= 0.25 mol

ii. 1 mole ≡ 6.02×10^{23} particles
∴ 0.25 mole ≡ $6.02 \times 10^{23} \times 0.25$ molecules of water
= 1.505×10^{23} molecules.

iii. Formula of water = H_2O

mole of O = mole of water = 0.25
number of O atoms = $0.25 \times 6.02 \times 10^{23}$
$= 1.504 \times 10^{23}$ atoms

iv. From the formula mass H_2O
mole H atoms = 2 x mole of water = 0.50 mol
Number of H atoms = $0.50 \times 6.02 \times 10^{23}$
$= 3.01 \times 10^{23}$ atoms

You can also convert amount (in moles) to mass (m gram).

Example 2

What is the mass of 2 moles of ammonia gas (NH_3) with the following relative atomic mass of atoms, N = 14 H =1

Solution

formula mass of NH_3 = 14 + 3 (1)
= 17
molar mass = 17g
1mole has a mass of 17g
∴ 2 moles = 17g x 2 = 34g

3.3 Finding the Formula of a Compound

This involves the calculation of the quantities of the elements involved in the formation of the compounds. The concept of a mole is central in this type of calculations.

Example 1

0.24g Magnesium is burnt in oxygen. The mass of the magnesium oxide is found to be 0.40g
Determine (i) the formula of magnesium oxide
(ii) the molar mass of magnesium oxide.

Solution

Mass of oxygen in magnesium oxide
= 0.40 - 0.24 = 0.16g,
0.24g Mg combines with 0.16g O
OR (multiply x 100 to remove the decimal point)
24g Mg combines with 16g O to form
40g of magnesium oxide.

mole of magnesium = $\frac{24}{24} = 1$

mole of Oxygen = $\frac{16}{16} = 1$

∴ 1 mole of magnesium combines with 1 mole of oxygen in magnesium oxide

∴ Formula = MgO
molar mass = 24 + 16 = 40g.

Example 2

Zinc oxide is found by chemical analysis to contain 80.3 percent (by mass) of zinc and 19.7 percent of oxygen.
Determine the formula of zinc oxide.

Solution

Assuming we analyse 100g sample.
80.3g Zn and 19.7g O

mole of Zn = $\frac{80.3}{65.4} = 1.23$
mole O = $\frac{19.7}{16.00} = 1.23$

1.23 mol of zn combines with 1.23 mot of oxygen
1 mol of Zn combines with 1 mol of oxygen
∴ formula = ZnO.

Example 3

2.00g mercury combines with 0.71 chlorine gas to form 2.7g of mercury chloride. What is the simplest formula of mercury oxide.

Solution

2.00g Hg combines with 0.71g chlorine

mole Hg = $\frac{2.00}{200} = 0.01$

mole Cl = $\frac{0.71}{35.5} = 0.2$

.01mol Hg combines with .02 mol Cl

1 mol Hg will combine with $\frac{1.0 \text{ x } 0.02}{0.01} = 2$

formula = $HgCl_2$

3.4 Calculating the Percentages of Elements in a Compound

Example 1

Calculate the percentage of oxygen in water.

Solution

formula of water H_2O

formula mass = 2 + 16 = 18g

mass of 1 mole = 18g.

% oxygen = $\frac{16}{18}$ x 100 = 88.9

Example 2

Calculate the percentage of nitrogen in ammonium nitrate fertilizer formula NH_4NO_2

Solution

NH_4NO_2 formula mass = 14 + 4 (1) + 14 + (16) 3
= 14+4+14+48
= 80
molar mass = 80 g
mass of nitrogen/formula mass = 28

% Nitrogen = $\frac{28}{80}$ x100

= 35 percent

3.5 Using Empirical Formula to Calculate Unknown Atomic Man of an Element

Example 1

The empirical formula of non metallic oxide is $X0_2$ where X = stands for the chemical symbol of the element with unknown atomic mass. 0.80g of X is burnt in oxygen and the mass of the metal oxide found to be 1.60g. Given that the atomic mass of oxygen is 16, determine the atomic mass of X.

Solution

Mass of oxygen in the compound = 1.60 - 0.80
= 0.80g

mole of oxygen $\frac{0.80}{16}$ = 0.05

from the formula XO_2
ratio X: O is 1:2
2 mol O combine with I mole x
0.05 mol O will combine with 0.05 x ½
= 0.025 mol X
0.025 mol X ≡ 0.8g

1 mo1 X = $\frac{0.8}{0.025}$ = 32g

molar mass is 32

3.6 The Mole in Chemical Reactions (Yield and Percentage Yield)

The above is very important in chemistry because it allows for calculation of yields expected even before an experiment is carried out. A balanced chemical equation of the reaction is all that is required.

$$FeO_{(s)} + CO_{(g)} \longrightarrow Fe_{(s)} + CO_{2(g)}$$

mole ratio of reactant	of product
1 : 1	1 : 1

1 mole FeO + 1 mole CO $\longrightarrow$ 1 mole $Fe_{(s)}$ + 1 mole CO_{2g} or in terms of masses.

72g FeO + 28g CO $\longrightarrow$ 56g $Fe_{(s)}$ + 44g CO_2

Suppose we require to produce 2.8g $Fe_{(s)}$

mole Fe(s) required = $\frac{2.8}{56}$ = 0.05

mole FeO needed as reactant
= 0.05 x 72 = 3.6g
mole CO = 0.05 x 28 = 1.4g

3.6.1 Calculation of Percentage Yield

In most reactions the conversion of reactant to product is not often complete. The percentage yield gives the ratio of the experimental yield to a theoretical yield assuming complete reaction.

Take the last example

Suppose 3.6g of FeO reacts with enough
CO(g), and only 2.2g of Fe(s) is formed.

$$\%\text{yield} = \frac{\text{yield from experiment}}{\text{yield expected}} \times 100$$

$$= \frac{2.2}{2.8} \times 100$$

= 78.571%

3.6.2 Molar Volume for Gases

At a temperature of 273 and pressure of I atmosphere commonly called (STP), a mole of any gas occupies a volume of 22.4 dm^3. This is called molar volume.

1 mole of a gas occupies a volume of 22.4de at S.T.P.
Now consider the reaction in the last section.

$FeO_{(g)} + CO \quad Fe_{(g)} + CO_{2(g)}$
At S.T.P.
22.44dm^3 of $CO_{(g)}$ = 1 mole $CO_{(g)}$
22 4 dm^3 $CO_{2(g)}$ = 1 mole CO_{2g}

Example 1

Calculate the volume of CO required at S.T.P. to react completely with 3.6g FeO.

$$FeO_{(g)} + CO_{(g)} \longrightarrow Fe_{(g)} + CO_{2(g)}$$

mole ratio FeO : CO = 1:1

Mole FeO in 3.6g FeO $\frac{3.6}{72}$ = 0.05

mole CO needed 0.05

Volume CO needed at S.T.P
= 0.05 x 22.4 dm^3
= 0.ll2dm^3

Example 2

What mass of Fe(s) is formed by complete reaction of 5.6dm^3 of CO at ST.P. in the above reaction.

Solution

mole CO = $\frac{5.6}{22.4}$ = 025 mole

mole ratio
FeO: CO: Fe(s) = 1 : 1 : 1
∴ mole Fe formed = 0.25
mass of Fe = 0.25 x 56
= 1.4g

4.0 CONCLUSION

From the many examples considered in this unit, you will agree that the mole concept is a very useful tool in chemical calculations. In the next unit the use of mole concept in volumetric analysis and solution preparations will he discussed. Learn to use the mole concept and you will be in a position to solve a mole of problems.

5.0 SUMMARY

The mole is defined and applied to chemical calculations. The mole concept is applicable to gas reactions as well as reactions with solid and liquid substances. The mole is a measurable quantity of substance and is more relevant to experiments in quantitative analysis.

6.0 TUTOR-MARKED ASSIGNMENTS

1. (a). Give the number of moles of atoms in the following:
   60g Carbon, 6.4g 0, and 10.00g Ca
   (b) What is the mass of the following?
      (i) 3.5 moles of Oxygen molecules
      (ii) 0.05 mole of CaC_2
2. (a) What is percentage of water of crystallisation in $CuSO_4$ $5H_2O$ and Na_2CO_3 $10H_2O$
   (b) 5.4g of carbon burns in excess oxygen Calculate
      (i) mole and mass of CO_2

(ii) formed volume of CO_2 at S.T.P.

3. The percentage composition by mass of a compound of sodium, sulphur and oxygen are 18.2, 12.7 and (9.1 percent respectively. Water of crystallisation is 50 percent (by mass). Determine the formula of the compound.

7.0 REFERENCES/FURTHER READING

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UNIT 5 ACIDS, BASES AND SALTS: GENERAL PROPERTIES

CONTENTS

1.0 INTRODUCTION

The concept of acids and bases is one of the most important and useful concepts in chemistry. Many chemical reactions are classified as acid-base reactions on the basis of the many theories that have been proposed.

Acids, bases and salts are compounds that we see, touch and use in our everyday life at home. We use comma' salt (sodium chloride) everyday in seasoning our food and sterilising fruits and vegetables. We use soap everyday in washing. Many of air food items contain these compounds. Citrus fruits like oranges and lemon contain citric acid and lactic acid is present in milk. Sodium chloride, tetraoxosulphate (vi) acid and calcium trioxocarbonate (iv) and many others are important raw materials in the chemical industry.

Water is one of the most abundant chemical in nature and saves as a good solvent for many substances bat ionic and covalent. Most acids dissolve in water to produce ions. This is also true of bases and salts. They are electrolytes. In this unit, we investigate the properties of acids, bases and salts and explain some of their behaviour in chemical reactions.

2.0 OBJECTIVES

At the end of this unit you should be able to:

- define acid, base and salt.
- describe some methods of preparation of each.
- list some properties and uses of each.
- explain hygroscope, deliquescence and efflorescence.
- differentiate between a weak and a strong acid.
- explain, why acid/base solutions conduct electrically.

3.0 MAIN CONTENT

3.1 Acids

3.1.1 Definition, Examples and Preparation

An acid can be defined as a compound,

i. containing hydrogen which can be replaced directly or indirectly by a metal;

ii that produces $H^+(H_3O^+)$ (hydroxanum ion) as the only positive ion in aqueous solution.

$$H^+ + H_2O \longrightarrow H_3O^+$$

By the above definitions HNO^3, H_2SO_4 CH_3 COOH are acids. HNO_3 and H_2SO_4 are sometimes referred to as inorganic or mineral acids while CH_3 COOH is referred to as an organic acid — Now add more acids to the list of acids.

Acids are generally prepared by the following methods.

(a) Action of water on acid anhydride (acidic oxides) e.g.

$$H_2O + SO_2 \longrightarrow H_2\,SO_3 \text{ trioxosulphate (iv) acid}$$
$$H_2O + N_2O_5 \longrightarrow 2HNO_3$$

(b) Displacement of a more volatile acid from its salt.

$$2KCI + H_2\,SO_4 \longrightarrow K_2\,SO_4 + 2HCl$$

(c) Oxidation of petroleum hydrocarbons

$2C_3H_g + 4O_2 \longrightarrow 3CH_3COOH + 2H_2O$ ethanoic acid.

(d) Ethanoic acid can also be prepared by the reaction of lead (ii) ethanoate with hydrogen sulphide.

$Pb\ (CH_3C00)_2 + H_2S \longrightarrow PbS_{(g)} + 2CH_3COOH$

(e) Direct reaction of the elements

e.g. $H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCI_{(g)}$

The first reaction requires a catalyst. The catalyst used is activated charcoal.

3.1.2 General Properties of Acids

Acids generally have sour taste and are corrosive. An aqueous solution of an acid turns blue litmus paper red. The chemical properties include

(a) Reaction with metals to liberate hydrogen gas e.g.

$Zn + 2HCl_{(aq)}^{2} \longrightarrow ZnSO_4 + H_2$

(b) Liberate CO_2 from trioxocarbonates and hydrogentrioxocarbortates

$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$

(c) Neutralise bases or alkalis to form salt and water only.

$NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$

$Cad + 2HNO_3 \longrightarrow Ca(NO_3)_2 + H_2O$

3.1.3 Uses of Acids

(a) In manufacture of dyes, fertilizers, paints, detergents and salts.
(b) They serve as solvents and electrolytes.
(c) As preservatives.
(d) For refining gasoline and lubricating oils.
(e) In dissolving metals.

3.2 Bases

3.2.1 Definition Examples and Preparation

A base is defined as a compound which produces OH as the only negative ion in aqueous solution.

A more general definition is that a base neutralises an acid to form salt and water only. A soluble base is called an alkali.

By the above definition CaO, NaOH $Ca(OH)_2$ are bases.

Bases are prepared by the listed methods.

(a) Burning of metallic elements in oxygen

$$4Na + O_2 \longrightarrow 2Na_2O$$
$$2Mg + O_2 \longrightarrow 2MgO$$

(b) Decomposition of carbonates.

$$Ca\ CO_3 \xrightarrow{heat} CaO + CO_2$$
$$Cu\ CO_3 \longrightarrow CuO + CO_2$$

(c) Alkaline solutions are prepared by dissolving metallic oxides in water.

$$Na_2O + H_2O \longrightarrow NaOH(aq)$$
$$CaO + H2O \longrightarrow Ca(OH)_2(aq)$$

(d) Direct reaction of electropositive metals with water.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$
$$Ca + 2H_2O \longrightarrow Ca(OH)2 + H_2$$

3.2.2 General Properties of Bases

Bases have bitter taste and appear slippery or soapy to touch. Their solutions will turn red litmus blue. Bases liberate ammonia gas when warmed with an ammonium salt:

$$NH_4\ Cl + NaOH \longrightarrow NaCl + NH_3 + H_2O$$
$$(NH_4)_2\ SO_4 + 2KOH \longrightarrow K_2SO_4 + 2NH_3 + 2H_2O$$

3.2.3 Uses of Bases

(a) In manufacture of dyes, paper, soap etc.
(b) Some are used as drying agents.
(c) Calcium hydroxide is used for making mortar, plaster of paris (POP) and cement. It is also used in sugar refining and neutralisation of soil acidity.
(d) Milk of magnesia is a suspension of magnesium hydroxide in water. It is used as a mild laxative and antacid in tooth pastes to neutralise acidity in the mouth.

3.3 Salts

3.3.1 Definition, Examples, and Preparation

A salt is formed when a base is neutralised by an acid. A salt consists of a metallic and acidic radicals. Examples are NaCl, $NaNO_3$, Na_2SO_4, KCI etc.

Salts are prepared by the listed methods.

(a) Neutralisation of a base by an acid.

(b) Reaction of an acid with a reactive metal

$Zn + H_2SO_4 \longrightarrow Zn\ SO_4 + H_2$

(c) Reaction of an acid with a trioxocabonate(iv)

$Na_2CO_3 + 2HC1 \longrightarrow 2NaC1 + H_2O + CO_2$

(d) Reaction of soluble salts to produce an insoluble salt.

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

$Ba\ Cl_{2(aq)} + Na_2\ SO_{4aq} \longrightarrow BaSO_{4(s)}$

(e) Direct combination of a gas with metal

$2Na_{(s)} + Cl_2 \longrightarrow 2Na\ Cl$

$2Fe + 3Cl_2 \longrightarrow Fe\ Cl_3$

3.3.2 Types Salts

(a) Normal Salt

This is the salt famed when all the replaceable hydrogens in the acid have been completely replaced by a meal.

e.g. NaCI, Na_2SO_4, $Ca(NO_3)_2$

(b) Add Salts

These are formed when the replaceable hydrogens lathe acid are partially replaced by a metal. They are produced by partial neutralisation of acids with more than one replaceable hydrogens.

They are acidic to litmus e.g. NaHSO4 NaHCO3 is alkaline to litmus because of hydrolysis.

(c) Basic Salts

They are formed when there is partial replacement of the hydroxyl ions in a metallic hydroxide by an acid. They are formed as a resift of partial neutralisation of a base.

$Zn(OH)_2 + HCl \longrightarrow Zn\ OH\ Cl + H_2O$

$Ba(OH)_2 + HC1 \longrightarrow Ba\ OHCl + H_2O$

They are alkaline to litmus.

(d) Double Salts

Double salts are formed when two salts crystallise from solution in definite proportion by mass. e.g. $K_2SO_4\ Al_2\ (SO_4)_3 . 24H_2O$

from $K_2\ SO_4$ and $Al_2(SO_4)_3$. They combine the chemical properties of the two salts.

(e) Complex Salts

These are salts of complex ions.
e.g. K_3 Fe $(CN)_6$, $Cu(NH_3)_4$ SO_4

Recall that complex ions are formed by coordinate covalent bonding of metal ions with covalent molecules that can donate lone pair of electrons for binding. These covalent molecules are called Uganda

(f) Hydrated Salts

These are salts that contain water of crystallisation e.g. Na_2CO_3. lOH_2O, $CuSO_4$, $5H_2O$. The water of crystallisation is chemically bonded to the salt.

On heating or on exposure, some hydrated salts lose their water of crystallisation.

3.4 General Characteristics of Acids, Bases and Salts

3.4.1 Ionisation

Acids, bases and salts ionise in aqueous solution to produce ions. They are all electrolytes. Acids and bases which are completely ionised in solution usually water are strong acids a bases e.g. HNO_3, HCl, NaOH, KOH and H_2 SO_4. Weak acids and bases are partially ionised in solution e.g. NH_4, OH. The extent of ionisation is given by the acid or base dissociation constant which has been discussed in an earlier unit (i.e. Unit 6).

Strong acids and bases are strong electrolytes and weak acids and bases are weak electrolytes. Very soluble salts ionise readily in solution and are strong electrolytes but sparingly soluble salts are weak electrolytes. Because acids ionise and increase the H' concentration in water, the pH of acid solutions is less than 7. Bases that dissociate in water increase the hydroxide ion concentration in water. The pH of base solutions is greater than 7. Salts solutions are generally neutral to litmus. Some however show acidic or basic behaviour because of hydrolysis. Recall that hydrolysis is ran of an earlier unit i.e. Unit 6.

3.4.2 Deliquescence

This is a phenomenon whereby substances absorb moisture and form solution. Substances that show this behaviour are called deliquescent compounds.

e.g. Na OH, KOH, NaCl, $MgCl_2$ and $CaCl_2$.

They have great affinity for water and are sometimes used as drying agents, e.g. fused calcium II chloride.

3.4.3 Hygroscopy

Hygroscopic compounds absorb moisture without forming solution when exposed to the atmosphere. They have very strong affinity for water and are used as drying agents.

e.g. CuO, CaO, and concentrated $H_2 SO_4$. They are sometimes called dessiccants.

3.4.4 Efflorescence

Efflorescence is a phenomenon whereby a hydrated salt loses all or part of its water of crystallisation when exposed to the atmosphere. An example is $Na_2 CO_3 . 10H_2O$. This salt loses nine molecules of water of crystallisation when exposed and is converted to the mono hydrate

$$Na_2 CO_3 .10H_2O \longrightarrow Na_2 CO_3 H_2O + 9H_2O$$

4.0 CONCLUSION

The acid-base concept used in this unit is based on ease of hydrogen ion or hydroxyl ion loss in solution which is the Arrhenius concept. With this concept, so many chemical compounds are classified as acids, bases and salts.

There are two more acid base theories not discussed here which when considered will allow many more compounds to be classified as acids, bases or salts. It is hoped that after studying this unit you now understand and appreciate these compounds more.

5.0 SUMMARY

- Acids, bases and salts are defined, examples are given and the methods of preparation discussed. Their uses are also enumerated.
- Some general characteristics are discussed in relation to their uses as electrolytes and drying agents.

6.0 TUTOR-MARKED ASSIGNMENTS

1. You are supplied with $0.01 mol dm^{-3}$ solutions of A, B, C, D, and E with pH values of 2, 12, 3.5, 6.5 and 9 respectively.

   a. Arrange the solutions in order of increasing +1* concentration.
   b. Which of the solutions will turn litmus from red to blue?
   c. Which of them is likely to be a solution of a strong acid? Explain.
   (d) Explain what happens when equal volumes of A and B are mixed. What is the pH of the mixture? 2

2. (a) Give two characteristics reaction each of an acid and a base
   (b) Give the product of the given reactions.
      (i) $CuCO_3 \longrightarrow$
      (ii) $CuCO_{3(s)} + HCl_{(aq)} \longrightarrow$
      (iii) $CH_3COOH + Ca(OH)_2 \longrightarrow$

7.0 REFERENCES/FURTHER READING

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MODULE 3

UNIT 1 CARBON AND ITS COMPOUNDS

CONTENTS

1.0 INTRODUCTION

Carbon is a non-metal that has been known for a very long time under the names charcoal, soot and diamond. It is the sixth most abundant element in the universe. Carbon exists naturally in the elemental states as diamond, graphite and coal; and in the combined state as petroleum, wood, natural gases and mineral deposits. The common mineral deposits of carbon are the metallic trioxocarbonates (IV) such as limestone (calcium trioxocarbonates (IV) and dolomite (Magnesium trioxocarbonates (IV)). Carbon as carbon (IV) oxide, CO:, constitutes

0.03% of the atmosphere. The carbon found in coal, petroleum, wood and natural gases are used essentially as source of fuels for energy supply.

Carbon is also an essential constituent of all living things. The study of the millions of carbon compound in living (organic) organism forms a separate branch of chemistry known as 'organic chemistry'. Organic compounds will be treated in separate units of this programme. Our concern in this unit is the chemistry of carbon as an element and its non-organic compound such as the oxides, sulphides, carbides, trioxocarbonata (IV) and hydrogen trioxocarbonates (IV) which are in the domain of in organic chemistry.

You will recall that carbon has six electrons with an electronic configuration of ls^2 $2s^2$ $2p^4$ and is therefore placed in group 4 and period 2 of the periodic Table. The energy required to lose its four outermost electrons or gain four electrons to form a stable configuration is enormous, therefore carbon combines essentially by forming four covalent bonds with other elements, Chemically, carbon is not very reactive element and that is why it exist naturally as elemental carbon in various forms e.g. diamond, graphite, coal, charcoal and carbon black.

2.0 OBJECTIVES

At the end of this unit, you should be able to:

- recognise the uniqueness of carbon as an element.
- define allotropy and describe the various allotropic forms of carbon.
- enumerate the important inorganic compounds of carbon.
- describe the characteristics of some carbon compounds.
- explain the carbon cycle,

3.0 MAIN CONTENT

3.1 Allotropes of Carbon

The existence in the same physical state of two or more structural form of an element is known as allotropy. Carbon exhibit allotropy. Diamond and graphite are two allotropic forms of crystalline carbon. The others like coal, coke, charcoal, lamp-black, sugar charcoal and animal charcoal are amorphous or non-crystalline forms of carbon

3.1.1 Diamond

Diamond exists naturally in South Africa, Brazil, India, and Venezuala. It is the purest form of naturally occurring carbon and is commonly obtained as a colourless solid, which can be transformed into brilliant gem. The structures of diamond consist of an infinite way of carbon atoms covalently bonded. The diamond crystal is octahedral in shape.

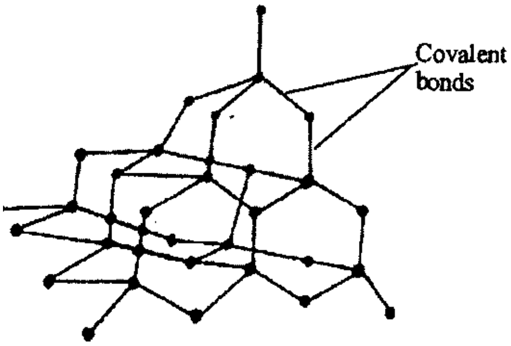

Fig. 3.1 The structure of Diamond

The network structure gives diamond its great strength and high melting temperature. Diamond is the hardest substance known. It is a poor conductor of heat and electricity. Diamond is resistant of chemical attack. Since diamonds are dense and hard, they are used industrially in drills for mining, as abrasives sharpen very hard tools, and for cutting glass and metals. They are also used as pivot supports in precious instruments and as dies for drawing wires. Its high refractive index and dispersion power give it a sparkling brilliance when it is cut and polished making it valuable as jewellery.

3.1.2 Graphite

The structure of graphite consists of hexagonal sheets of covalently bonded carbon atoms. The sheets are held together by weak intermolecular forces in parallel layers. Graphite occurs naturally as an opaque black solid called plumbago, which is formed by the action of volcanic heat on coal deposits. It is found as deposit in Austria, China, West Germany, Madagascar and Srilanka.

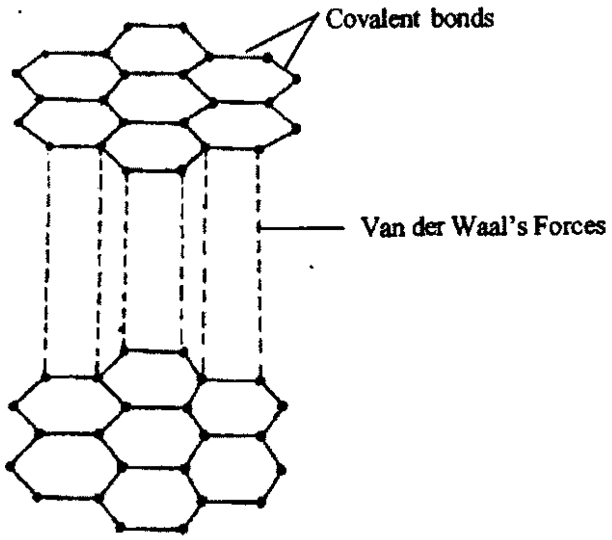

Fig 3.2 The Structure of Graphite

Graphite can also be artificially produced by heating coke at a high temperature in an electric furnace.

Because graphite is in sheets, it is soft and slippery. Although it has a high melting point, it is less dense than diamond. Graphite is relatively inert chemically but can be oxidized under suitable conditions. Unlike diamond, graphite is a good conductor of electricity because of the presence of mobile electrons within the sheets. The electrons exist since only three of the four valence electrons of each carbon atom in the graphite crystal are involved in bond formation.

Graphite is an excellent dry lubricant. This is because its layered structure allows one layer to slide over another easily. Unlike oil, it is non volatile and not sticky. It is usually used on bicycle chains and for the bearings of some motor cars. When mixed with oil it forms a high temperature lubricant. Being a good conductor of electricity and is often used as electrodes in electroplating and batteries. A non-conductor can be made conductive by coating it with graphite. Because graphite is soft and marks paper, it used to make pencils. It is used as black pigment in paint.

3.1.3 Amorphous (Or Non-Crystalline) Carbon

Carbon exists in several other forms which have no definite crystalline structure. They are not considered as true allotropes of carbon. Apart from coal, which is naturally occurring, all the other amorphous forms of carbon can be prepared in various ways.

a. Coal was formed from vegetation protected from complete decay by water-washed earth deposit. Decomposition occurred gradually under pressure and in the absence of air. Carbon (IV) oxide, methane and water were liberated, leaving behind a material that contained a very high percentage of carbon. During this process of carbonization, the vegetable material was converted in stages into peat, lignite (or brown coal), bituminous (or soft) coal, and finally anthracite (or hard coal) which is about 95% pure carbon. Impurities present may include nitrogen, sulphur and phosphorus. Coal is used mainly as a fuel to generate power for steam engines, factories and electrical plants. It is also used for making various chemicals e.g. benzene, methane. Nigeria has a large coal deposit in Jos, Plateau State and Enueu, Anambra State. There are large lignite deposits in °nestle and Asaba.

b. Coke is obtained by heating bituminous coal to very high temperatures in the absence of air to burn off the volatile constituents. This process is referred to as the destructive distillation of coal. Coke is used mainly as a fuel as it burns with practically no smoke and leaves very little residue. It is an important industrial reducing agent, used in the extraction of metals, e.g. iron from their ores. It is also used in the production of gaseous fuels, like water gas and producer gas and for the manufacture of graphite, calcium carbide, silicon carbide and carbon (1V) sulphide.

c. Carbon black (soot) is made by heating carbon containing materials in a limited supply of air to give finely divided particles of carbon. Lamp-black is obtained front vegetable of lamp oils, while carbon black itself obtained from coal gas, natural gas or fuel oils. Carbon black is used in manufacturing tuber tyres, black shoe polish, printer's ink, typewriting ribbons and carbon paper.

d. Charcoal can be made by burning wood, nut shells bones, sugar and even blood. Wood charcoal, the most common, is prepared by burning wood in a limited supply of air.
Sugar charcoal is formed when sugar is dehydrated i.e. `the hydrogen and oxygen it contains are removed in the form of water either by burning the sugar in a limited supply of air or by the action of concentrated tetraoxosulphate(V1) acid. It is the purest form of amorphous carbon.

Animal charcoal is produced when bones and animal refuse are burnt in limited supply of air. Charcoal has a very porous structure and allows molecules of gases and dyes to adsorb or adhere to its internal surfaces. Thus, it is a good adsorbent particularly when activated. Wood charcoal is used in gas-masks for adsorbing poisonous gases. It is also used for the purification of the noble gases and the recovery of industrial solvents. Similarly, animal charcoal, is used in removing the brown colour front crude sugar and colourizing petroleum jelly. Like coal, wood charcoal is also used mainly as a domestic fuel in Nigeria.

3.2 General Properties of Carbon

All the different allotropes of carbon are black solids or greyish-black except diamond which is colourless when pure.

All the allotropes have high melting point, about 3,500°C.
Carbon is insoluble in all common solvents e.g. water, petrol, alkalis, acids and carbon (IV) sulphide. This characteristic explains why carbon deposits in motor engines have to be mechanically removed decarbonization of motor engines.

3.3 Chemical Properties of Carbon

Carbon is not a very reactive element. Its compounds are formed by covalent bonding and most of them are stable. Carbon forms single or multiple covalent bonds with itself and other elements such as hydrogen, oxygen, nitrogen and sulphur. The unique ability of carbon to form long chains and rings of carbon-carbon bonds (known as catenation) enables it to have numerous compounds. The molecules range from small to extremely large ones and most of them are organic in nature and therefore beyond the scope of this unit.

All the allotropes of carbon have similar chemical properties; however diamond and graphite are less reactive than the amorphous form.

3.3.1 Combustion Reactions

Carbon burns in excess oxygen to produce carbon (IV) oxide only, although the temperature required differ for the allotropic forms.

$C(s) + O_2(g) \longrightarrow CO_2(g)$
When the supply of air is limited, combustion may not be complete. Carbon (II) oxide is formed instead of carbon (IV) oxide

$2C(s) + 0_2(g) \longrightarrow CO(g)$

Charcoal Fires: In countries like Nigeria and Ghana, charcoal is used extensively for making fires. As the charcoal burns, carbon (IV) oxide and carbon (II) oxide are formed at various levels inside the charcoal pot, depending on the level of oxygen supply.

Some of the reactions involved are:

$2CO + O_2 \longrightarrow CO_2$
$CO_2 + C \longrightarrow 2CO$
$C \quad + \longrightarrow CO_2$

The carbon (II) oxides that are not oxidized and escapes to the atmosphere is dangerous if the area around the fire is not well ventilated because the gas is poisonous.

3.3.2 Combination Reactions

Carbon combines directly with certain elements such as sulphur, hydrogen, calcium and aluminium at very high temperatures usually greater than 500°C.

$C(s) + S(s) \longrightarrow CS_2(1)$ carbon (IV) sulphide
$C(s) + 2H_2(g) \longrightarrow CH_4(g)$ methane
$2C(s) + Ca(g) \longrightarrow CaC_2$ (s) Calcium Carbide
$3C(s) + 4A1(s) \longrightarrow A1_4C_3$ (g) aluminum carbide

3.3.3 Carbon as Reducing Agent

Carbon is a strong reducing agent i.e. as oxygen acceptor. It reduces the oxides of less active metals such as iron and copper; at high temperatures.

$Fe_2O_3 + 3C \longrightarrow 2Fe + 3CO$
$2CuO + C \longrightarrow 2Cu + CO_2$

At high temperatures, carbon can also reduce steam to hydrogen and carbon (IV) oxide. $C + H_2O \longrightarrow CO + H_2$

3.3.4 Reactions with Strong Oxidizing Agents

When carbon is heated with concentrated trioxonitrate(V) acid or concentrated tetraoxosulphate(VI) acid, it is oxidized to carbon(IV) oxide.

$C_{(s)} + 4HNO_{3(aq)} \longrightarrow 2H_2O(1) + 4NO_{2(g)} + CO_{2(g)}$
$C_{(s)} + 2H_2SO_{4(aq)} \longrightarrow 2H_2O(1) + 2SO_{2(g)} + CO_{2(g)}$

3.4 Oxides of Carbon

Carbon forms two types of oxides, carbon(IV) oxide, $CO2$, and carbon(I1) oxide CO. Both oxides are obtainable as products of the combustion of carbon.

3.4.1 Carbon(IV)Oxide

Carbon(IV) oxide constitute about 0.03% by volume of air and is also found, in Very small quantity, dissolved in water. In the earth crust, carbon(IV) oxide is found as metallic trioxocarbonates(IV) and hydrogen trioxocarbonate(IV) e.g. in limestone regions and coral reefs. Carbon (IV) oxide is very important to green plants which use it to produce food substances e.g. starch, by the process of photosynthesis.

(a) Preparation

Carbon(IV) oxide is prepared in the laboratory by the action of dilute hydrochloric acid or trioxonitrate(V) acid on metallic trioxocarbonate(IV) or hydrogen trioxocarbonate.

$CaCO_3 + 214CI \longrightarrow CaCt_2 + 11_20 + CO_2$
$CaCO_3 + 2HNO_3 \longrightarrow Ca(NO_3) + 14_30 + CO_2$
$NaHCO_3 + HC1 \longrightarrow NaCI + 14_704\ CO_2$

Action of heat on metallic trioxocarbonate(IV) except those of sodium and potagsium, or the hydrogen trioxocarbonates(IV) of sodium, and potassium; is also used for the production of carbon(IV) oxide

$CaCO_3 \longrightarrow CaO + CO_2$
$2NaHCO_3 \longrightarrow Nap + H_2O + 2CO_2$

Carbon(IV) oxide is obtained industrially as by-product in the fermentation process and manufacture of quicklime (CaO) from limestone ($CaCO_3$).

(b) Physical properties

i. $CO2$ is a colourless, odourless and tasteless gas.
ii. It is about 1,5 times denser than air.
iii. It is moderately soluble in water
iv. It turns damp blue litmus paper red because CO_2 dissolves in water to form an acidic oxide, trioxocarbonate(IV) acid.
v. It can readily be liquefied and solidified (-78°C). Solid $CO2$ is known as cardice or dry ice.

(c) Chemical properties

i. CO_2 does not burn nor does it support combustion. However burning magnesium, decomposes CO_2 leaving behind carbon deposit and magnesium(I1) oxide ash.
$CO_2 + 2Mg \quad 2MgO + C$

ii. CO dissolves in water to form trioxocarbonate(IV) acid (soda water). It is a weak acid. $CO_2 + H_2O\ {}^{-}\text{ts} \longrightarrow H_2CO_3$ 2H + I

iii CO_2 reacts directly with alkalis e.g. NaOH, to form trioxocarbonates(IV).
$CO_2 + 2NaOH \longrightarrow Na_2CO_3 + H_2O$

In the presence of excess, CO_2, the trioxocarbonate (IV) is converted to hydrogen trioxocarbonate(IV)

$$Na_2CO_3 + CO_2 + H_2O_4 \longrightarrow 2NaHCO_3$$

When CO, is passed through the alkali, calcium hydroxide (funewater), the lime water turns milky due to the precipitation of insoluble calcium trioxocarbonate(IV). This reaction is used to test for CO_2.

$$\underset{\text{limewater}}{Ca(OH)_2} + CO_2 \longrightarrow \underset{\text{insoluble}}{CaCO_3} + H_2O$$

However, when excess gas is bubbled, the milkiness disappears leaving a clear solution because the insoluble trioxocarbonate(IV) is converted to soluble hydrogen trioxocarbonate(IV).

$$CaCO_3 + H_20 + CO_2 \longrightarrow Ca(HCO3)2$$

(d) Uses

i. Many fire extinguishers use carbon(IV) oxide to put out fires as the gas does not support combustion. Being heavier than air, it envelops the burning material and cuts off the oxygen supply. Carbon(IV) oxide is more effective than water in putting out petrol or oil fires because these materials float on water.

ii. Carbon(IV) oxide is used in the manufacture of:

- Sodium trioxocarbanate0V) (14/asking soda) by the Solvay process.
- Sodium hydrogentrioxocarbonate0V).
- Lead (H) trioxocarbonate(IV).
- Urea and turanonium tetraoxosulphate(VI) - important fertilizers.

iii. Yeast and baking powder are used in baking to produce carbon (IV) oxide which causes the dough to rise, making dough 'light'.

iv. Solid carbon(IV) oxide i.e., dry ice is used as a refrigerant for perishable goods, e.g. ice-cream. On warming, it sublimes and provides a lower temperature. Gaseous carbon (IV) is used to preserve fruits. Carbon(IV) oxide is also used as a coolant in nuclear reactors.

v. Carbon(IV) oxide is used to give carbonated (aerated) drinks a pleasant and refreshing taste e.g. Coca-cola, pepsi-cola, beer, cider, wine and champagne. `Soda-water' drink is a solution of CO2 gas in water under pressure.

vi. Green plants use CO2 during photosynthesis.

3.4.2 Carbon(H) Oxide

CO is produced by the incomplete combustion of carbon compounds, such as octane,C_8H_{18}, found in petrol.

$2C_8H_{18(1)}$, + $170_{2(g)} \longrightarrow 16CO_{(g)} + 18H_2O_{(1)}$

CO occurs in traces as an impurity in the air. The percentage may increase in cities where the gas is released in the exhaust fumes of motor cars, and in industrial areas due to the combustion of fuels. It is a poisonous gas and as little as 0.5% of it in the air may lead to death. As the gas has no colour or odour, its presence is difficult to detect, so it is very dangerous.

a. Preparation

Carbon(H) oxide is prepared by passing carbon (IV) oxide over red-hot carbon $CO_2 + C \longrightarrow 2C0$

It can also be prepared by dehydrating methanoic acid, HCOOH, or ethanedioc acid, $C_2H_2O_4$, using concentrated tetraoxosulphate(VI) acid, which act as dehydrating agent.

$$HCOOH \xrightarrow[-H_2O]{Conc.H_2SO_4} CO$$

$$C_2H_2O_4 \frac{Conc.H_2SO_4}{-H_2O} \quad CO_2 + CO$$

The preparation of CO must be done in a fume cupboard as the gas is poisonous.

In the latter reaction, the CO_2 is removed by passing gaseous products through concentrated sodium hydroxide.

b. Physical Properties

Carbon (II) oxide is a colourless, tasteless and odourless gas. It is insoluble in water, but dissolves in a solution of ammoniacal copper (I) chloride.

i. It is slightly less dense than air,
ii. It has no effect on litmus
(c) Chemical Properties

i. As a reducing agent: Carbon (II) oxide is a strong reducing agent. It reduces some metallic oxides to metals and is itself oxidized to carbon (IV) oxide. $PbO_{(s)} + CO_{(g)} \longrightarrow Pb, + CO_{2(g)},$

ii. It also reduces steam to hydrogen

$$H_2O + CO \longrightarrow CO_2 + H_2,$$

it burns in air with a blue flame to give carbon(IV) oxide

$$2CO + O_2 \longrightarrow 2CO_2$$

iii. Carbon (II) oxide poisoning is as a result of its reaction with haemoglobin in the red blood cells and thus prevents the haemoglobin from transporting oxygen in our body. Death resulting from CO poisoning occur when the supply of oxygen to the body become inadequate because the carrier haemoglobin is not available for this purpose.

(d) Uses

i. CO is used as reducing agent in the extraction of metals e.g. iron from its ores.
ii. CO is an important constituent of gaseous fuels like producer gas ($CO + N_2$) and water gas ($CO + H_2$).

(e) Fuel Gases

These are gas mixtures produced by heating coke with oxygen or steam. All the raw materials coke, air and water are cheap and readily available.

i. When hot coke combines with oxygen of the air, the product is a gaseous mixture, of one-third carbon (II) oxide and two-third nitrogen by volume, called producer gas. The formation of producer gas is exothermic.

$$\underset{\text{Coke}}{2C} + \underbrace{O_2 + 4N_2}_{\text{Form air}} \quad \underbrace{\underset{\text{2 volumes}}{2CO} + \underset{\text{4 volumes}}{4N_2}}_{\text{Producer gas}} + \text{heat}$$

Producer gas has a low heating power because it contains about 67% non-combustible nitrogen and 33% carbon (II) oxide. However, it is inexpensive and is widely used to heat furnaces retorts e.g. in the manufacture of zinc, steel and glass. It is also a source of nitrogen for the manufacture of ammonia (Haber process).

ii. When hot coke combines with steam, the product is a gaseous, mixture, of equal volumes of carbon (II) oxide and hydrogen, called water gas. The formation of water gas is endothermic.

$$\underset{\text{Coke}}{C} + \underset{\text{steam}}{H_2O} \longrightarrow \underbrace{\underset{\text{1 volume}}{CO} + \underset{\text{1 volume}}{H_2}}_{\text{water gas}} - \text{heat}$$

During the process, the coke quickly cools to a temperature too low for reaction if heat is not supplied externally. Industrially both producer gas and water gas are made in the same plant, known as the producer, by passing air and steam alternately through the heated coke. The heat produced when producer gas is formed is sufficient for water gas formation.

Water gas is used to manufacture hydrogen and as an important industrial fuel (both CO and H_2 bum in air releasing a lot of heat).

4.0 CONCLUSION

5.0 SUMMARY

6.0 TUTOR-MARKED ASSIGNMENTS

7.0 REFERENCES/FURTHER READING

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UNIT 2 THE GAS LAWS (I): BOYLE'S AND CHARLES' LAWS AND THE GENERAL GAS EQUATIONS

CONTENTS

1.0 INTRODUCTION

In the previous unit the kinetic theory was used to explain the physical nature of the different states of matter as well as the changes that occur when matter is heated. Can you recall the following as explained in that unit? Melting point, boiling point, vaporisation and condensation.

You will also recall that particles in the gas state are in random motion in all directions and at very high speed with virtually no force of attraction between the particles. The physical behaviour of a gas is very much different from those of the solid and liquid. This physical behaviour of gases was investigated by early scientists and that led to the establishment of gas laws named after them. There is a need therefore to increase the postulates of the kinetic theory to account for the special behaviour of gases.

This is the first of two units on gases and you will learn about Boyle's law and Charles' law and the general gas equation. Statement of the gas laws will be examined and the gas behaviour as established by each law explained by the kinetic molecular theory.

2.0 OBJECTIVES

At the end of this unit you should be about to:

- state the postulates of the kinetic molecular theory for gases.
- state the Boyle's and Charles' laws for gases.
- illustrate Boyle's and Charles' laws graphically.
- use the kinetic molecular theory to explain Boyle's and Charles' laws.
- derive the general gas equation from Boyle's and Charles' laws.
- carry out calculations based on the three gas laws.
- explain the effect of temperature and pressure on the volume of a gas.

3.0 MAIN CONTENT

3.1 The Kinetic Molecular Theory of Gases

Two assumptions are made about gases in the kinetic theory:

i. There is virtually no force of attraction between gas molecules
ii. Gas molecules move independent of each other.

These assumptions are only true for an ideal gas. An idea gas is also often referred to as a perfect gas

Real gases in general do not quite follow the assumptions made for an ideal gas. They constitute what is called the kinetic molecular theory. They specifically deal with the gas molecules. These following six statements describe the behaviour of an ideal gas:

1. A gas consists of small identical particles called molecules moving randomly in all directions colliding with each other and also with the walls of the containing vessel.
2. There is no force of attraction between the gas molecules.
3. Molecular collisions are perfectly elastic i.e. no energy is lost when molecules collide with each other or with the container wall.
4. The volume of gas molecules is negligible compared to the container volume.
5. The collisions of gas molecules with the container walls give rise to what is called the gas pressure.
6. The temperature of the gas is directly proportional to the average kinetic energy of the molecules.

- When of the above will not always be true for a real gas?
- Can you predict the conditions under which a real gas can be assumed to behave ideally?

3.2 Boyle's Law

3.2.1 Statement of Boyle's law

Boyle's law states that at a constant temperature the volume of a given mass of gas is inversely proportional to the pressure.

Note

The words in bold type face is the statement of the law. Boyle's law describes the relationship between the pressure and volume of a gas at constant temperature.

Mathematically

$V \propto \frac{1}{P}$ at constant temperature
OR V = Volume and P = Pressure of gas

$V = \frac{k}{P}$ where k is the constant of proportionality.

$PV = k$

Boyle's law can also be stated mathematically as $P_1 V_1 = P_2 V_2$
P_1 is the pressure when gas volume is V_2 and P_2 is the pressure when gas volume is V_2.

The above implies that the product of the pressure and volume is always a constant. What this means is that as the gas pressure increases, the volume decreases and vice versa as long as the temperature is constant.

Boyle's law can also be presented graphically

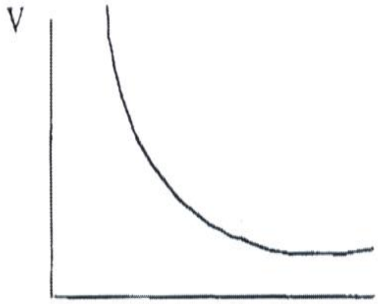

Fig. 3.1

From the above graphical illustration, you see That as pressure increases, the volume decrease.

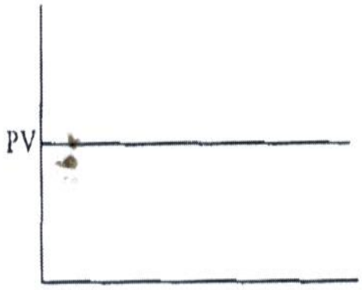

Fig. 3.2

This figure shows that the product of pressure and volume is constant.

Consider the equation in the form PV = k

The equation can be written as $P = \frac{k}{V}$

The above equation is the mathematical representation of the alternative statement of Boyle's law:

The pressure of a given mass of gas is inversely proportional to the volume provided the temperature is constant

Examples

1. At 25°C and 750mm Hg pressure the volume of a given mass of oxygen is 520cm^3. Calculate the gas volume when the pressure is increased to 760mm Hg pressure at constant temperature.
2. A gas occupies a volume of 1.20dm^3 at 25°C and 760mm Hg pressure. What will be gas pressure when the gas expands into 2.80dm^3 container at the same temperature?

Answers

1. Let P_1 = 750mmm Hg and V_1 = 520cm^3
   P_2 760mm Hg and V_2 = ?

   Using $P_1V_1 = P_2 V_2$

   $$\frac{P_1 V_1}{P_3} = V_2$$

   $$\frac{750 \text{ x } 520}{760} = V_2 = 513 \text{ cm}^3$$

2 Let P_1 = 760mm Hg V_1 = $1.20dm^3$
P_2 = ? V_2 = $2.80dm^3$

Using $P_1 V_1 = P_1 V_2$

$$\frac{P_1 V_1}{V2} = P_2$$

$$\frac{760 X 1.20}{2.80} = P_2 \; 326\text{mm Hg}$$

3.2.2 How kinetic theory explains Boyle's Law

According to the kinetic molecular theory of gases an increase in pressure of the gas implies an increase in the number of molecular collisions per unit area of the containing vessel. You will also recall that for a fixed mass of gas at constant temperature, the number and the average kinetic energy of the molecules remain constant for the number of molecular collisions per unit area of the containing vessel to increase under a condition of fixed mass and average kinetic energy of the gas molecules, the area and therefore the volume of the containing vessel must decrease. This implies that an increase in gas pressure is accompanied by a corresponding decrease in the gas volume. The reverse of the above is true when the gas pressure is reduced at constant temperature.

3.3 Charles' Law

3.3.1 Charles' Law and the Kelvin Temperature Scale

The term temperature is a measure of the degree of hotness or coldness of a body. The Celsius or centigrade (℃) scale is the most used temperature scale in scientific work. The volume of matter generally increase with increasing temperature but the increase is most pronounced for gases.

- Can you recall the name of the instrument used to measure temperature?
  Charles study on gases show that the volume of a gas increases by 1/273 of its volume at. O°C for every. °C rise in temperature.

Mathematically stated

$$V\tau = Vo + \frac{T}{(273)} V_o$$

$V\tau$ = Volume of the gas as T℃
V_o = Volume of gas at 0°C.

According to the above equation, a gas has zero volume at - 273°C. This temperature at which the gas volume is theoretically zero is the lowest temperature that can be reached. It is called the absolute zero temperature. The Kelvin temperature scale represented with a capital K has this temperature (-273) as its starting point and measures temperatures upwards from it.

Recall the relationship between temperature and the average kinetic energy of gas molecules and explain why it is assumed that the average kinetic energy of gas particles at -273°C is zero.

The Celsius and the Kelvin (absolute) scales are related by the equation.

$$T(K) = t°C + 273$$

Let us now try using the above formula to complete the table below.

t°C	**T(K)**
20	-
-	325
-15	-
77	-
-	40

In calculations involving the gas laws, temperature readings must be converted to the Kelvin scale.

Note

Temperatures on the Kelvin scale are in K units with no degree sign.
e.g. 40K and 374K

3.3.2 The Statement of Charles' Law

Charles' law states that for a fixed mass of gas, the volume is directly proportional to the absolute temperature provided the pressure remains constant.

Mathematically

$V \propto T(K)$ V - Volume of gas
OR T - Kelvin temperature
$V = kT$ Where k in constant.

$$\frac{V}{T} = k$$

OR

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ where V_1 is the volume at T_1 and V_2. is the volume at T_2

Charles' law can also be stated in an alternative form when the gas volume is constant and pressure changes with temperature. In this case

$P \propto T_2$ $P = kT; = k\frac{P}{T} *$ $= \frac{P_1}{T_1}$ $\frac{P_2}{T_2}$

at constant volume, the pressure of a fixed mass of gas is directly proportional to the Kelvin temperature. Charles' law can also be presented graphically as shown below.

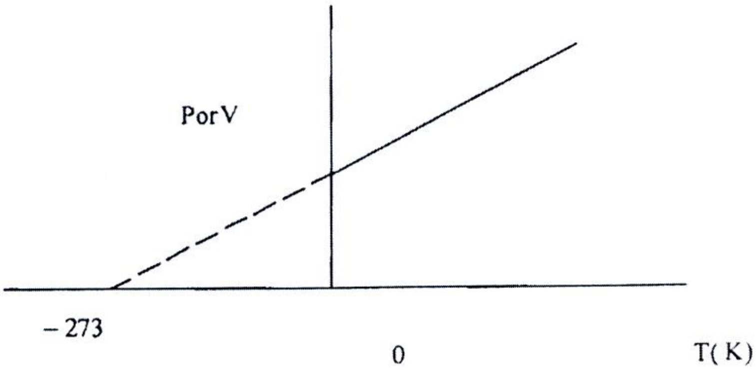

Fig 3.2

We can now solve this problem

At 27°C the volume of mass of helium gas is 1.85de. At what temperature will the volume be 2.25dm3 if pressure is kept constant.

V_1 = $1.85dm^3$

T_1 = 273 + 17 = 300K

T_2 = ?

V_2 = $2.25dm^3$

Using equation $\frac{V_1}{T_1} = \frac{V_1}{T_2}$

$V_1T_2 = V_2\, T_1;\ T_2 = \frac{V_2\ T_1}{V_1} = \frac{2.25\ X\ 300}{1.85}$

= 364.9K

SELF ASSESSMENT EXERCISE

The pressure of a mass of oxygen gas at 25°C is 783nun Hg in a 1.0dm' container. What pressure will the gas exert when the temperature is increased to 40°C without any change in the volume?

3.3.3 How the Kinetic Theory Explains Charles' Law

According to the kinetic molecular theory, an increase in gas temperature leads to increase in the average kinetic energy of the gas molecules. This will result in more random motion of the molecules. For a fixed mass of gas, the number of gas molecules is constant and at constant pressure of gas the number of collisions per unit area of the containing vessel is also constant. The only way for the pressure to remain constant despite the increase in the random motion and more collisions of particles with the wall of the container is for the area and therefore the volume of the container to increase. This implies that an increase in the temperature of the gas at a constant pressure leads to a corresponding increase in the gas volume.

In the alternative, consider the case when the volume of a fixed mass of gas is kept constant and the temperature is increased. At the higher temperature, the more random motion of the gas molecules will lead to more molecular collisions per unit area on the wall of the container and thus increase in the gas pressure.

3.4 The General Gas Equation

A gas sample is described completely by four quantities: the amount, the temperature, pressure and volume. According to Boyle's and Charles' laws both pressure and temperature can affect the volume of a gas.

According to Boyle's law

$V \propto \frac{1}{P}$ and according

to Charles' law $V \propto T$

when both T and P change for a gas sample $V \propto \frac{T}{P}$; or $V \frac{kT}{P}$

$PV=kT; \quad \frac{PV}{T} = k$

$\therefore \quad \frac{PV}{T} = k$ (constant)

Sometime the above is given as

$$\frac{P_1\ V_1}{T_1} = \frac{P_2\ V_2}{T_2}$$

The last is the general gas equation and is Boyle's and Charles' law combined. For one mole of gas, the constant of proportionality (k) in the general gas equation is the molar gas constant (R) and the general gas equation then reads;

PV = RT. Fore moles of gas the equation has the form
PV = nRT.

The value and units of the gas constant R depends on the units of P, V, and T.

P in atm.
V in dm^3
T in K

From the general gas equation

$$PV = nRT$$

$$\text{and } R = \frac{PV}{nT}$$

The value of R is 0.082 atm dm^3 K^{-1} mot^{-1} when pressure is in atmosphere, volume in dm' and temperature in Kelvin. But when pressure is in Nm^2, volume in m^3 and temperature in Kelvin, the value of R is 8.31 JK^{-1} mot^{-1} (Joules per degree Kelvin per mole).

Example

1. A mass of oxygen gas occupies a volume of 2 x 10^4 cm^3 at a pressure of 1.01 x 10^5 Nm^{-2} and temperature of 17°C. Calculate the volume when the pressure is increased to 1.35 x 10^5 Nm^{-2} and the temperature changed to 43°C.

Solution

P_1	=	1.01 x 10^5 Nm^2	P_2	=	1.35x 105/4m-2
T_1	=	273+17 =300K	T_2	=	773 + 43 = 316K
V_1	=	2 x 10^4 cm^3	V_2	=	?

Using the general gas equation in the form

$\frac{P_1 V_1}{T_1} = \frac{P_2\ V_2}{T_2}$ and making V_2 the subject of the formula

$$V_2 = \frac{1.10\ X\ 10^5\ X\ 2\ X\ 10^4\ X\ 316}{300\ X\ 1.35\ X\ 10^5} = 1.58 \text{ x } 10^4 cm^3$$

4.0 CONCLUSION

In this unit the physical behaviour of an ideal (perfect) gas is summarised in the assumptions of the kinetic molecular theory of gases. Some of the assumptions of the theory may not always be true for real gases, explaining why real gases sometimes deviate from ideal behaviour. The quantities, pressure, temperature, volume and amount (mole) are important in measurements and calculations involving gases. These parameters describe the behaviour of the gas and a change in one of them will result in modification of their behaviour.

5.0 SUMMARY

This unit presents the kinetic molecular theory of gases. The theory is used to explain Boyle's and Charles' laws. Apart from giving the statements of the two laws they are graphically illustrated and employed in simple calculations.

The general gas equation was derived from Boyles and Charle's laws. Note that in calculations involving the gas laws, the Kelvin temperature is used.

6.0 TUTOR-MARKED ASSIGNMENTS

1. Identify two assumptions in the Kinetic theory that can account for observed deviations of real gases from ideal behavior.
2. Calculate the pressure of 0.20 mole oxygen gas placed in a 5.0dm^3 vessels at 25°C. State the effect if any on (i) the pressure (ii) the volume (iii) the number of moles, when the gas is heated.

7.0 REFERENCES/FURTHER READING

Osei Yaw Ababio (2002) New School Chemistry. Onitsha African-Fep Publishers.

UNIT 3 THE GAS LAWS II DALTON'S, GRAHAM'S, AVOGADRO'S AND GAY LUSSAC'S LAWS

CONTENTS

1.0 INTRODUCTION

In unit 2 the ideal or perfect gas is described in the postulates of the kinetic theory. The ideal or perfect gas does not exist. It is only hypothetical.

Real gases deviate from ideal behaviour and do not obey the gas laws perfectly. Imagine a real gas subjected to low temperatures and very high pressures. The gas molecules are compressed into a very small volume and are not very energetic. Intermolecular forces of attraction e.g. Vander waals and dipole-dipole attractions come into play between the gas molecules. The volume of gas molecules will also become significant when compared with the space occupied by the gas.

In this unit you will learn more about gas behavior in Dalton's, Graham's, Avogadro's and Gay Lussac's laws. These laws are also adequately explained by the kinetic theory and like the previous ones, are obeyed perfectly by ideal gases only. Most real gases however obey these laws with very little deviations at ordinary laboratory conditions of temperature and pressure.

2.0 OBJECTIVES

At the end of this unit you should be able to:

- account for deviations of real gases from ideal behaviour.
- state the statement of the laws of dalton, graham, avogadro and gay lussac.
- what avogadro constant is and relate it to the mole.
- use gay lussac's law to predict the volume of a gaseous product in a gas reaction.
- apply the laws of dalton, graham, avogadro and gay lussac in simple calculations.
- explain why gases diffuse faster than liquids and solids.

3.0 MAIN CONTENT

3.1 Dalton's Law

3.1.1 Dalton's Law of Partial Pressures

Dalton's law states that the total pressure exerted by a mixture of gases in a container is equal to the sum of the partial pressures of the gases present in the mixture provided no chemical reaction occurs.

The partial pressure of a gas in a mixture is the pressure the gas would exert if it occupies the total volume of the mixture alone. Mathematically for a mixture of three gases A, B, C in a container.

$P_{Total} = P_A + P_B + P_C$ where P_{Total} is the total pressure of the mixture and $P_A, + P_B, + P_C$ are the partial pressures of A, B and C respectively.

Recall the general gas equation PV = nRT

For gas A, $P_AV = n_A RT$
For gas B, $P_BV = n_B RT$
For gas C, $P_CV = n_C RT$ where n_A, n_B and n_C are the number of moles of A, B and C in the mixture.

$P_{Total} V = (n_A + n_B + n_C) RT$ and

$P_A = x_A P_{Total}$, $P_B = x_B P_{Total}$ and $P_C = x_C P_{Total}$ where x_A, x_B and x_C are the mole fractions of A, B, and C in the mixture. It follows from the above that the partial pressure of a gas in a mixture is equal to the product of total pressure and the mole fraction of that gas in the mixture.

Dalton's law finds application in calculations and determination involving volumes of gases that contain water vapour as the case in the laboratory preparations of gases that are collected over water e.g H_2, O_2 and N_2.

Example

A 1.0dm^3 gas jar contains 0.02 mole oxygen and 0.125 mole nitrogen at 27C

Calculate,

a. the partial pressures of oxygen and nitrogen.
b. the total pressure of the gas mixture R = 0.082 litre atm K^{-1} mot^{-1}

Recall the general gas equation

PV = nRT

$P = \frac{n\,RT}{V}$ T = 273 + 27 = 300K

(a) For O_2; $PO_2 = \frac{0.02\ (.082)\ x\ 300}{1.0}$

$= 0.492$ atm

For N_2; $PN_2 = \frac{0.125\ (.082)\ x\ 300}{1}$

$= 3.075$ atm

(b) P_{Total} $Po_2 + PN_2$

$= 0.492 + 3.075$

$= 3.567$ atm

3.1.2 How the Kinetic Theory Explained Dalton's Law

When two or more gases mix without a chemical reaction, each gas behaves independently of the other since there are no attractive forces between the molecules. Each gas will exert its own pressure on the container wall. The total pressure of the gas mixture will therefore be the sum of the pressures of the gases in the mixture.

3.2 Graham's Law of Diffusion

Diffusion is a natural phenomenon. It involves the movement of particles through a medium from a region of high particle concentration to a region of lower particles concentration, Diffusion occurs in all the states of matter. By virtue of the very high kinetic energy of gas particles, diffusion is fastest in gases.

Graham's law expresses the relationship between the rate of diffusion of a gas (R) to its vapour density and explains why the law is called Grahams's Law of diffusion. The law states that at a constant temperature and pressure, the rate of diffusion of a gas is inversely proportional to the square root of Its density.

Mathematically;

$RO_{\frac{1}{1}}$ where R is the rate of diffusion and τ the gas density.

From the above relationship the higher the density of the gas the slower it will diffuse. When the rates of diffusion for two gases 1 and 2 are competed or divided the equation becomes

$\frac{R_1}{R_2} = \sqrt{\frac{\rho_2}{\rho_1}}$ where R_1 R_2 are rates and ρ_1 ρ_2 are densities respectively

Graham's law can also be stated in terms of the relative molar masses of the gases

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

The rate of diffusion is inversely related to the time of diffusion i.e. the higher the rate of diffusion, the shorter the time for a given mass of gas to diffuse. It follows from the above that

$\frac{R_1}{R_2} = \sqrt{\frac{t_2}{t_1}}$ where t_1 t_2 are the the times for the same mass of gases 1,2 to diffuse under the same conditions.

SELF ASSESSMENT EXERCISE

It takes 20 minutes for a given mass of oxygen to diffuse through an opening. How long will it take the same mass of the following gases to diffuse through the same opening under the same conditions of temperature and pressure?

(i) SO_2 (ii) CO_2 (iii) H_2S (iv) NO_2

3.3 Avogadro's Law and its Applications

This law states that equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

Avogadro's law provides information on all number of molecules in a given volume of gas and allows for change over from statements relating volumes of gases to statements relating to the number of molecules or moles.

e.g. $H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$

1 vol of $H_{2(g)}$ reacts with 1 volume of $Cl_{2(g)}$ to give 2 volumes of $HCl_{(g)}$. Applying the Avogadro's law this implies that 1 mole of $H_{2(g)}$ reacts with 1 mole of $Cl_{2(g)}$ to give 2 moles of $HCl_{(g)}$. Application of the law as above has allowed for the establishment of formulae for gases. At 0°C and 760 mmHg pressure, $22.4dm^3$ of any gas contains approximately 6.02×10^3 molecules i.e. 1 mole of gas molecules. 0°C and 760 mmHg pressure are referred to as standard temperature and pressure (S.T.P) respectively and 22.4 dm^3 is the molar volume of gas at S.T.P.

3.4 Gay Lussac's Law of Combining Volumes

This law states that when gases react, they do so In volume, which bear a simple ratio to another and to other products (if gaseous) provided temperature and pressure remain constant.

Applying Avogadro's law, Gay Lussac's law can also be restated as follows:

When gases react they do so in small whole numbers of molecules of reactants to produce small whole numbers of the products provided temperature and pressure remain constant.

e.g. $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

SELF ASSESSMENT EXERCISE

In the contact process for the production of tetraoxosulphate (vi) acid, a step of the production process involves the oxidation of sulphur (iv) oxide gas to sulphur (vi) oxide gas. 2 volumes of the lower oxide reacts with l volume of oxygen to give 2 volumes of the higher oxide.

i. Use Gay Lussac's and Avogadro's law and write a balanced chemical equation for the reaction.
ii. Calculate the volume of oxygen gas required at S.T.P. to completely react with 6.4g of sulphur (iv) oxide.

4.0 CONCLUSION

This unit concludes our discussion on gases and gas behaviour though the theory used in explaining the laws talks of an ideal or a perfect gas. Most real gases obey the gas laws at normal laboratory conditions with very little deviation. Understanding gas behaviour should help us to manage, control and improve our environment for better living. Remember that the earth is surrounded by air which is a mixture of gases.

5.0 SUMMARY

- This unit presented to you four gas laws: Dalton's. Graham's. Avogadro's and Gay-Lussac's
- Statements of these laws arc recalled and interpreted mathematically and applied to numerical problem.
- The concept of standard temperature and pressure and the molar volume are introduced.
- Diffusion in the gas phase is a very rapid process and explains why gases mix.
- Provided there is no chemical reaction, each gas in a mixture behaves independently of the other.

Can you relate the above statement to the air in our environment.

6.0 TUTOR-MARKED ASSIGNMENTS

1. 50.0cm^3 of nitrogen gas was collected over water at 17°C and 770 mmHg pressure. Calculate the volume of the dry gas at S.T.P. Saturated vapour pressure of water at 17°C is 14 mmHg.
2. 0.30g of a vapour of a volatile liquid occupies a volume of 0.062dm^3 at 27°C' and 757 mmHg. Calculate the molar mass of the liquid substance.

7.0 REFERENCES/FURTHER READING

Rajah, S. T, Teibo, B. O., Onwu. 6 and Obikwere, A. (2002). Senior Secondary Chemrsity Textbook I Lagos. Longman Publishers.

Osei Yaw Ababio (2002), New School Chernictry. Onitsha. Africana-HP Publishers.

UNIT 4 LIQUIDS

CONTENTS

1.0 INTRODUCTION

In unit 11 the general physical features of the solid, liquid and gas state of matter were highlighted. According to the Kinetic theory the liquid particles have lower average Kinetic energy than the gas and movement is a intermediate between those of the solid and the gas. Liquid particles are much wider apart and are held by weaker forces of attraction than in the solid state.

You will recall from the heating curve that, when solid matter is heated it changes to the liquid form at the melting point. The cooling of the vapour (gas) also results in liquid formation, the process is called condensation. The liquid state is therefore an intermediate state between the solid and gas states, liquids are fluids, like gases. They flow and can be poured from one container to another. Unlike gases they have fixed volumes and are uncompressible.

- Now list some differences and similarities between solids and liquids.

In this unit, the liquid state is revisited with a view to highlight and explain some concepts about liquids and the liquid state.

2.0 OBJECTIVES

At the end of this unit you should be able to:

- list some properties of liquids that make them resemble gases.
- define saturated vapour pressure.
- list factors that affect the saturated vapour pressure of liquids.

- define boiling and boiling point
- define the term 'normal boiling point'.
- describe a method to determine the boiling point of a liquid.

3.0 MAIN CONTENT

3.1 The Vapour Pressure of Liquids

You will recall that a liquid sample consists of molecules with varying Kinetic energies. Some particles have very high energies while some have low energies. Now consider a liquid sample placed in a closed container in the laboratory. Some energetic particles that are close to the liquid surface will vaporise and constitute a vapour pressure above the liquid. This pressure builds up until the space above the liquid is saturated with vapour. When this happens a dynamic equilibrium is set up between the liquid particles in the vapour and the liquid phase.

The pressure of the vapour under this equilibrium condition is called the saturated vapour pressure (SVP) of the liquid. It is the maximum vapour pressure of the liquid at that temperature. The SVP is not dependent on the liquid volume in the container and it increases with increasing temperature Fig14.1

— Try and explain this increase with temperature using the Kinetic theory.

Liquids with strong intermolecular forces generally have low saturated vapour pressure than those with weak intermolecular forces.

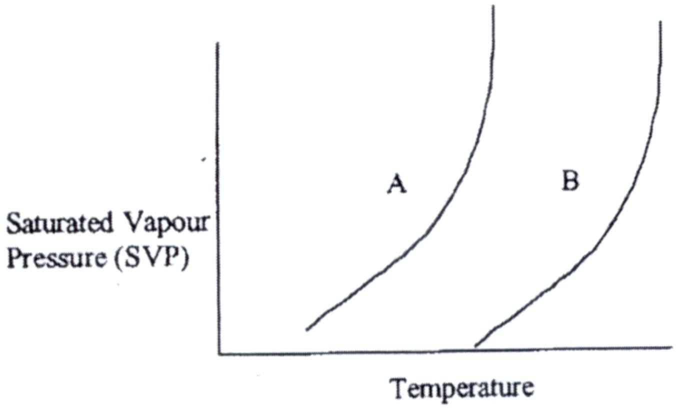

Fig 14.1: Change of saturated vapour pressure with temperature.

3.2 Boiling and Boiling Point

The atmosphere around us exerts a pressure on all objects on the earth surface. The pressure of the atmosphere is called atmospheric pressure. Its value decreases with altitude (height above sea level). This implies that the atmospheric pressure on top of a high mountain is less than at the base.

A heated liquid is also acted upon by the atmospheric pressure but as it is heated, its saturated vapour pressure increases opposing the downward effect of the atmospheric pressure. At a particular temperature characteristic of the liquid, the saturated vapour pressure becomes equal to the atmospheric pressure. At this temperature vaporisation becomes most rapid and bubbles of vapour form even from the interior of the liquid. This is boiling and the temperature is called boiling point. By definition the boiling point is the temperature at which the saturated vapour pressure of the liquid equals the prevailing atmospheric pressure.

The above definition does not allow for a reference point of measurement because the atmospheric pressure is not a constant at every point on the earth's surface. This is because the atmosphere contains many gases including water vapour. It is thus a mixture and depending on the place in the earth and the conditions, its composition is not always the same. The normal (standard) boiling point of a liquid is the temperature at which its saturated vapour pressure equals 760nunHg is called 0 atmosphere pressure. The menial boiling point of water is 100°C but because the atmospheric pressure in Zaria is slightly less than 76OrnmHg water boils at less than 100°C. Measurements show that water boils at 98°C in Zaria. Do you know at what temperature water boils in your locality?

The saturated vapour pressure of acetone at 20°C is much higher than that of water at the same temperature which liquid will have the higher boiling point, water or acetone? Give reasons for your answer.

3.3 Boiling Point - A Criterion of Purity

Most substances in nature exist in the inpure state. For instance polluted water poses danger to the public therefore water is purified before drinking. In the study of chemistry, purification techniques, are studied. Now recall some of the purification techniques taught in Unit 1.

Even after purification, a test of purity is important so as to ascertain the success of the purification process. For liquid sample, boiling point is used to ascertain the purity of a liquid sample. A pure liquid has a fixed boiling point i.e. the boiling is always at a fixed and constant temperature provided the external pressure is constant.

- What will be the effect of pressure increase on the boiling point of a pure liquid?

Until all the liquid is vaporised at the boiling point the liquid/vapour mixture will maintain a constant temperature.

3.4 Methods of Boiling Point Determination

Several methods have been developed for determining the boiling point of liquids. Here we shall consider two methods. The first is used for liquids with non poisonous vapours while the second, is used for liquids with poisonous vapours.

The easiest method is to insert a thermometer into a boiling Liquid in an open glass container. The temperature must be read with the thermometer inside the boiling liquid. This method can be used for liquids whose vapoured are not poisonous or offensive. The distillation method is recommended for liquids with poisonous or offensive vapour.

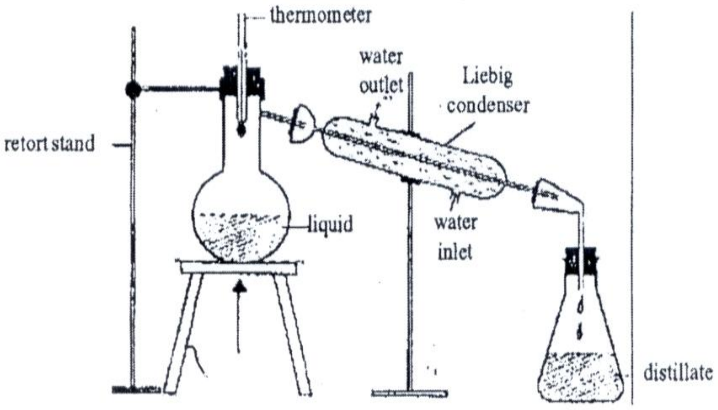

Fig 14.2 Simple distillations

The liquid is placed in the flask labelled A. When the flask is heated the satured vapour pressure of the liquid increases until it becomes equal to the pressure above the liquid. Boiling occurs and the thermometer records the temperature of the liquid/vapour mixture for liquids that are flammable the heating must be done electrically or indirectly using a liquid bath. For liquids with boiling points lower than I00°C, a water or steam bath may be used.

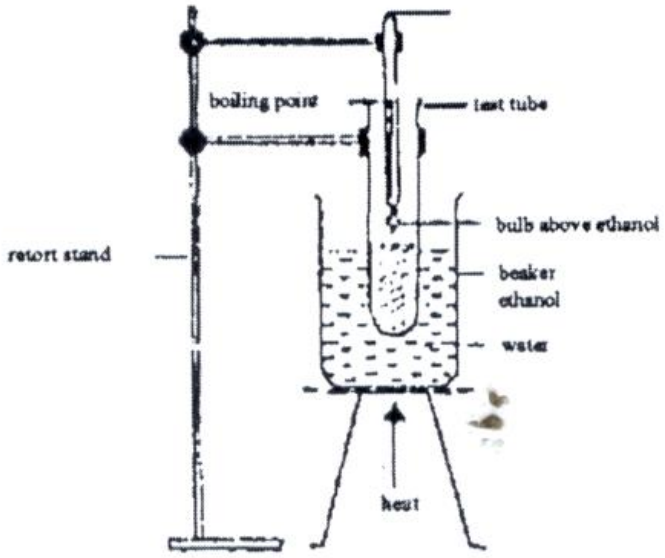

Fig 14.3 Determination of boiling point of alcohol

4.0 CONCLUSION

In this unit you are presented with some facts about the liquid state of matter. Liquids, like gases, are used daily in domestic and commercial ventures. The saturated, vapour pressure of a liquid expresses its relative volatility. A liquid with very high normal boiling point is less volatile than one of low boiling point. Volatile liquids are stored in air fight containers more so when the vapour is poisonous or offensive. e.g. perfumes, petrol, kerosene, mercury and laboratory chemicals that are liquids.

5.0 SUMMARY

- The liquid state is presented and its intermediate nature between the solid and gas explained. Concepts such as, saturated vapour pressure, boiling and boiling points are explained.
- Normal boiling point is defined. Values of normal boiling points can be used to ascertain the relative volatility of liquids.
- Boiling point is a (criterion) of purity for liquids. A pure liquid boils at a 'constant and fixed temperature provided the external pressure is constant.
- Methods of boiling point determination are discussed.

6.0 TUTOR-MARKED ASSIGNMENTS

1. The saturated vapour pressures of four liquids A, B, C, and D are 45, 15, 135 and 75 mm Hg at 25°C respectively.

a. Arrange the liquids in the order of increasing boiling points.
b. Arrange in the order of increasing volatility. Give reasons for the order of arrangement.
2. Define the terms normal boiling point differentiate between boiling and boiling point.

7.0 REFERENCES/FURTHER READING

Rajah, ST., Teibo, B.O., Onwu, G. and Obikwere A. (2002). Senior Secondary Chemistry Textbooks Lagos Longman Publishers.

UNIT 5 SOLIDS

CONTENTS

1.0 INTRODUCTION

In unit 11, the general features of the solid state are highlighted and explained using the kinetic theory. You will recall that the closeness and low average kinetic energy of solid particles allow for strong attractive forces keeping particles in fixed positions. Solid matter therefore occurs naturally in different shapes and sizes e.g. small stones and big rocks and mountains.

Scientists have also synthesised, modified, shaped and reshaped solid mater e.g. polymetric materials such as resins, fibres, rubbers and plastics are available in various shapes, sizes and colours: plastic buckets, cups etc, textile materials, electrical parts: plugs; cables and sockets, motor vehicle and machinery parts are just few examples. Unlike the gas and liquid matter that require containers for storage and moving them from place to place, solid matter is the easiest to handle and move around. The following are some of the very important use of solids; support, shelter, for movement and for food.

In this unit, the focus is on the solid state. Classification, properties and particle arrangement are also discussed. The significance of solid melting point and its use as criterion of purity for the solid are also discussed. In particular the structures of the element, carbon, diamond and graphite forms are fully discussed.

2.0 OBJECTIVES

At the end of this unit you should be able to:

- list some use of solid matter,
- identify the different classes of solids.
- list some properties of the different types of solids.
- compare and contrast the properties of diamond and graphite.
- describe a simple experiment to determine the melting point of a solid.
- recall that the melting point is a criterion of purity.

3.0 MAIN CONTENT

3.1 Solid Classification

Solids are classified as crystalline or amorphous by virtue of the order in particle arrangement.

Crystalline solids are built from simple structural units which repeat themselves in three dimensions to give the solid its characteristic geometric form, Most metal and ionic solids are crystalline in nature.

Other solids such as glass, rubber and plastics do not have definite geometrical forms and are called amorphous The regular and repeated arrangement of particles in the crystalline form a pattern called the crystal lattice.

3.2 Solid Properties

Solids are generally characterised by hardness/firmness and high mechanical strength. The mechanical strength of any solid material is determined by the type of bonds holding the particles in position in the solid structure. The type of force in turn depends on the chemical nature of the units that occupy the fixed position in the Lattice. Solids are grouped into four classes based on the nature of the intermolecular forces between the particles of the solid matter.

3.2.1 The Molecular Solid

The lattice points in molecular solids are occupied by molecules which do not carry any charge. The forces holding the molecules together are of two types, the dipole-dipole interaction for polar molecules and Van der wools forces which are the non-polar molecules. You will recall that these forces i.e. the dipole-dipole and Van der Wools forces are relatively weak forces. This is why molecular solids are usually soft and of very low melting points. The lack of electrical charge in the structure makes them non conductors. They are used mainly as insulators. Examples of compounds that form molecular solids are ammonia, ice water, candle wax and carbon(iv) oxide. Many molecular solids are liquids or gases at normal laboratory temperatures and pressures. Water for instance is a liquid and carbon(iv) oxide is a gas at normal laboratory temperature.

3.2.2 The Metallic Solid

The lattice points in the metallic solid is occupied by the positive core of the metallic atom surrounded by an electron cloud. The force of attraction here is electrostatic between the positive metal cores and the electron cloud. This electrostatic force is strong and is responsible for the compact structure of metals and their very high boiling and melting points. The mobility of the electron cloud under the influence of heat and electricity accounts for the high thermal and electrical conductivities of metals. Examples of metallic solids are copper, iron, lead, aluminum and zinc.

Can you now explain the following?

- An aluminum spoon when placed in hot water soon feels hot to the hand.
- Exposed electrical cables when touched or stepped on can lead to shock or death.

Apart from the high thermal and electrical conductivities, metallic solids are malleable and ductile i.e. can be beaten into shapes and drawn into thin wires e.g. copper wires and aluminum doors.

3.2.3 The Ionic Solid

The units occupying the lattice points in ionic solids are positive and negative ions. There is very strong attraction between the positive and negative ions. In the solid structure each ion is surrounded by as many ions as possible of the opposite charge leading to a giant assembly of ions held together in a triad structure. Ionic solids are characterised by the following properties

i. High melting and boiling points.
ii. They are non conductors of electricity in the solid sate but conduct when molten or in solution.

iii. Because the energy to separate the ions is relatively high, ionic solids are hard and brittle. Examples of ionic solids are NaC1, KCI, KNO_3 and $CuSO_4$.

- Why do ionic solids conduct electricity in the molten state or in solution and not in the solid form?

3.2.4 The Covalent Solid

The lattice points in covalent solids are occupied by atoms linked together by a 'continuous system of covalent bonds. The shared electrons between the atoms in the lattice result in strong bonding between the atomic nuclei. Examples of covalent solids are diamond, graphite and quartz (Si01). When all the valence electrons

are involved in beading as in diamond and quartz, the solid will not conduct electricity. In some arrangement where all valence electrons are not used in bonding as in graphite, electrical conductivity is possible because the free electrons move under the influence of applied electrical potential.

3.3 The Structure of Diamond and Graphite

You will recall that diamond and graphite are allotropes of carbon

- What are allotropes?

Both of than are also covalent solids but while diamond is hard, have very high density and melting point, graphite is soft, less dense and have a much lower melting point. In both diamond and graphite, the lattice points art: occupied by carbon atoms. In diamond, carbon atoms are tetrachedrally bonded using Sp^3 hybridised orbitals. All the four valence electrons of carbon are used in bonding resulting in a rigid three dimensional network. This accounts for the very hard nature of diamond, Diamond is the hardest substance known. It cat not be cut by any other substance, hence it is used in cutting glass and metals.

In the graphite structure however, each carbon atom is linked to three others using Sp' hybridised orbitals. The carbon atoms are arranged in layers one above the other. The layers are held together by weak van der weal fates. Graphite has a relatively high melting point and it is less dense than diamond. The presence of mobile electrons in the crystal lattice makes graphite a good conductor of electricity.

Recall that carbon has four valence electrons awl only three are used in bonding in graphite but all the four electrons are used in diamond.

Graphite is used as an inert electrode in electrolysis. The layered structure of graphite allows one layer to slide over another easily wanting for the lubricating property of graphite. Unlike oil, it is non-sticky and is usually used on bicycle chains and for bearings of some motor cars.

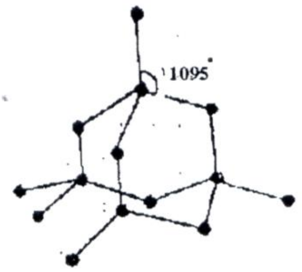

Fig. 15.1: Structure of diamond

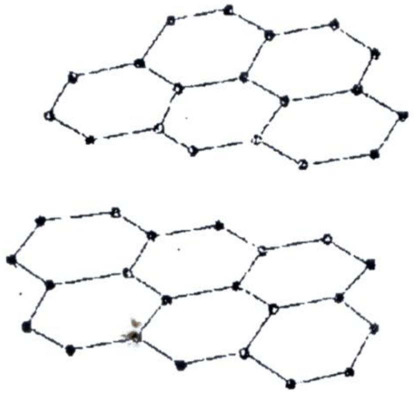

Fig. 15.2: Hexagonal rings in graphite

Some properties of diamond and graphite are summarised in the table 3.1

Property	Diamond	Graphite
1. Appearance	Colourless, transparent solid, sparks when cut	Black opaque with metallic lustre
2. Crystalline structure	Three dimensional network	Two dimensional network
3. Density	Very dense	Less dense
4. Hardness	Hardest natural substance known	Soft
5. Electrical conductivity	Non conductor	Good conductor
6. Chemical activity	Inert unreactive	More reactive
7. Uses	(i) used as a precious jewel	1. dry lubricant
	(ii) drilling rocks and cutting glass	2. electrodes in electrolysis.
	(iii) to sharpen tools	3. used in 'lead' pencil
	(iv) pivot supports in precision instruments	4. black pigment in paints
	(v) dies for drawing wires	5. neutron moderator in atomic reactors

3.3.1 Melting and Melting Point

Heating of solid matter increases the average kinetic energy of the particles. You will recall that at a particular temperature characteristic of the solid the particles will acquire sufficient energy and overcome the attractive forces holding them in fixed positions. At this temperature the solid structure collapses and a liquid is formed. This is the melting point. As earlier explained in unit II, melting of a pure solid sample occurs al a constant temperature and is a test of purity.

The melting point of a solid sample is determined by the capillary tube method. A small amount of the solid sample is put in a glass capillary tube. The capillary tube is attached to a thermometer using a rubber band. The capillary tube is heated in a transparent liquid bath and the temperature at which the solid starts melting is noted. For a pure solid sample, this temperature remains constant until all solid is melted.

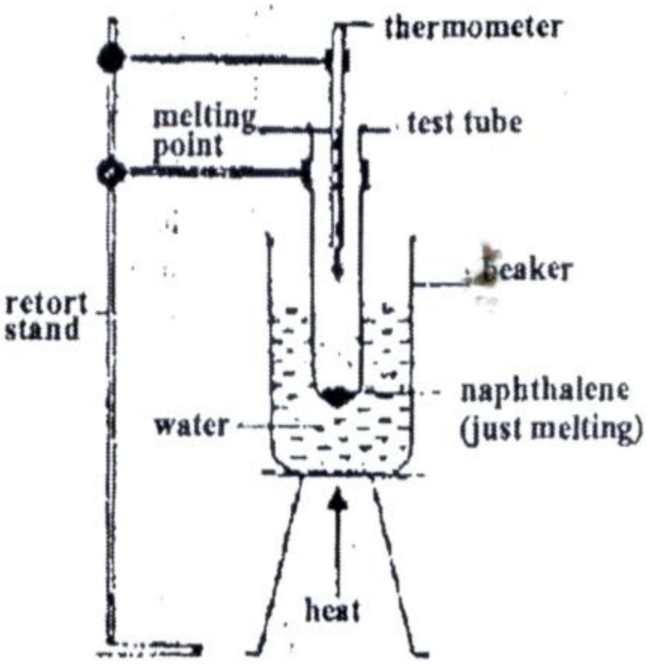

Fig. 15.3: Determination of melting point of naphthalene

Electrically operated melting point apparatus can also be used but the equipment is fairly expensive. The apparatus is fitted with a thermometer. It also has apartments for placing the capillary tube. After the capillary tube containing the solid sample is put into the apparatus, the apparatus is switched on. As the solid heats up, the process is monitored through a magnifying glass in front of the apparatus until melting occurs.

4.0 CONCLUSION

In this unit you have learnt that the physical behaviour of the solid is not only determined by the particle arrangement but also on the nature of the particles that occupy the lattice points which in turn determines the type and strength of the intermolecular forces holding the particles in the solid. Solids with weak intermolecular forces have low melting points while those with strong cohesive forces have fairly high melting points.

Solids are used for a variety of applications and the choice of a solid for a particular application must be only after its physical and chemical behaviour have been properly investigated. Where a solid must be used pure, the melting point determination offers a way of ascertaining its purity.

5.0 SUMMARY

In this unit, the solid state is presented. The important roles of solids in our daily life were highlighted. Solids are classified according to the degree of order in the particle arrangement and the type of particles that occupy the lattice point & the ionic, covalent, metallic and molecular solids are discussed and their physical properties explained in relation to the nature of the particles in the solid structure. The melting point is again presented and its use as a criterion of purity emphasised.

Graphite and diamond are discussed, their structures and properties compared and contrasted. Their uses differ and these uses are presented as part of the study.

6.0 TUTOR-MARKED ASSIGNMENTS

1. Explain the following observations

    a. Chlorine (Cl_2) is a gas while bromine is a liquid and iodine is a solid at normal laboratory temperatures and pressures.
    b. Diamond is hard but graphite is soft.
2. List some uses of graphite and diamond.

7.0 REFERENCES/FURTHER READING

Osei Yaw Ababio (2002) New School Chemistry. Onitsha. Africana-FEP Publishers.