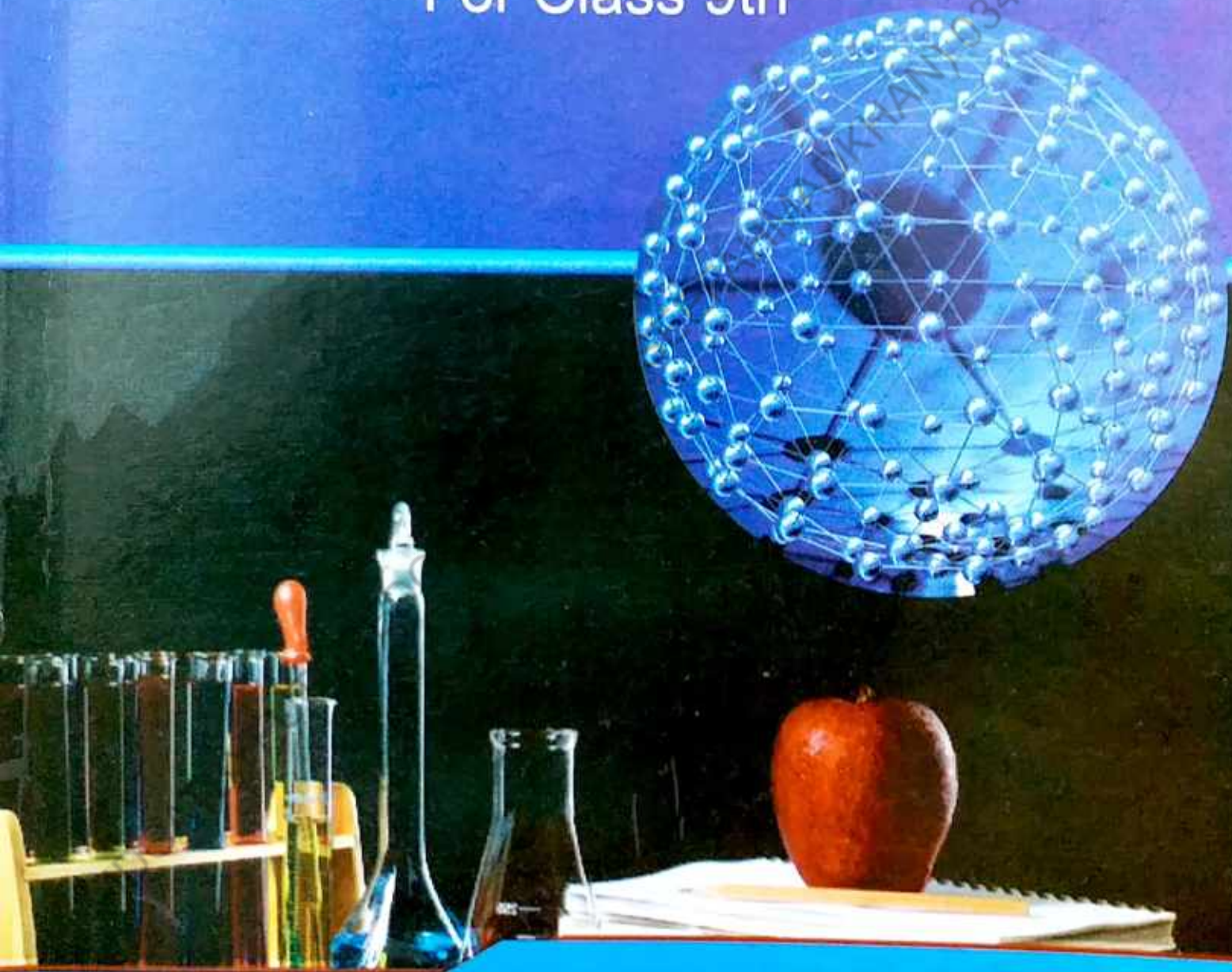


Chemistry

For Class 9th



Balochistan Textbook Board, Quetta.

بِسْمِ اللَّهِ الرَّحْمَنِ الرَّحِيمِ
شروع اللہ کے نام سے جو بڑا مہربان نہایت رحم کرنے والا ہے

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Appeal.

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Preface

With the beginning of new era, the need for reviewing and rewriting of syllabus was felt, so new curriculum for courses has been set by the Ministry of Education in 2006.

The new curriculum offers a radical shift from the traditional curriculum is based upon the multi text book system, creating the students being able to think independently, asking questions and looking for the answers at their own. The understanding of subject must be more developed in them rather than the traditional remembering procedures.

This chemistry '9' book is also the part of new curricula system, and has been written in very simple language, so the learner feel easy in understanding the basic concepts of chemistry, and cope with the challenging demands of the today's world.

Some of the main aims of this book are:

- (i) Writing the chemistry in an easy and approachable manner so the students from the remote and backward areas feel easy to understand the important ideas, concepts of the subject.
- (ii) An overview of the basic principles involved in the subject.
- (iii) Knowledge of the practical approach in the subject, for this purpose many activities and interesting information have been introduced in the book.
- (iv) Most of the topics also explain the role of chemistry tools in our society so the reader be able to understand the importance of the subject and its role for improving the life standard of us.
- (v) There are also solved examples to guide the students how to cope with the problems on the topic.
- (vi) Each chapter is accompanied with an exercise at the end to check the learning ability of the students.

As it is obvious that a text book can do nothing alone and the most important part of studies is the role of teachers and instructors, so the teachers are requested to apply the aims of this book in the light of the SLOs (Students learning outcomes) which are stated at the beginning of each chapter, that describes the time to be allotted to each chapter, tests and assignments of the chapter and the main areas that they should focus on.

At the end we are thankful to all those who helped us in writing this book, and made it possible for us to try to bring a change in the traditional education system.

We also appreciate for further suggestions from readers and educators to improve the quality of this text in future.

Thanking you all in anticipation.

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INTRODUCTION OF CHEMISTRY

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the:

- Basics of Science.
- Divisions of Science
- Importance of Science

TIME ALLOCATION

Teaching periods	= 16
Assessment periods	= 03
Weightage	= 14%

LEARNING OUTCOMES

After reading this chapter, the Student will be able to:

- Identify and provide the examples of different branches of chemistry (Applying)
- Differentiate between branches of Chemistry (Understanding)
- Distinguish between matter and a substance. (Analyzing).
- Define ions, molecular ions, formula units and free radicals. (Remembering)
- Define atomic number, atomic mass, atomic mass unit. (Remembering)
- Differentiate among elements, compounds and mixtures. (Remembering)
- Define relative atomic mass based on C-12 scale. (Remembering)
- Differentiate between empirical and molecular formula. (Understanding)
- Distinguish between atoms and ions. (Analyzing)
- Differentiate between molecules and molecular ions. (Analyzing)
- Distinguish between ion and free radical. (Analyzing)
- Classify the chemical species from given examples. (Understanding)

- Identify the representative particles of elements and compounds. (Remembering).
- Relate gram atomic mass, gram molecular mass and gram formula mass to mole. (Applying)
- Describe how Avogadro's number is related to mole of any substance. (Understanding)
- Distinguish among the terms **gram atomic mass**, **gram molecular mass** and **gram formula mass**. (Analyzing)
- Change atomic mass, molecular mass and formula mass into gram atomic mass, gram molecular mass and gram formula mass. (Applying)

1.0 Introduction to Chemistry



Fig 1.0 Chemistry in action

The Science has been divided into many branches among which the chemistry is of key importance. In this field of Science, the matter, and to some extent energy, the two entities by which our universe is made up of are studied. But the main focus, here is of course upon the matter, and Chemistry deals with the structure of matter on both microscopic and macroscopic level, transformation in matter, in terms of chemical reaction, preparation of different new compounds, the use of these compounds and obviously, also the use of matter which exists already.

The chemistry is in action since the prehistoric era. Burning of wood, purifying the metals like gold for making ornaments, extracting ingredients for usage in drugs, all these involve chemical principles and are known far before our known history.

The chemistry plays very important role in our life. It has a direct application in every sphere of our life, whether concerned to our food we eat, our home we dwell, our cloths, our medicines, our entertaining materials, our educating books, all are direct result of chemistry, and chemistry is related with every discipline of our life. Without the applications of chemistry, the life will be same as of the prehistoric or even below than that, and knowledge and applications of chemistry are the basic constituent of our modern life.

So, all this shows that in order to maintain the living standards of humanity, and meet the new challenges of coming world, the chemistry is very essential to be studied in more depth and it is evident that the 21st century is the century of chemistry.

1.1 Branches of Chemistry

As already discussed that chemistry deals with a vast field of studies, so it has been divided into several branches to make it easier for study.

Here we discuss some of the important branches of chemistry.

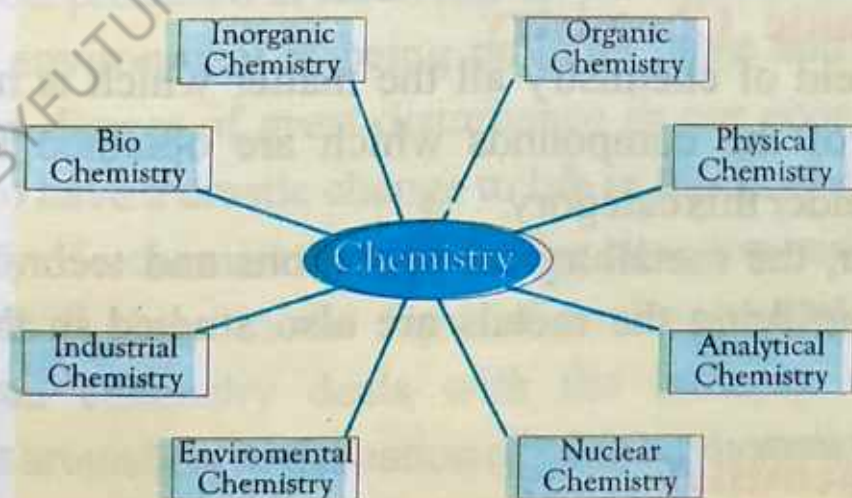


Fig. 1.1 Some branches of Chemistry

1.1.1 Physical Chemistry

This branch deals with the study of physical aspects of matter, such as the colour, melting and boiling points, density etc and all the factors which govern and relate with these characters.

The physical chemistry also deals with the study of atomic structure, role of energy in chemical and physical changes, and reaction kinetics.

Physical chemistry builds the hypothetical basis of structure elucidation of matter, enabling us to understand the basic nature of matter.

1.1.2 Organic Chemistry

The organic chemistry deals with the study of hydrocarbons and their derivatives. Hydrocarbons are the compounds that are composed of carbon and hydrogen e.g. methane, propane, butane, benzene, plastics etc.

Although in the organic chemistry we are primarily concerned with the Carbon and hydrogen elements, but due to the properties of Carbon element, the list of compounds formed are tremendously large, making the subject too heavy and complex. Millions of organic compounds have been in notice at now and each year new compounds are even added in this list.

1.1.3 Inorganic Chemistry

In this field of chemistry all the matter which is not organic is studied. Most of the compounds which are obtained from mineral sources come under this category.

Moreover, the metallurgical operations and techniques that are employed in purifying the metals are also studied in this branch of chemistry.

1.1.4 Biochemistry

The word bio stands for life, so this branch of chemistry deals with the

matter which is directly obtained from the living organisms. Thus it deals with the compounds like carbohydrates, proteins, lipids and others.

1.1.5 Industrial Chemistry

As the name indicates, the industrial chemistry is the chemistry of industries. All the materials produced by the industries are discussed here. This branch is important in the sense that all the matter produced in the laboratory must come in our life usage to make our life easier and comfortable, which is done by industries. e.g. the cement to make our homes, the fertilizers to grow up the amount of our food, the medicines to keep us healthy, the fabric industry to give us the cloths and almost all of the compounds which we use in our life is the result of research and comes in our reach by their production in the factories, which is dealt with the industrial chemistry.

1.1.6 Nuclear Chemistry

This branch of chemistry deals with the study of nucleus of an atom. This gives us the information for the production of energy from matter by nuclear reaction, weapons of mass destruction, transformation of elements, and preparation of nuclear medicines used in radiotherapy.

1.1.7 Environmental Chemistry

As the world is developing more and more and we are using the advanced tools produced in factories, the fossil fuels usage is increasing rapidly, the environment is being polluted more and more. This has increased the danger of great disturbance in our ecosystem, which is considered to have a drastic change to life in this planet, so it is also the responsibility of a chemist to improve this situation and to meet all such dangers. This is done in the field of environmental chemistry. The environmental chemistry deals with the building of life friendly environment around us, and creation of such ecosystem that is inevitable for our survival in the universe.

1.1.8 Analytical Chemistry

The basic aim of chemistry is the production and usage of compounds, which can be done only when the composition and structure of an existing compound is known, which is the main aim of analytical chemistry, which deals with the determination of structure and composition of matter i.e. in this branch of chemistry we analyze a chemical compound and determine the molecular formulae of the chemical compounds.

1.2 Basic Definitions

Before we proceed further, let us understand some of the key and basic terms used frequently in the chemistry.

1.2.1 Elements, Compounds and the Mixtures

The universe is made up of two basic entities, the matter and the energy. Chemistry deals with the matter, so we will focus on this exclusively.

The matter is the substance that has mass. It is distributed all round us and this matter is classified into the two main categories, i.e. the pure substances and the mixtures. A pure substance is that which shows uniform properties every time, for example melting point and boiling point of such substances are always same, while in a mixture the quantity of constituents can vary, e.g. in the mixture of sugar syrup the amount of sugar can be less or greater. The pure substances are further divided into two sub divisions i.e., the elements and the compounds, this classification is shown in the following flow diagram. These divisions will be discussed in the subsequent sections in detail.

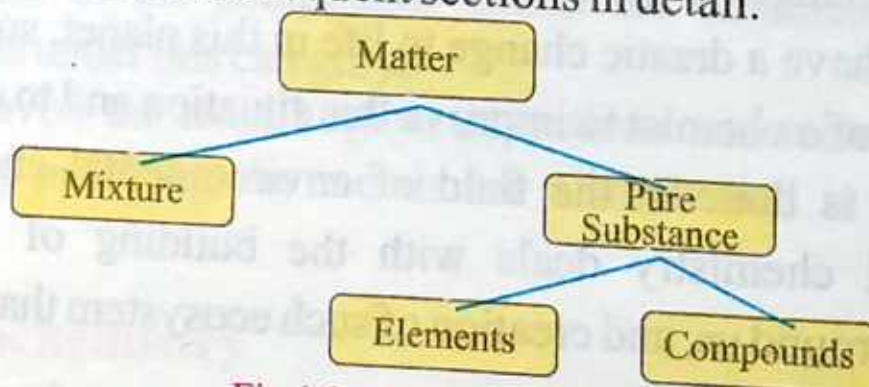


Fig 1.2 Divisions of matter

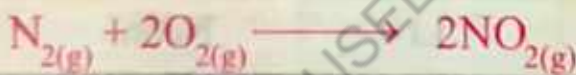
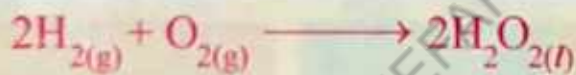
1.2.1.1 Elements and their Symbols

The element is defined as the sample of a pure substance in which all the atoms are alike.

For example, if we take a rod of iron metal, it consists of same type of atoms, so iron is an **element**, and on the other hand, a piece of wood consists of different types of atoms, so wood is not an element.

The elements found in all three states of matter i.e., the gas (Nitrogen, Oxygen, Hydrogen etc.), liquid (Bromine and Mercury) and solid (Aluminum, Sodium, Copper etc.).

Different elements show different physical and chemical properties, for example, if we take the above described three gaseous elements i.e., N_2 , O_2 and H_2 , all these have different boiling points and show different chemical products on reactions, such as:



In Chemistry, the atoms are identified by their atomic numbers, (The atomic numbers are discussed in detail in the section 1.2.2.1). So an element can be defined as the sample of pure substance that has all the atoms with same atomic numbers. (e.g. in the example of iron rod, all atoms have atomic number = 26). In our daily life we normally deal with the atoms in the form of elements, e.g. O_2 , Au, Ag, Cu and many more. Usually we use a short name for an element called the Symbol. A symbol is chosen normally by taking the first letter, written as the capital letter of the English name of that element, e.g. 'H' is used for hydrogen, 'O' for oxygen, 'C' for carbon.

Here is a table (Table 1.1) showing symbols and English names of some of the common elements.

Name of Element	Symbol	Name of Element	Symbol
Hydrogen	H	Fluorine	F
Oxygen	O	Iodine	I
Nitrogen	N	Boron	B
Carbon	C	Sulphur	S

Table: 1.1: Some elements whose symbols are derived from their english names

Sometimes the first and second letter from the English name are used, e.g. 'He' for helium, 'Ca' for calcium. When two letters are used, the second letter must be written in small alphabets. Here are some examples of such elements (Table 1.2).

Name of Element	Symbol	Name of Element	Symbol
Lithium	Li	Neon	Ne
Beryllium	Be	Nickel	Ni
Aluminum	Al		

Table: 1.2: Symbols of some elements having first and second letters from English names

In some cases, first and any other letter is used, as described in the table 1.3.

Name of Element	Symbol	Name of Element	Symbol
Chlorine	Cl	Cadmium	Cd
Chromium	Cr	Magnesium	Mg

Table: 1.3: Symbols of some elements having first and a prominent letter from their English names

Some elements which are known from prehistoric era, are represented by the symbols of their Greek names, e.g. (see table 1.4).

English Name of Element	Greek/Latin Name of Element	Symbol
Gold	Aurum	Au
Silver	Argentum	Ag
Mercury	Hydrargyrum	Hg
Sodium	Natrium	Na
Copper	Cuprum	Cu

Table: 1.4: Symbols of some elements derived from their Greek/Latin names

So far 118 elements have been discovered of which 92 are found naturally while rest are synthesized artificially in laboratories, these elements found in both the free and in combined states.

Interesting To Know

In our solar system top 3 elements are:

H = 74%, He = 24%, and O = less than 2%

In the earth crust top 3 elements are:

O = 46.6%, Si = 27.7% and Al = 8%

In our atmosphere top 3 elements are:

N = 78%, O = 21% and Ar = 0.9%

In our body top 3 elements are:

O = 65%, C = 18% and H = 10%

1.2.1.2 The Compound

When different elements combine, they form a compound, thus generally a compound has two or more than two types of elements, example of such substances include water, table salt, glucose.

In a compound the proportion of elements is fixed and can not be changed, e.g. in H_2O the proportion of H atoms to O atom is 2:1, and this ratio is fixed otherwise the compound will be changed, for example in above example of hydrogen and oxygen, if the ratio between these two elements changes 2:2 the result is formation of H_2O_2 (hydrogen peroxide) which is totally different from H_2O (water). It is also notable that the formation of a compound is a chemical change and can not be easily reversed back to its constituent elements.

A compound which is formed of different type of elements shows properties quite different from the elements that form it, e.g., water which is a liquid is formed of two gases hydrogen and oxygen, the hydrogen burns while the oxygen helps in burning process, on the other hand the compound water extinguishes the burning objects.

A compound is represented by its short name called its formula or molecular formula. A formula shows all the elements present in the compound with the number of each type of atoms.

A compound can be divided into two sub types, the ionic and the covalent. Ionic compounds consist of a large number of particles forming a huge crystal structure, while the covalent exists in the molecular form.

The mixture is a combination of two or more substances (Elements or compounds) in such a manner that each component shows its own characters, i.e. they are not combined chemically.

The common examples of such mixtures include sugar syrup, air and fabric woven of yarns combining various colors.

The mixtures do not have the uniform composition and properties, and the components of a mixture can easily be separated by simple

Formula	Name of Compound
CO_2	Carbon Dioxide
NaCl	Sodium Chloride
$\text{C}_{12}\text{H}_{22}\text{O}_{11}$	Sugar (Sucrose)
$\text{C}_6\text{H}_{12}\text{O}_6$	Glucose
$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	Sodium Carbonate
NaHCO_3	Sodium Bicarbonate
CaCO_3	Calcium Carbonate
SiO_2	Silicon Dioxide
$(\text{C}_6\text{H}_{10}\text{O}_5)_n$	Cellulose

Table:1.5: Some Compounds and their Formulae



Fig 1.3 Fabric--A mixture of yarns

physical methods e.g. the sugar from the syrup can be separated simply by evaporating the water from the syrup.

The Table 1.6 that shows some basic differences between compound and mixture.

Activity: 1.1

How can you separate the components of a mixture of table salt (NaCl) solution?

Try to separate these in the laboratory with the help of your teacher.

S/No	Compound	Mixture
i	They are pure substances.	They are impure.
ii	Compound is formed by chemical reaction of atoms/elements.	It is formed by the mixing of substances.
iii	Its properties are entirely different from the component substances by which it is made up of, e.g water is quite different from its building elements i.e. hydrogen and oxygen. $\text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{H}_2\text{O}_{(l)}$	This shows the properties of its constituent.
iv	It cannot be converted back to its original constituents easily by physical methods.	A mixture can be easily converted back to its parent components by simple physical methods.
v	The constituents have fixed composition of the elements.	The constituents here do not have fixed composition.
vi	They show fixed melting and boiling points.	They do not show fixed melting and boiling points.

Table: 1.6: Some basic differences between compounds and mixtures.

1.2.2 Atoms, the Building Blocks of Matter

The name atom has been derived from the Greek word "atomos" which means indivisible or which cannot be cutted further. This originated because scholars of ancient Greek believed all substances are made up of tiny indivisible particles which they named atomos.

In the modern era John Dalton proposed his atomic theory and named the basic constituent of matter as atoms.

Although atoms are the smallest unit of matter, and hence it is the building block of matter, but still they are not indivisible now. So, now a day's atom is defined as the smallest unit of matter which takes part in chemical reactions and shows all properties of an element. e.g. if a piece of gold is divided into two fragments, each fragment will show the characters of gold, the gold is identified by its peculiar physical and chemical properties. When it is divided in more fragments again each fragment is the representative of chemical properties of gold, the division is continued until the last fragment achieved is called an atom.

Atoms are so small that they can't be seen by the naked eye or any traditional microscope, a recent technique called scanning tunneling microscopy (STM) have made atoms visible up to a little extent, An image of gold foil by STM is shown here in figure 1.4.

The atoms normally do not exist freely in nature, though some exceptions are there, e.g. all the members of group-18 in periodic table, called the noble gases exists as single atom, but their existing atoms are rendered as mono atomic molecules.

Although the atoms are invisible,

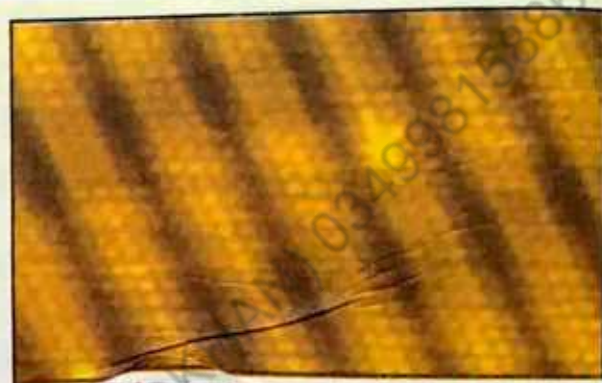


Fig 1.4 An image of gold foil by STM

Interesting Information

The size of nucleus is very small as compared to volume of an atom and is about 10,000 time less than the size of an atom.

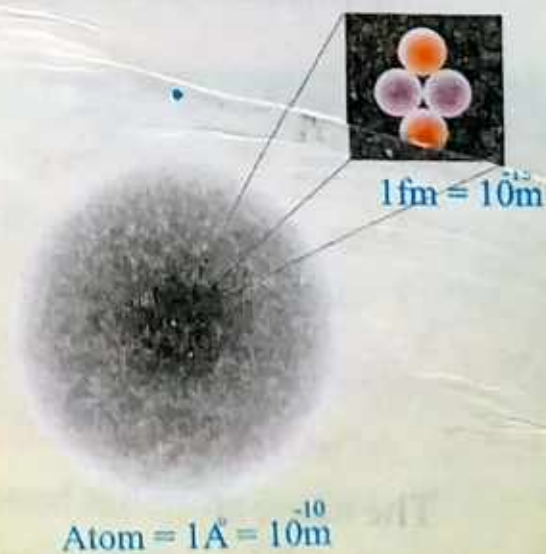


Fig 1.5
The modern view of an Atom

but by using special techniques and experiments, Scientists have created a picture of atom, that defines the atom as being buildup of more smaller particles, these particles are called electron, proton and neutron, of these three, the proton and the neutron lie in the center of atom, this center is termed as the nucleus of atom.

The electron revolves around the nucleus in definite areas called shell.

The electrons and the protons are charged particles and electron bears unit negative, while proton bears unit positive charge. The number of electrons and protons are equal in the atoms so the atom is electrically neutral. A simple picture of an atom described is thus given as in figure 1.5.

This picture 1.5 is showing the basic characteristics of the atom, the darker area represents the density of electron, and a comparison of the sizes of atom and nucleus of atom, note that an atom here has the size of 10^{-10} meters ($1\text{\AA}$) while the size of the nucleus is 10^{-15} meters (1 fm).

The tiny center is the nucleus, and the cloudy area is the region of revolving electrons, the strength of clouds show the density of electron charge distribution.

1.2.2.1 Atomic Number and the Mass Number

In order to identify the atoms, each atom has been allotted a number, called **atomic number** (symbolized as **Z**). (This is same just like the students in a class are allotted the roll numbers to identify them easily). The atomic number is based upon the number of protons in the nucleus of the corresponding atom, e.g. H has 1 proton so its atomic number is also 1, the oxygen has 8 protons in its nucleus, so its atomic number is also 8. The concept of giving identity to atoms is of great importance, as nothing can be described about its properties unless it is identified separately.

The **mass number** (symbolized as **A**) is the total number of mass building particles in an atom, the mass building particles of atom are the neutrons and the protons. e.g. the 'N' has 7 protons and 7 neutrons so its mass number is '14', whereas, the 'C' has 6 protons and 6 neutrons, so its mass number is 12. The mass of an atom has great effect upon the physical and sometimes even the chemical properties of the atoms.

Table 1.7 shows the atomic number and mass number of some of the

Element	Number of Protons	Number of Neutrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1
Helium	2	2	2	4
Carbon	6	6	6	12
Nitrogen	7	7	7	14
Oxygen	8	8	8	16
Sodium	11	12	11	23
Calcium	20	20	20	40

Table 1.7: Some common elements with their atomic and mass numbers

The modern science has enabled us to increase or decrease the number of particles in the nucleus and thus we can change the type of atoms.

Naturally there exist 90 different type of atoms while many have been synthesized in the laboratories.

1.2.3 Relative Atomic mass and the Atomic mass unit

The relative atomic mass of an atom is its mass when compared with the light isotope of carbon i.e. $C^{12.000000}$.

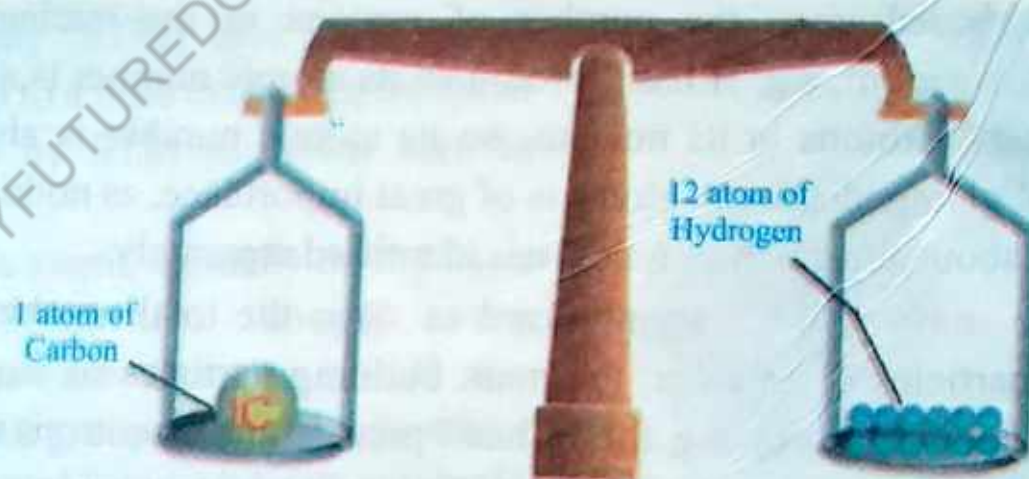


Figure 1.6: A hypothetical balance to show the concept of phenomenon of the comparing the mass of Hydrogen atoms with standard $C^{12.000000}$

The reason that relative unit is used rather than the absolute mass unit is that it is experimentally much easier and more precise to compare masses of

atoms and molecules than to measure their absolute masses. Masses are compared with a mass spectrometer. The atomic and molecular masses are measured in the unit of atomic mass unit (a.m.u), or sometimes as unified atomic mass unit, abbreviated as 'u'. It is defined to be one twelfth of the mass of an atom of the C^{12} nuclide. In kilograms it is written as:

$$1 \text{ a.m.u} = 1.6605402 \times 10^{-27} \text{ kg}$$

$$1 \text{ a.m.u} = 1.6605402 \times 10^{-24} \text{ grams}$$

In the periodic table, the atomic mass listed is the weighed average of all of the isotope masses for that element. In this weighed average, the weights are the percent abundance that each isotope occurs in nature. For example, if we analysed a lump of pure carbon from the planet Earth, it is observed that:

98.9 % of all carbon atoms are $C^{12.0000}$ atoms,
and

1.11 % of all carbon atoms are $C^{13.0039}$ atoms.

So the weighted average mass of carbon is:

$$\frac{(98.9 \times 12.000000) + (1.11 \times 13.0039)}{100} = 12.01 \text{ a.m.u}$$

In the same way, the natural abundance of $Cu^{62.93}$ is 69.09 % and for $Cu^{64.93}$ is 30.19 % respectively.

Thus the average atomic mass for copper is calculated as:

$$\frac{(69.09 \times 62.93) + (30.91 \times 64.93)}{100} = 63.55 \text{ a.m.u}$$

These average values of mass of respective elements are written in a periodic table along with the symbol of that element.

1.2.4 Empirical and the Molecular Formulae

The constitution of a compound can be represented in many ways, this representation is called formula of the respective compound. There formulas are written in different ways.

1.2.4.1 The Empirical Formula

The empirical formula represents the simple ratio of the atoms present in a compound, e.g. in glucose ($C_6H_{12}O_6$) the simplest ratio of atoms is 1:2:1

i.e. for each atom of Carbon there are two Hydrogen atoms and 1 Oxygen atom.

So the empirical formula of glucose is " CH_2O ".

In water (H_2O) the simplest ratio of Hydrogen to Oxygen atoms is 2:1 so the empirical formula of water is " H_2O ", remember this is also the molecular formula of water.

Activity: 1.2

What are the empirical formulae of following, Benzene (C_6H_6), Sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$), Carbondioxide (CO_2), and Cellulose ($\text{C}_6\text{H}_{10}\text{O}_5$)_n

1.2.4.2 The Molecular Formula

In Chemistry a short name is used to represent the constituents of a compound, this short name is called the formula of that specific compound, e.g. for water the formula is " H_2O ".

In the table 1.8 some formulae of common compounds are written:

As it is clear from the above table that a formula shows two basic information about that specific compound, i.e.

- The composition of the compound, stating types of the elements present in that compound.
- The number of atoms of each element present in that compound.

Name of Compound	Formula
Carbon dioxide	CO_2
Table salt (Sodium chloride)	NaCl
Sugar (Sucrose)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}$
Glucose	$\text{C}_6\text{H}_{12}\text{O}_6$
Washing soda (Sodium carbonate)	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Baking soda (Sodium bicarbonate)	NaHCO_3
Lime stone (Calcium carbonate)	CaCO_3
Sand (Silicon dioxide)	SiO_2
Starch	$(\text{C}_6\text{H}_{10}\text{O}_5)_n$

Table 1.8: Some common compounds with their formulae

Sometimes a compound consists of a large number of repeating unit called polymer e.g. Cellulose or Starch $(\text{C}_6\text{H}_{10}\text{O}_5)_n$

and Polyvinyl Chloride (P.V.C) $\left(\text{---CH}_2\text{---CH}_2\text{---} \right)$
 $\quad \quad \quad |$
 $\quad \quad \quad \text{Cl}$

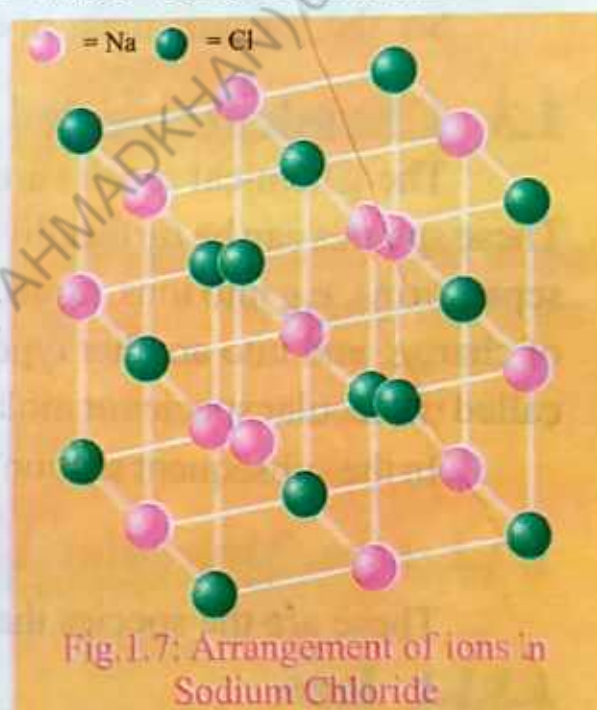
Such enormously giant structures are called polymers.

1.2.4.3 The Formula Unit

Sometimes a molecule exists as a large entity of particles, for ease the simplest ratio of the formulae of such huge structures is written, this simplest ratio is termed as the unit of their formula or the formula unit, e.g Sodium chloride exists in such manner, and we write "NaCl" as its formula unit.

The NaCl consists of a large number of cations, The Na^+ ions, and the anions, Cl^- . Thus the NaCl exists as a large entity of crystal structure, in which the cations are surrounded by the anions and vice versa, this is shown here in the figure 1.7.

The simplest ratio of ions here is 1:1, so its formula unit is NaCl, in the same manner the ratio of ions in calcium bromide is 1:2, so the formula unit of Calcium bromide is CaBr_2 .



1.2.5 The Molecular Mass and the Formula Mass

Molecular mass of a compound is the total mass of all the atoms present in that particular molecule. This is also known as the molar mass and is calculated easily by taking sum of masses of all the atoms present, e.g. for the molar mass of water,



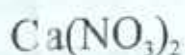
$$\text{H} = 1 \times 2 = 2$$

$$\text{O} = 16 \times 1 = 16$$

$$\text{Sum of these two is } 2 + 16 = 18$$

So, the molar mass of water is 18 a.m.u.

The formula mass is also calculated in the same manner, e.g. the formula mass of Calcium nitrate, $\text{Ca}(\text{NO}_3)_2$ is:



$$\text{Ca} = 40 \times 1 = 40$$

$$\text{N} = 14 \times 2 = 28$$

$$\text{O} = 16 \times 6 = 96$$

Sum of these three is

$$40 + 28 + 96 = 164$$

So, the formula mass of calcium nitrate, $\text{Ca}(\text{NO}_3)_2$ is 164 a.m.u.

Activity: 1.3

Calculate the mass of following compounds,

NaCl , CO_2 , Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$), and $\text{Al}(\text{OH})_3$

1.3 Chemical Species

The Chemical species are the building blocks of chemical compounds. These species can be further divided into different types on the basis of charge separations, e.g. into ions, molecular ions and free radicals that bear some sort of charge, and into another types of stability which do not have such charge, called molecular species or molecules.

In the subsequent section we explain each of these types in detail.

1.3.1 Ions, Molecular Ions & Free Radicals

These are the species that bear some sort of charges on them:

1.3.1.1 Ions

An ion is an atom or group of atoms that has lost or gained one or more electrons.

The simplest ions are the proton (a hydrogen ion, ${}^1\text{H}^+$, positive charge), and alpha particle (helium ion, ${}^4\text{He}^{2+}$, consisting of two protons and two neutrons). The ions are of two types by the charges, negatively charged ions are called anions, and the positively charged are called cations, these names come from the fact that the negatively charged ions move towards the anode in the process of electrolysis, while the positively charged ions move towards the cathode so named cations. The table 1.9 shows some common monatomic cations.

Common Cations

Common Name	Symbol	Common Name	Symbol
Aluminium	Al^{3+}	Nickel	Ni^{2+}
Calcium	Ca^{2+}	Magnesium	Mg^{2+}
Chromium (II) (Chromous)	Cr^{2+}	Chromium(III) (Chromic)	Cr^{3+}
Copper(I) (Cuprous)	Cu^{+}	Copper(II) (Cupric)	Cu^{2+}
Lead(II) (Plumbous)	Pb^{2+}	Lead (IV) (Plumbic)	Pb^{4+}
Silver (Argentum)	Ag^{+}	Potassium (Kalium)	K^{+}
Hydrogen	H^{+}	Sodium (Natrium)	Na^{+}
Iron (II) (Ferrous)	Fe^{2+}	Iron (III) or Ferric	Fe^{3+}

Table 1.9 List of Common Cations

The ions that have the negative charge are called anions, (The name comes from the fact that in an electrolytic cell the negatively charged particles moves towards the anode, so anion— anode)

The table 1.10 lists some important anions:

Common anions

Common Name	Symbol	Common Name	Symbol
Bromide	Br^{-}	Iodide	I^{-}
Chloride	Cl^{-}	Nitride	N^{3-}
Fluoride	F^{-}	Oxide	O^{2-}
Hydride	H^{-}	Sulphide	S^{2-}

Table 1.10: Some Important Anions

Although the existence of ions independently is hard due to their greater instability, but there are molecules that have combination of both positive and negative ions in a manner that both types neutralize and as a whole the compound is electrically neutral, as described above in the case of NaCl.

1.3.1.2 Molecular Ions

An ion consisting of multiple atoms is called a polyatomic ion, or molecular ion, e.g. NH_4^+ ion.

The molecular ions also may be of two types, the cations and the anions. Following tables show some common molecular cations (Table 1.11) and molecular anions (Table 1.12).

Polyatomic or Molecular Cations	
Ammonium	NH_4^+
Hydronium	H_3O^+
Nitronium	NO_2^+

Table: 1.11: Some Polyatomic Cations

Molecular anions	
Carbonate	CO_3^{2-}
Bicarbonate / Hydrogen Carbonate	HCO_3^-
Dichromate	$\text{Cr}_2\text{O}_7^{2-}$
Nitrate	NO_3^-
Phosphate	PO_4^{3-}
Hydrogen Phosphate	HPO_4^{2-}
Sulphate	SO_4^{2-}

Table: 1.12: Some Polyatomic Anions

Naming Ionic Compounds

The names of ionic compounds are written by listing the name of the positive ion first, followed by the name of the negative ion e.g.

NaCl Sodium chloride

$(\text{NH}_4)_2\text{SO}_4$ Ammonium sulfate

NaHCO_3 Sodium Bicarbonate / Sodium Hydrogen Carbonate

KClO_3 Potassium chlorate

1.3.1.3 Free Radicals or Radicals

The radicals or free radicals are atomic or molecular species with unpaired electrons.

Due to these unpaired electrons the radicals are highly reactive, and are produced only during the course of a chemical reaction for a very short time e.g. the chlorine radical is formed by the dissociation of chlorine molecule by light energy and is readily react with the neighboring species.



The electron of the free radical is written by a dot upon the symbol of

free radical, as shown above in the radical of Cl. Radicals play an important role in combustion, atmospheric chemistry, polymerization, plasma chemistry, biochemistry and many other chemical processes, including human physiology. For example, superoxide and nitric oxide regulate many biological processes, such as controlling vascular tone.

The upper atmosphere of earth contains radicals which are formed there by the action of UV radiations that is coming from sun and thus the energy of this highly injurious UV radiation is blocked.

1.3.2 Molecular Species

A molecule is the smallest stable form of matter and is composed of generally two or more atoms. The constituting atoms may be same or different from each other, e.g. O_2 is a molecule having both of its atoms same while the CO_2 is a molecule having different types of atom. Table 1.7 shows some molecules and their formulae.

On the basis of its composition molecules are classified into different sub types. These types are:

- (a) **Monoatomic Molecules:** Although the molecule is generally considered of being formed by the combination of two or more atoms, but even a molecule could be composed of a single atom, this is because the main difference between the molecule and an atom arises from its stability, molecules are the species which are stable, so even a single atom which is stable could be considered as the molecule, important examples of such monoatomic molecules include the He, Ar, Ne and other noble gases.
- (b) **Diatomic molecules:** When a molecule consists of two atoms it is called diatomic molecule, for example H_2 , O_2 , HCl , CO etc.
- (c) **Polyatomic molecules:** These are the molecules that are composed of more than two atoms, like $C_6H_{12}O_6$, HNO_3 , $CuSO_4 \cdot 5H_2O$ etc.

In addition to this the molecule may also be classified as **homo atomic molecules** (homo-similar) e.g. O_3 , O_8 , which are composed of same type of atoms and **heteroatomic molecule**, (hetero-different) e.g. NH_3 , HNO_3 , H_2SO_4 etc, which are composed of different types of atoms.

1.4 Mole and the Avogadro's Number

1.4.1 The Mole (mol)

In Chemistry we often make reactions by combining two or more reactants, and for this we must have to take the reactants in correct amount, otherwise the product(s) may not be formed. (this is very similar to the behavior that in making a good tea one have, to take all the ingredients in correct amount). But in actual we are unable to count and take atoms of reacting substances in desired quantities.

So, to solve this, we use the concept of mole, which states the ratios of given substances in much easier way in the units of grams, and this makes it very easy and perfectly practical to measure the amounts of reactant(s) or product(s).

According to this concept the atomic mass or in case of molecules the molecular mass is taken in the unit of grams instead of atomic mass unit and hence it becomes easy to measure these using ordinary instruments (balances).

Thus the atomic mass or molecular mass expressed in terms of grams is called 1 mole of that substance, (or sometimes it is also called gram atom or gram molecule or gram formula mass), as shown in figure 1.8.

Here are some examples that elaborates this concept,

'H' has '1' atomic mass so 1g of it is one mole. 'O' has '16' a.m.u atomic mass so its 16g = 1 mole

'H₂' has '2' atomic mass so 2g of it is one mole. 'O₂' has '32' a.m.u atomic mass so its 32g = 1 mole H₂O has 18 a.m.u molecular mass and its 18g = 1 mole

And

36 g of water = 2 moles

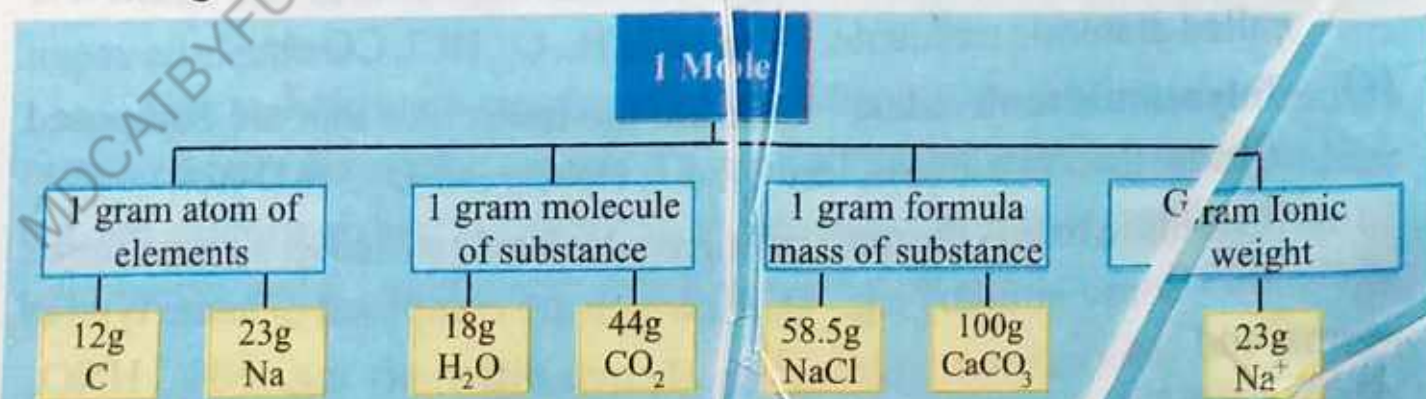


Figure 1.8: Mole—Mass Relationship

A simple relation that is used to interconvert the mole—mass relation, which given by:

$$\text{Mole} = \frac{\text{Given Mass}}{\text{Atomic / Molecular Mass / Mass Formula / Ionic Mass}}$$

The concept of mole is really very helpful for the Chemists and is applied widely.

Example: Calculate number of moles in 250g of glucose.

Solution: As the moles of glucose can easily be determined by using the formula:

$$\text{Mole} = \frac{\text{Given Mass}}{\text{Atomic or Molecular Mass}}$$

The molecular mass of glucose can be calculated by the help of its molecular formula, which is $\text{C}_6\text{H}_{12}\text{O}_6$.

Here,

$$\text{C} = 6 \times 12.01 = 72.06$$

$$\text{H} = 12 \times 1.0076 = 12.0912$$

$$\text{O} = 6 \times 16.00 = 96$$

$$\text{Total mass} = 72.06 + 12.0912 + 96 = 180.15$$

By inserting the values in the above formula, we get

$$\text{Mole} = \frac{250}{180.15} = 1.39 \text{ moles}$$

1.4.2 Avogadro's Number (N_A)

Avogadro practically found that one mole of any substance contains 6.02×10^{23} particles (atoms/ molecules) i.e 1 mole of 'O' has 6.02×10^{23} atoms and 1 mole of H_2O has 6.02×10^{23} molecules. This number is called Avogadro's Number.

This is to be remembered that in atomic species the Avogadro's Number represents



Amedeo Avogadro
(Italian Scientist)
1776 — 1856

number of atoms, e.g., Helium (He) gas which has atomic mass '4', 4gm helium will have 6.02×10^{23} atoms.

In case of molecular species this Avogadro's Number represents number of molecules e.g., 32g of oxygen (O_2) will have 6.02×10^{23} molecules.

Example:

Calculate total number of atoms in 1.0 g of water.

Solution

As 1 mole of water has 6.02×10^{23} molecules = 18g

1 mole of H_2O = 18g = 6.02×10^{23} molecules

So, 18g of H_2O has 6.02×10^{23} molecules and $6.02 \times 10^{23} \times 3$ atoms (as one molecule of H_2O has 3 atoms)

And 1g of H_2O will have:

$$\frac{6.02 \times 10^{23} \times 3 \text{ atoms}}{18g H_2O} \times 1g H_2O$$

$$= 1.0033 \times 10^{23} \text{ atoms}$$

Activity: 1.4

Calculate the number of atoms in 100g of CO_2 , and Glucose ($C_6H_{12}O_6$).

SUMMARY OF THE CHAPTER

- ❖ **Chemistry** is the basic branch of Science that deals with the study of matter.
- ❖ The important **branches of Chemistry** are Physical Chemistry, Organic chemistry, Inorganic Chemistry, Biochemistry, Industrial Chemistry, Nuclear Chemistry, Environmental Chemistry, and Analytical Chemistry.
- ❖ The **Element** is a sample of pure substance that has all atoms with same atomic number.
- ❖ A **Compound** is a sample of pure substance that has two or more than two types of elements.
- ❖ A **mixture** is not a pure substance and is consists of by mixing two or more than two pure substances, the quantities of the components of a mixture could be changed.
- ❖ Atomic number is the identity number of an atom or element. It is equal to the number of protons found in the nucleus of atom.
- ❖ The sum of number of protons and neutrons in an atom is the **Atomic mass number**.
- ❖ **Relative atomic mass** is the mass of an atom or molecule which is obtained by comparing the mass of an atom or molecule with the standard $C^{12.0000}$.
- ❖ The atoms are very small and their mass is measured in a very small unit called **atomic mass unit (a.m.u)**, in grams it is equal to:
$$1 \text{ a.m.u} = 1.6 \times 10^{-24} \text{ g}$$
- ❖ Empirical formula is the formula which shows the simplest ratio of atoms in a molecule.
- ❖ **Molecular formula** is the formula of a chemical compound and it shows the exact number and type of atoms present in that compound.
- ❖ **Molecular mass** is the sum of masses of all atoms in a molecule.
- ❖ **Ions** are the charged atoms. Ions are **anions** when they bear negative charge, and **cations** when they have positive charge.

- ❖ Radicals or free radicals are the species (atoms or group of atoms) that have an unpaired electron.
- ❖ The molecules are said to be **monoatomic** when they are composed of single atom, **diatomic** when they are composed of two atoms and **polyatomic** when they are formed by more than two atoms.
- ❖ The molecules are **homoatomic** if they are composed of same type of atoms and **Heteroatomic** when they are formed by different types of atoms.
- ❖ When atomic or molecular masses are represented in terms of grams, mass obtained is termed as **1 mole**.
- ❖ The number of particles in one mole is called **Avogadro's number** (N_A), this number is equal to 6.0225×10^{23} .

EXERCISE

Q.1: Fill in the blanks.

- i) Physical chemistry deals with of matter.
- ii) The branch of chemistry in which the nucleus of atom is studied is called chemistry.
- iii) A sample of pure substance having all atoms is called an element.
- iv) For the mass comparison of atoms, the standard atom chosen is
- v) The number 6.0225×10^{23} is called
- vi) moles of water constitute 36g.
- vii) The process of formation of ions is called
- viii) The Atomic mass unit is expressed by the symbol
- ix) The molecular formula shows number of atoms.
- x) Glucose is type of molecule.

Q.2: Tick the correct answer from the given list.

- i) The smallest component of an element that still retains properties of the element is called

- (a) electron (b) proton (c) neutron (d) atom
- ii) The atom is a Greek word meaning
 (a) indivisible (b) small particle
 (c) which enter in reactions (d) basic particle
- iii) The branch of chemistry which deals with the determination of molecular formula is called
 (a) Nuclear chemistry (b) organic chemistry
 (c) Analytical chemistry (d) Industrial chemistry
- iv) Empirical formula shows the
 (a) ratio of atoms (b) ratio of molecules
 (c) exact number of atoms (d) exact number of molecules
- v) Molecular mass is calculated by
 (a) Counting number of atoms of C^{12}
 (b) Taking sum of the masses of atoms in the molecule
 (c) By taking the sum of protons and neutrons number
 (d) None of above
- vi) The value of a.m.u in grams is
 (a) 1.6×10^{23} (b) 1.6×10^{27}
 (c) 1.6×10^{-23} (d) 1.6×10^{-24}
- vii) The atoms can exists freely in nature
 (a) Always (b) never
 (c) some exist and some do not (d) none of these is true
- viii) The formula unit is used for the
 (a) ionic substances (b) molecular substances
 (c) Mixture (d) none of these
- ix) The gram atomic or molecular mass is called
 (a) one mole (b) two moles
 (c) three moles (d) 4 moles
- x. The concept of Avogadro's number is used to calculate the
 (a) number of atoms (b) number of molecules
 (c) number of ions (d) all of these

Q3. Answer the following questions in short.

- i) Why understanding of chemistry is important to us?

- ii) Give a short note on classification of matter.
- iii) What is meant by the term Atomic Number?
- iv) How will you differentiate between the Atomic mass number and the Relative atomic mass?
- v) Distinguish among atoms and ions with examples.
- vi) Explain with an example how the molecular mass is calculated.
- vii) What are the different types of matter found in nature?
- viii) What is the mole?
- ix) What is meant by a.m.u?
- x) Describe with example how simple inorganic compounds are named.

Q4. Answer the following questions with reasoning.

- i) How Avogadro's Number and the mole are related with each other?
- ii) Comment that concepts of vital for chemists.
- iii) Distinguish between ion and the free radical?
- iv) What is the difference between empirical formula and the formula unit?
- v) Why C^{12} is taken as the standard for calculating atomic or molecular masses?

Q5: Nicotine a deadly poison which is found in tobacco has the formula $C_{10}H_{14}N_2$. An average cigarette contains 1mg of nicotine. What will be the number of molecules in this quantity?

Q6: Calculate the molar mass of following,

- | | | |
|----------------|--------------------------|----------------|
| (a) K_2SO_4 | (b) $MnCl_2$ | (c) $Al(OH)_3$ |
| (d) $Fe(NO)_3$ | (e) $CuSO_4 \cdot 5H_2O$ | |

Q7: Write down the molecular formula of following compounds.

- | | |
|--------------------------|------------------------|
| (a) Ferrous sulphate | (b) Ferric sulphate |
| (c) Copper(II) sulphide | (d) copper(I) sulphite |
| (e) Ammonium bicarbonate | |

Q8: What are the mass and mole of sulphur in 50g of sulphuric acid?

Q9: Define Chemistry? How Chemistry effects our life?

Q10: Why the Science is divided into various branches? Discuss the ing of

branching of chemistry in detail?

Q11: What is the difference between the Element, mixture and the compound?

Q12: What do you understand by the term symbol of an element? Explain various rules for assigning symbols to an element?

Q13: What is the molecular formula of a compound? How the formulae are written?

Q14: Discuss the concept of mole in chemistry? How it is related with the Avogadro's number?

STRUCTURE OF ATOMS

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the:

- Basics of Atom.
- Basics of sub atomic particles
- Properties of charged particles

TIME ALLOCATION

Teaching periods	= 16
Assessment periods	= 03
Weightage	= 14%



LEARNING OUTCOMES

After reading this chapter, the Student will be able to:

- Describe the contribution that Rutherford made to the development of the atomic theory. (Applying)
- Describe the contribution of Bohr made in development of atomic structure. (Applying)
- Explain how Bohr's theory differs from Rutherford's theory. (Analyzing)
- Describe the structure of an atom including the location of the proton, electron and neutron. (Understanding)
- Define isotope. (Remembering)
- Compare isotopes of an atom. (Analyzing)
- Draw the structures of different isotopes from mass number and atomic number. (Applying)
- State the importance of isotopes in various fields of life. (Remembering)
- Describe the presence of sub shells in a shell. (Understanding)
- Distinguish between shells and sub shells. (Understanding)
- Write electronic configuration of first 18 elements in the Periodic table. (Applying)

2.0 Introduction

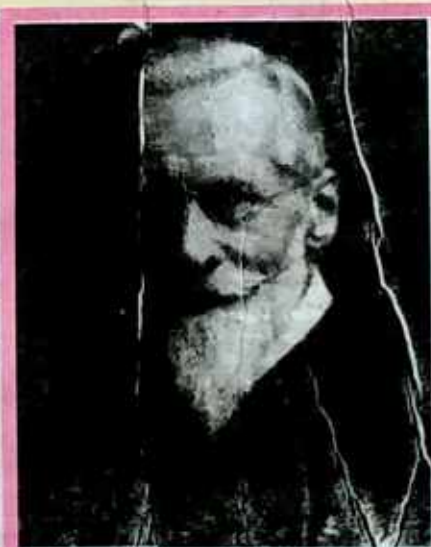
Atoms which are the building blocks of matter are very important to understand in Chemistry which is the science of matter. As already describe that the name atom is derived from the Greek word Atomos which means "indivisible", this initial concept of atom is believed to be given by the Greek philosopher Democritus (400BC) who suggested that matter is composed of tiny indivisible particles.

But with the advancement of science they are no more indivisible now, and can be easily broken into smaller particles. These particles are called subatomic particles. The important subatomic particles are the Electrons, protons and the neutrons. In the subsequent pages you will learn about the structure of atom, the arrangement of subatomic particles and discovery of these subatomic particles.

2.1 Theories and Experiments Related to Atomic Structure

Although, the Scientists from the ancient eras were trying to understand the atomic structure, but the important success in understanding this was made in and after the mid of eighteenth century, after the advent of cathode ray tube by William Crook.

A cathode ray tube is a glass tube linked with high voltage source of current at its two electrodes, the air pressure in the tube is reduced by using vacuum pump, William Crook observed that at ordinary pressure, when electricity is passed through this tube, nothing happens, but at reduced pressure 0.01 torr high voltage produces a beam of radiation which cause fluorescence on striking with some fluorescent material.



William Crook
(British)
Born 182—Died 1919

The structure of this cathode ray tube is shown in figure 2.1.

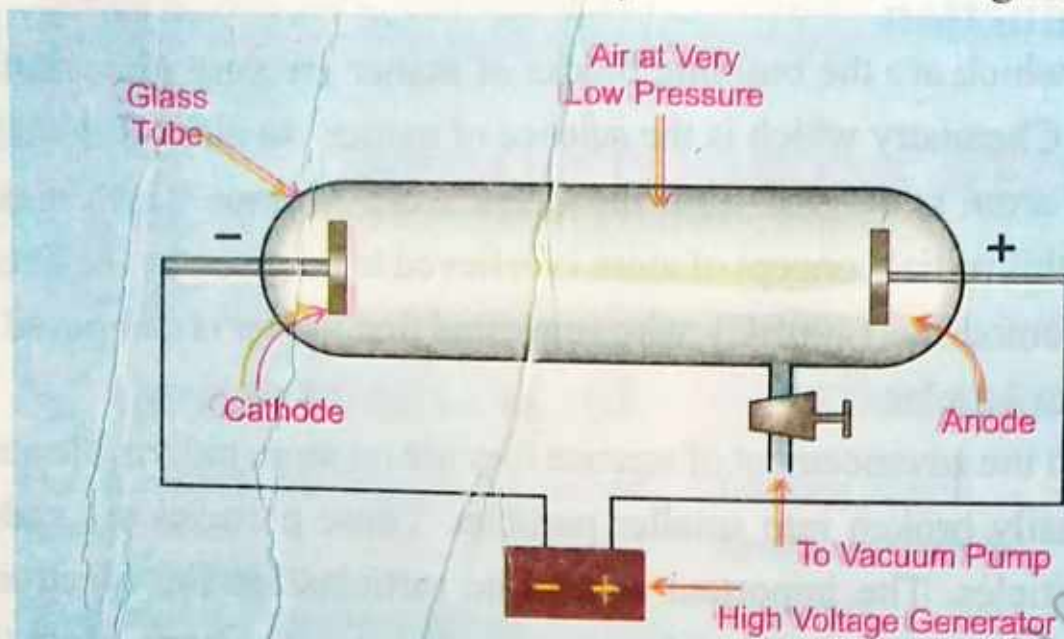


Figure 2.1: The Cathode Ray Tube

Discovery of Electron

It was observed that when high voltage current is passed through this tube at a very low pressure 0.01 torr then a beam of ray is produced. The direction of these rays is from cathode to anode, hence they originate from the cathode electrode, so are called cathode rays, fig 2.2.

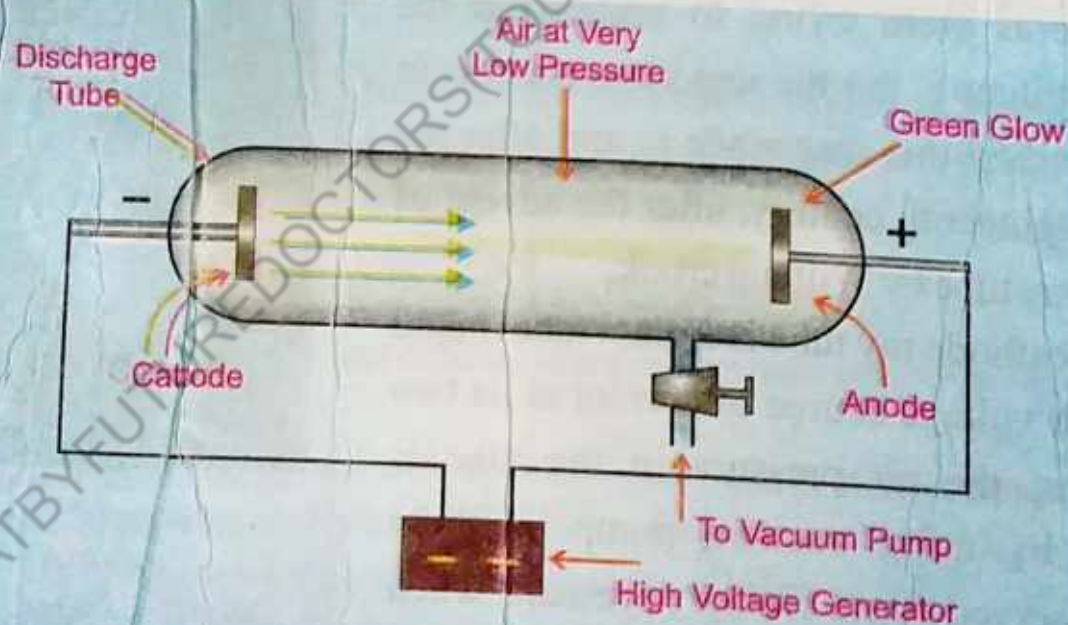


Fig 2.2: Cathode rays emission in Cathode ray tube

When an object is placed in the path of these rays, it is observed that they cast a sharp shadow; it means that they travel in straight line. (Figure 2.3)

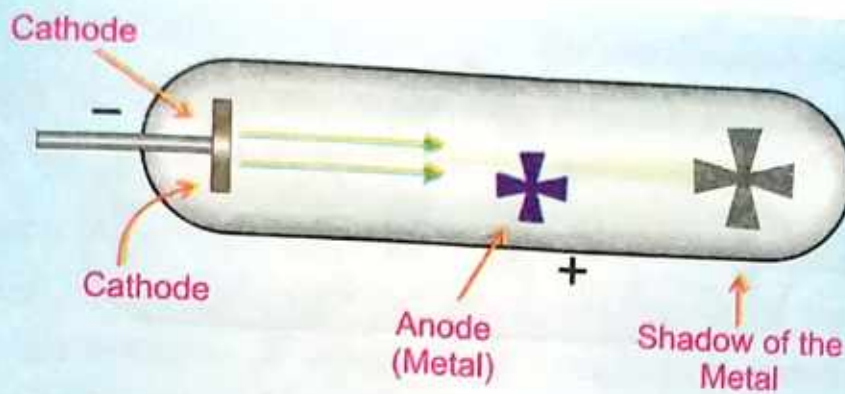


Fig 2.3: Experiment showing that cathode rays cast shadow

When a flywheel of mica is placed in their path it is moved showing that these rays are actually particle in nature, (Figure 2.4)

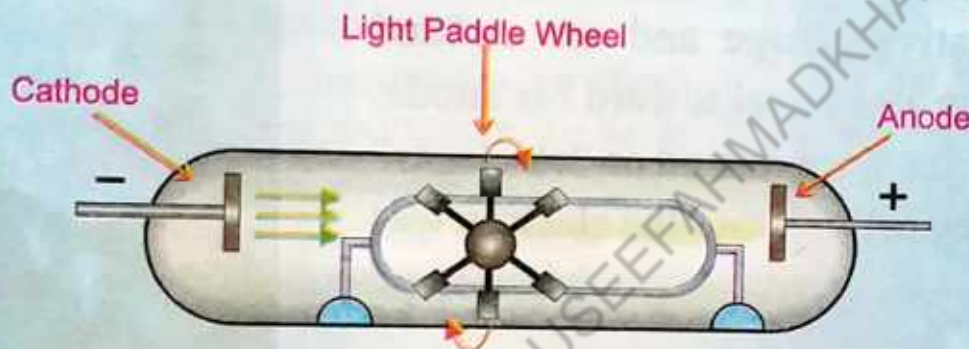


Fig 2.4: Experiment showing that cathode rays move fly wheel

When these rays are subjected to electric or magnetic fields, they were deflected towards the positive plate showing their negative nature. (Figure 2.5 and 2.6)

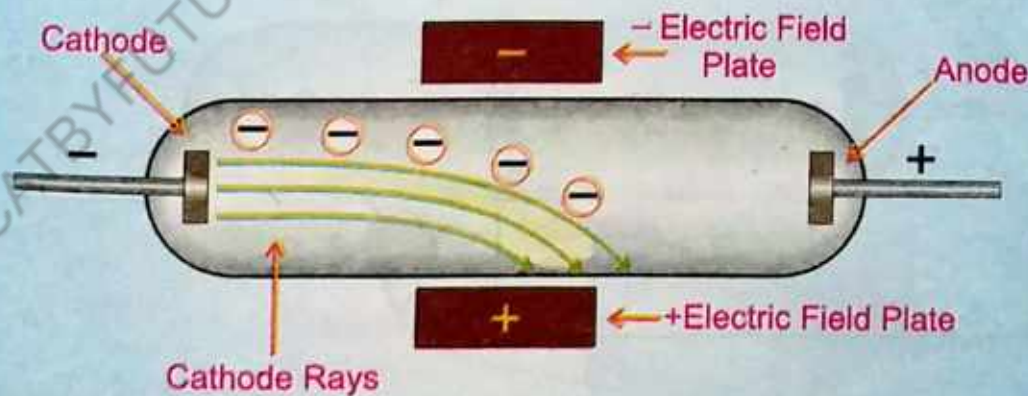


Fig 2.5: Experiment showing that cathode rays bent towards positive plate

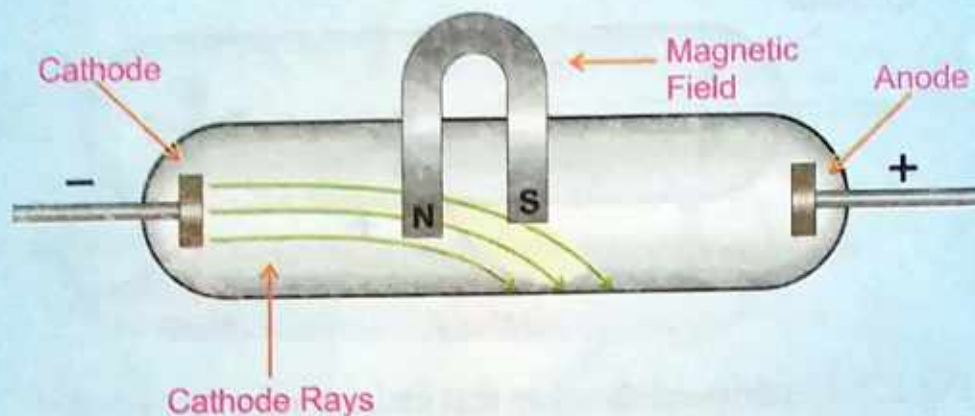


Fig 2.6: Experiment showing that cathode rays bent perpendicular in magnetic field

These experiments proved that cathode rays are actually tiny particles having negative charge and are emitted from cathode and travel toward his anode. These particles were named as “electrons” by G.J Stoney in 1891, thus the electron was discovered.

Discovery of Proton

In 1886 Gold Stein observed that glow appeared in the discharge tube at opposite end of anode at back of cathode if a perforated cathode is used (see figure 2.7).



Goldstein
(German Physicists)
1850—1930

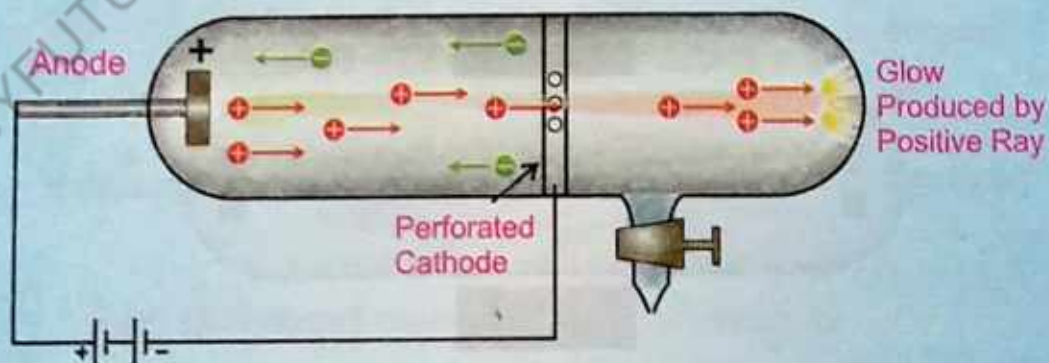


Fig 2.7: Canal rays produce by anode in cathode ray tube

The rays that form from the gas inside the discharge tube.

Further investigations showed that these rays have positive charge, have momentum. They were named Proton.

The Discovery of Neutron

In 1920, Ernest Rutherford postulated that there were neutral, massive particles in the nucleus of atoms. This conclusion arose from the disparity between an element's atomic number (protons = electrons) and its atomic mass (which is in excess of the mass of the known protons present). James Chadwick proved the existence of these particles in 1930 and he received the

Noble prize in 1935 for his success.

Thus all the basic fundamental sub atomic particles were discovered.

Soon after the discoveries of these particles Scientists tried to explain the basic structure of atom by their models.

The important atomic models are discussed here in the following sections.

2.1.1 Rutherford's Atomic Model

In 1910 (note that at this time only proton and electrons were known) Rutherford bombarded a thin gold foil (4×10^{-7} m) and observed that:



Rutherford
(New Zealand)
Noble Laureate
Chemistry,
Born 1871, Died 1937

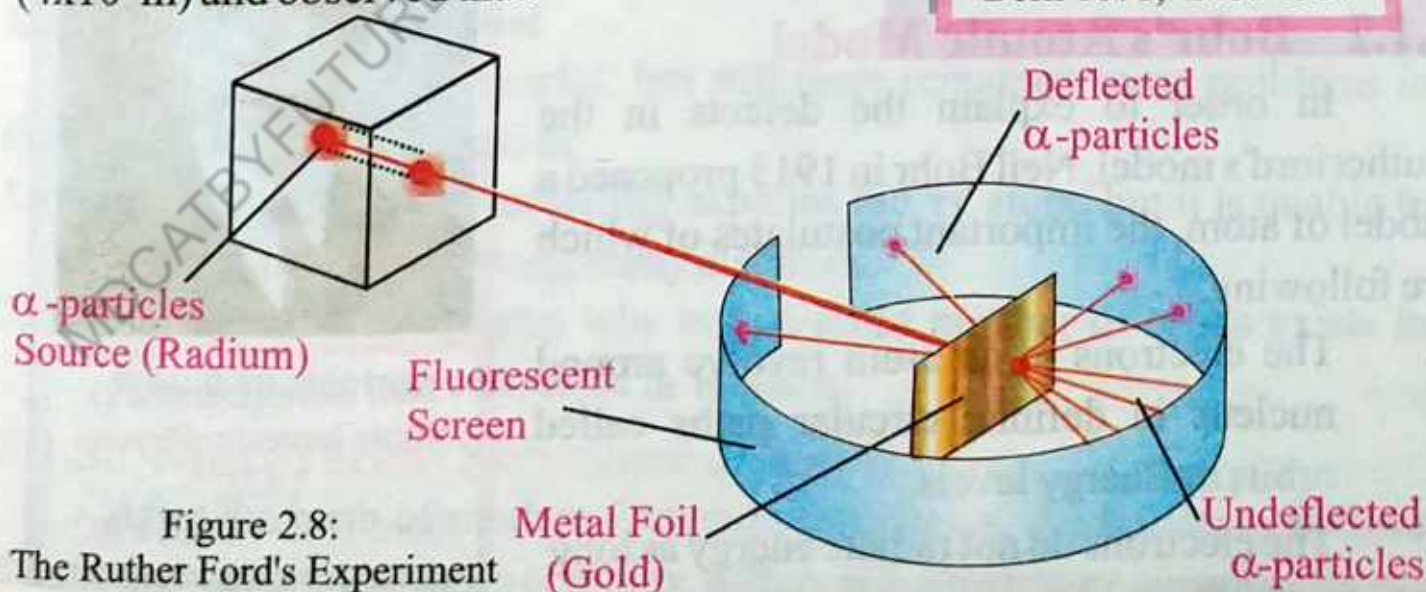


Figure 2.8:
The Rutherford's Experiment

Most of the alpha ray particles passed through foil. Some particles deflected (one in 20,000) at various angles. Very few deflected back. Rutherford concluded that, as most of the α -ray particles passed through foil without any resistance, so most of area in atom is empty. Some particles deflected so in the atom there is a very small area of positive charge with the mass of whole atom in their.

He named the positive centre of atom as Nucleus of atom.

On the basis of his experiment, Rutherford developed a model for atomic structure, **which has following important postulates.**

- An atom has a positive centre called nucleus.
- Around this nucleus the electrons revolve.
- The number of electrons are equal to positive particles so atom as a whole is neutral.

This model is just like the solar system where the centre is sun around which planets move. Rutherford's model gave the basic idea for structure of atom, although there were some problems. e.g why electrons do not fall in nucleus, as by classical mechanics a charged particle moving should radiate energy, and when all energy is lost by electron, it should fall in nucleus. Moreover due to radiating electron, a continuous spectra should be observed, but this is not the case.

Recent studies show that nucleus is about 10^{-15} m in diameter and atom has about 10^{-10} m diameter, (See figure 1.5).

2.1.2 Bohr's Atomic Model

In order to explain the defects in the Rutherford's model, Neil Bohr in 1913 proposed a model of atom, the important postulates of which are following.

1. The electrons in an atom revolve around nucleus in definite circular paths called orbits or Energy levels.
2. The electrons do not radiate energy as long



Neil Bohr (Denmark)
Noble laureate Physics
1922,
Born 1885—Died 1962

as they remain in an orbit.

3. For electrons only that orbit is allowed for which the angular momentum, (mvr) of electron is equal to integral multiple of $\frac{h}{2\pi}$.
4. The electron jumps to higher energy level or emits energy coming to lower energy level by absorbing or emitting energy respectively.
5. This energy is quantized and is given by: $E = h\nu$

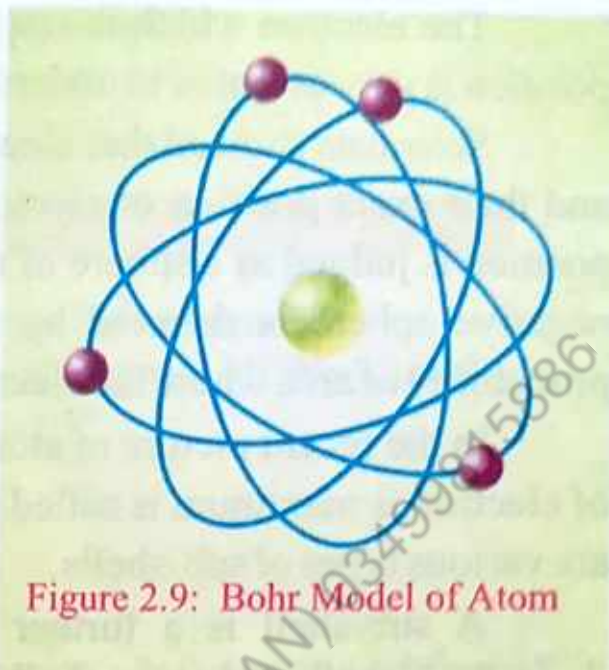


Figure 2.9: Bohr Model of Atom

Where h = Planck's constant and
where ν (spelled as nu) is the frequency of radiation.

The Bohr model has following merits over classical models:

1. Bohr model was a successful approach in explaining many basic characteristics of atomic structure.
2. His theory was conflicting with classical mechanics, and thus he started modern mechanics.
3. He told that the principles at macroscopic level do not apply on microscopic level.
4. Bohr explained the spectral lines of 'H' atom successfully.
5. He calculated Radius and energy of electron in an orbit successfully

Limitations Of Bohr Model

Although very successful, but still there remained some problems in the Bohr atomic model, such as

1. The Bohr model successfully explains the 'H' atom, but it is unable to explain the atoms having many electrons.
2. It also do not explains why only a fixed number of orbits exists in which an electron can reside in 'H' atom.

2.1.3 More recent Structure of Atom

After the birth of modern Quantum mechanics, the scientists believe the structure of atom as more complex, different from earlier concepts.

The electron which is responsible for the properties of atoms and its position is very essential to understand.

Scientists showed that electron behaves as a particle as well as a wave, and their exact position of electron is very hard to determine. Although its position is judged as a sphere of negative charge cloud around nucleus. This negative sphere is denoted by the dots and density of dots show more probability of area where the electrons are located.

In the recent picture of atom the areas where the probability of finding of electron is maximum is called the Sub-shell or orbital and in an atom there are various types of sub-shells.

A sub-shell is a further division of shell, and as the shell number increases, the size of shell also increases and more orbitals are included in this, this is explained in the following table 2.1.

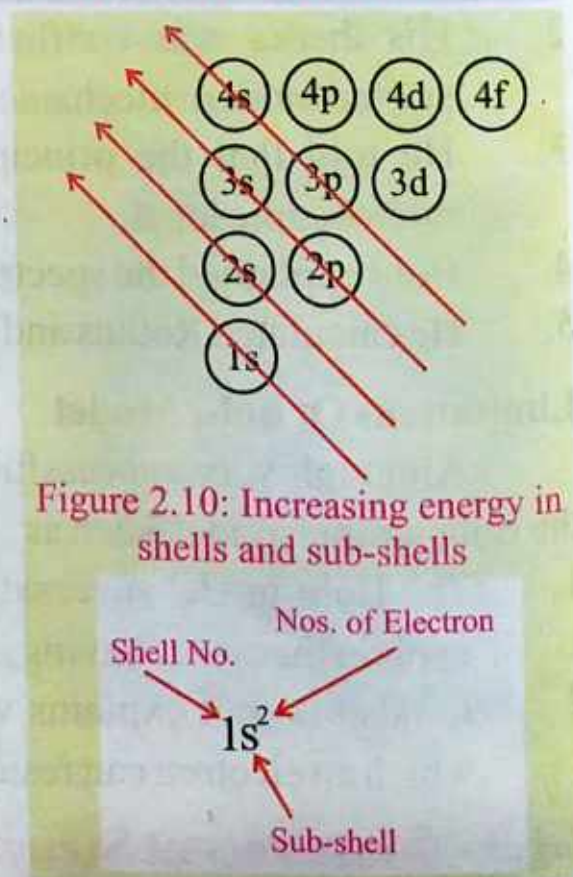
No of shell	Shell name	Sub-shells
1	K	s
2	L	s,p
3	M	s,p,d
4	N	s,p,d,f

Table 2.1: Table showing number of shell, name of shell and the sub shells that it contains.

The orbitals grow towards more complexity as their number increases, e.g the s orbital is spherical in shape and is directionless, while p orbital is dumbbell in shape and are located in 3 directions, i.e. along x-axis, y-axis, and z-axis. (d) sub-shell have 5 orbitals and (f) sub-shell have 7 orbitals.

2.2 The Electronic Configuration

The position of electrons in an atom is called its electronic configuration. In order to calculate electronic configuration it is kept in mind that electrons will fill that orbitals first whose energy is least, The filling of orbitals is easily calculated by following figure 2.10, the arrow sign shows the filling of electrons in specific sub shells.



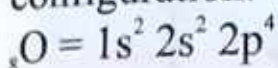
This is further explained by the following examples:

Example:

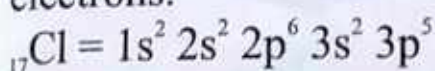
Write down the electronic configuration of 'O' and 'Cl'.

Solution

O having 8 electrons has electronic configuration.



Cl has 17 atomic number so it has 17 electrons.



Activity: 2.1

Draw the electronic configuration of "N=7" and "P=15".

Following table 2.2 that shows the electronic configuration of first 18 elements.

Atomic No.	Symbol	Electronic Configuration	Atomic No.	Symbol	Electronic Configuration
1	H	$1s^1$	10	Ne	$1s^2 2s^2 2p^6$
2	He	$1s^2$	11	Na	$1s^2 2s^2 2p^6 3s^1$
3	Li	$1s^2 2s^1$	12	Mg	$1s^2 2s^2 2p^6 3s^2$
4	Be	$1s^2 2s^2$	13	Al	$1s^2 2s^2 2p^6 3s^2 3p^1$
5	B	$1s^2 2s^2 2p^1$	14	Si	$1s^2 2s^2 2p^6 3s^2 3p^2$
6	C	$1s^2 2s^2 2p^2$	15	P	$1s^2 2s^2 2p^6 3s^2 3p^3$
7	N	$1s^2 2s^2 2p^3$	16	S	$1s^2 2s^2 2p^6 3s^2 3p^4$
8	O	$1s^2 2s^2 2p^4$	17	Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$
9	F	$1s^2 2s^2 2p^5$	18	Ar	$1s^2 2s^2 2p^6 3s^2 3p^6$

Table 2.2: Electronic configuration of first 18 elements

2.3 Isotopes

The term isotope has been originated from two words, 'iso' that means same, and 'topes' (from topos) which means place.

The isotopes are the atoms forms of same element that have same

atomic number but different atomic masses. e.g hydrogen has three isotopic forms i.e Protium (${}_1\text{H}^1$) deuterium (${}_1\text{D}^2 / {}_1\text{H}^2$) and Tritium (${}_1\text{T}^3 / {}_1\text{H}^3$).

These isotopes of hydrogen are drawn in the figure 2.11.

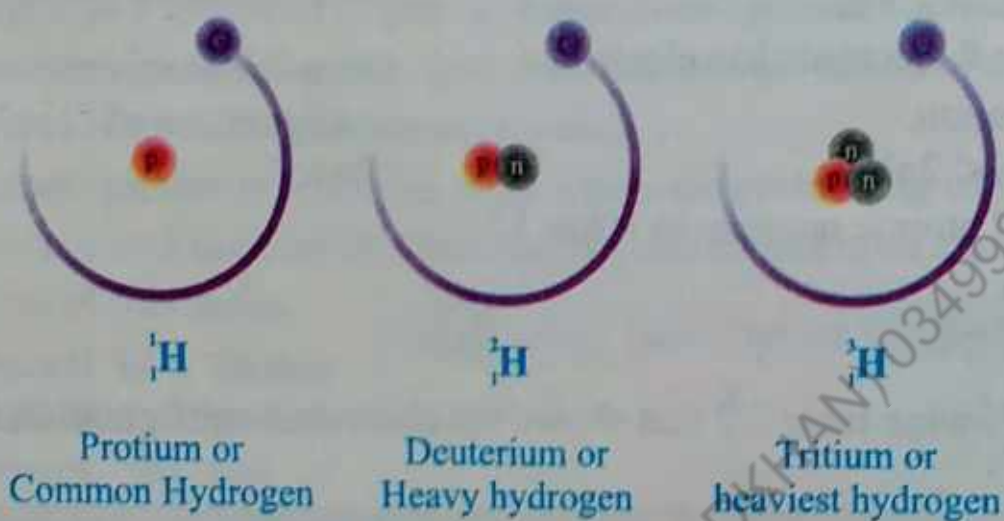


Figure 2.11: Isotopes of Hydrogen

Similarly the Carbon also has 3 isotopic forms, which are drawn in figure 2.12.

Carbon-12

Carbon-13

Carbon-14

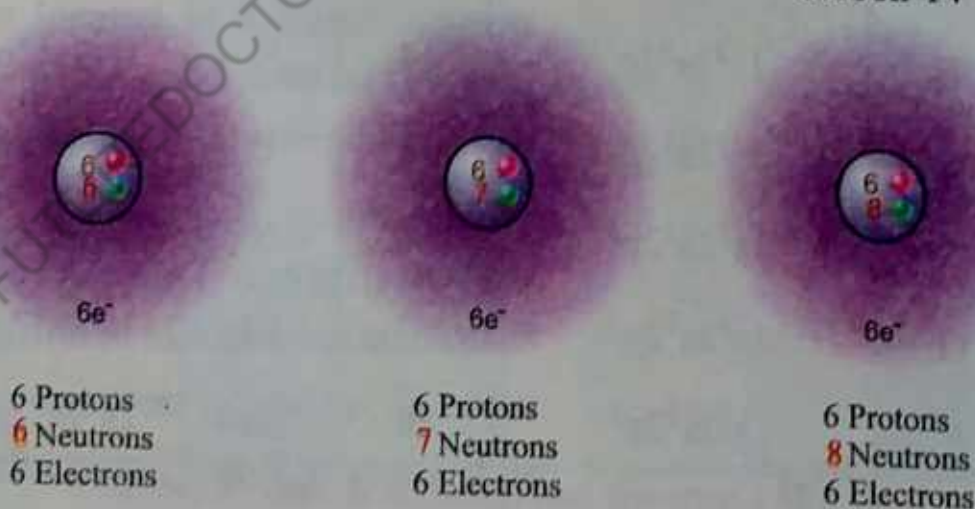


Figure 2.12: Isotopes of Carbon

Almost all of the elements form isotopes, in the following table 2.3 shows isotopes of some elements and their percentage in nature.

TABLE OF ISOTOPES

Element	Protons	Mass Number	Symbol	Abundance %
H	1	1	${}_1\text{H}^1$	99.985
	1	2	${}_1\text{D}^2$	0.015
	1	3	${}_1\text{T}^3$	Trace
C	6	12	${}_6\text{C}^{12}$	98.90
	6	13	${}_6\text{C}^{13}$	1.10
	6	14	${}_6\text{C}^{14}$	Trace
Cl	17	35	${}_{17}\text{Cl}^{35}$	75.77
	17	36	${}_{17}\text{Cl}^{36}$	Trace
	17	37	${}_{17}\text{Cl}^{37}$	24.23
U	92	233	${}_{92}\text{U}^{233}$	Trace
	92	234	${}_{92}\text{U}^{234}$	0.0055
	92	235	${}_{92}\text{U}^{235}$	0.7200
	92	238	${}_{92}\text{U}^{238}$	99.2745

Table: 2.3: Table of some isotopes

Activity: 2.2

Using the data given in the table 2.3 draw the structures of Chlorine isotopes.

2.3.1: Uses of Isotopes

The isotopes are used in many fields of life, e.g.

- 1- They are used in isotopic labeling, normally, atoms of a given element are indistinguishable from each other. However, by using isotopes of different masses, they can be distinguished by mass spectrometry or infrared spectroscopy.

- 2- As tracers or markers in chemical reactions, this technique is called radio isotopic labeling. If radioactive isotopes are used in chemical reactions, they can be detected by the radiation they emit, and thus mechanisms of different reactions can be understood quite easily.
- 3- The isotopes are also used to understand the mechanisms of various natural processes like photosynthesis.
- 4- A technique similar to radio isotopic labeling is radiometric dating; using the known half-life of an unstable element we can calculate the amount of time that has elapsed since a known level of isotope existed. The most widely known example is radiocarbon dating used to determine the age of carbonaceous materials.
- 5- Several forms of spectroscopy rely on the unique nuclear properties of specific isotopes. For example, nuclear magnetic resonance (NMR) spectroscopy can be used only for isotopes with a non zero nuclear spin.
- 6- The isotopes also used for the treatment of different types of cancers in the body, the cancerous cells are killed by using radio active rays of radio isotopes. This process is called radiotherapy.
7. Isotopes are also used for the investigation of different diseases, e.g. in order to determine the efficiency of kidneys heavy water (D_2O) is past in the body and its traveling is monitored. The malfunctioning kidney does not passes water, so the relative functioning of each kidney is determined.
8. The isotopes are also used to generate electricity. This is because of the tendency of radio isotopes to emit radioactive rays, these rays are also a good source of energy and this energy is carefully used to run turbines. In Pakistan Kanupp (Karachi), Nilore (Islamabad) and Chashma Nuclear Energy centers produce electricity based on this method.

SUMMARY OF THE CHAPTER

- ❖ **Atoms** are the building blocks of matter; they are the smallest particle of matter having all the characteristics of their element.
- ❖ The electron and proton, two important sub atomic particles of atoms were discovered by the series of experiments with **cathode ray tubes**.
- ❖ Rutherford discovered the **nucleus** of atom, and proposed the basic structure of atom.
- ❖ Bohr told that electron revolve in fixed paths around nucleus called **orbits**, during revolving around the nucleus the electrons do not radiate energy in normal circumstances. Bohr told that in special cases, electron absorb energy and get excited, jump to higher energy level (or orbit.), but as the normal conditions reach, these excited electron release their extra energy and come back in previous energy level.
- ❖ The Bohr energy levels or orbits were later named as **shells** and they are further divided into **sub shells** (orbitals). In an atom different types of sub shells are found, they are named, s, p, d, f.
- ❖ The electrons in an atom are filled in the increasing order of energy, i.e. the orbital with least energy is filled first.
- ❖ Filling of electrons in an atom is called its **electronic configuration**. The electronic configuration plays an important role in predicting the properties of atoms.

EXERCISE

Q1. Fill in the blanks with suitable words.

- i) Experiments with tube helped a lot in understanding the nature of sub atomic particles.
- ii) According to Bohr the angular momentum is of $h/2\pi$.
- iii) The most stable isotope of C has mass
- iv) Rutherford bombarded the gold foil with
- v) Cathode rays were given the name of electron by in

1891.

- vi) Nucleus was discovered by
- vii) The presence of neutron was confirmed by in 1930.
- viii) Electrons radiate energy when jumps from to energy levels.
- ix) The modern research showed that electron behaves both as a --- and a particle.
- x) Number of neutrons in the atom of $_{17}\text{Cl}^{35}$ are

Q2. Choose the correct answer.

- i) Structure determination helps to determine the:
 - (a) Physical characters
 - (b) Chemical characters
 - (c) Both physical and Chemical characters
 - (d) Does not helps in any way
- ii) Recent Studies shows that atom has an average size of:
 - (a) 10^{-10} meters
 - (b) 10^{-8} meters
 - (c) 10^{-15} meters
 - (d) 10^{-12} meters
- iii) 1 angstrom is equal to the:
 - (a) 10^{-10} meters
 - (b) 10^{-12} meters
 - (c) 10^{-15} meters
 - (d) 10^{10} meters
- iv) Neil Bohr mainly concentrated upon:
 - (a) Nucleus of atom
 - (b) Behavior of electrons
 - (c) protons
 - (d) Neutrons
- v) Electronic configuration tells about the:
 - (a) Number of sub-shells
 - (b) Number of shells
 - (c) Valency
 - (d) All sub-shell, shell and valency

Q3. Answer the following questions in short.

- i) Describe the structure of cathode ray tube.
- ii) Do you agree that atoms are indivisible?
- iii) Write a note on discovery of neutrons.
- iv) Why position of electrons is hard to know?

- v) Discuss the shape of 's' and 'p' orbitals.
- vi) What is the role of electronic configuration in determining the chemical nature of an element.
- vii) Write electronic configuration of Na and Na^+ .
- viii) What do you know about shells of atoms?
- ix) Describe division of shells into sub shells?
- x) Why Rutherford used α -rays and not γ -rays.

Q4. Answer the following question with reasoning.

- i) Can you explain the experiment that lead to the development that e^- has negative charge and proton has the +ve charge?
- ii) Why Rutherford used gold foil instead of any other metal foil?
- iii) Why scientist concluded that electrons revolve around nucleus and not neutrons?
- iv) Why Tritium, is not found in larger quantities in nature.
- v) Can you explain why in the cathode ray tube pressure is reduced?

Q5. Discuss various experiments that lead to establish the Atomic structure?

Q6. Can you describe the experiment and result deduced by Rutherford for explaining atomic structure?

Q7. Explain how Bohr helped in understanding the structure of atoms?

Q8. What do you understand by the term "isotopes"? Explain with examples.

Q9. Discuss use of isotopes in medical science?

PERIODIC TABLE & PERIODICITY OF PROPERTIES

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the,

- Symbols of Elements
- Characteristics of sub atomic particles

TIME ALLOCATION

Teaching periods = 12
Assessment periods = 03
Weightage = 11%



LEARNING OUTCOMES

After reading this chapter, the Student will be able to:

- Distinguish between a period and a group in the periodic table. (Understanding)
- State the Periodic law. (Remembering)
- Classify the Elements. (Analyzing)
- Determine the demarcation of the Periodic table into s-block and p-block. (Remembering)
- Explain the shape of the Periodic table. (Analyzing)
- Determine the location of the families in the Periodic table. (Understanding)
- Recognize the similarity in the chemical and physical properties of the Elements in the same family of elements. (Understanding)
- Identify the relationship between electronic configuration and the position of an element on the periodic table. (Analyzing)
- Explain how shielding effect influences periodic trends. (Applying)
- Describe how electronegativities change within a group and within a period in the periodic table. (Analyzing)

3.0 Introduction to classification of Elements

As already discussed that Chemistry deals with the study of matter or the elements, so far 118 elements have been discovered, if they are studied separately it would be a difficult job for any one so there is a need of classifying such a huge number of elements. As the number of discovered elements increased, the need to classify them also increased, because it became more difficult to remember the characteristics of the elements individually.

The early classifications were full of errors, but with the passage of time, the more précised classifications were presented by various scientists. Dobriener a German scientist in 1820 classified the elements into triads (a classification where each group has 3 members). Afterward an English Scientist Newland in 1865 presented the concept of classifying the elements based on law of octaves. A very important classification was presented by a Russian Chemist Mendeleev in 1869 which was based upon the atomic masses of the elements.

And after the discovery of protons, the classification based upon the number of protons built up, which finally resulted in a comprehensive periodic table called the Modern periodic table.

This classification made very easy the understanding and remembering the properties of elements, and also various predictions regarding the physical and chemical properties of these elements are now possible with very accurate results.

3.1 The Periodic Table

With the discovery of protons a new era of classification of elements started, and scientists observed that if elements are arranged in increasing



Dmitri Ivanovich
Mendeleev (Russian)
(1834-1907)

order of number of their protons (or simply their atomic numbers) all the discrepancies in classification can be removed.

Hence the modern periodic law originated stating that the properties of elements are periodic functions of their atomic numbers.

Thus all the known elements are put in the table in the increasing order of atomic number in the manner that elements with similar properties come in the same column.

In the modern periodic table the columns are called groups, and the rows are called periods. There are 18 groups and 7 periods in this periodic table (previous classification had 8 groups). Thus in the new era the periodic table has the set up shown in figure 3.1.

3.1.1 Periods in Periodic Table

A period is a horizontal row in the periodic table of the elements. In the periodic table, there are 7 periods.

In the periodic table the number of period shows actually the numbers of shell in that particular row of elements, it means the elements in the period-1 have just one shell, and the elements in 2nd period have two shells, and so on. The first period which contains only two elements (H and He) is the smallest one, 2nd and 3rd periods have 8 elements each and 4th and 5th periods have 18 elements, the 6th period contains 32 elements and is the longest one, 7th period is incomplete one and new elements discovered are being added in this period. From the 6th and 7th periods two rows are withdrawn and are written below the others, because their properties are remarkably similar with each other. These series are termed as "Lanthanides" and "Actinides" series.

3.1.2 Groups in Periodic Table

As mentioned earlier there are eighteen (18) groups in the modern periodic table (in previous classification there were 8 main groups), a group is also known as family.

The groups are named by the numbers beginning from 1 to 18. These groups are shown with all member in the figure 3.1. Some of the group are

1		18																VIIA	
IA																			
1																		2	
H Hydrogen																		He Helium	
2																		3	
Li Lithium																		Be Beryllium	
3																		4	
Na Sodium																		Mg Magnesium	
4																		5	
K Potassium																		Ca Calcium	
5																		6	
Rb Rubidium																		Sr Strontium	
6																		7	

also known by their specific names e.g., group '1' is called alkali metals, group '2' is alkaline earth metals, group '16' is called chalcogens, group '17' is called halogens and group '18' is called noble gases.

Some groups are also represented by the head element of their group e.g., group '13' as boron family, group '14' as carbon family, group '15' nitrogen family and group '16' as oxygen family.

In a group the members show many same properties like same last electronic configuration, same valencies and same type of chemical characters.

Group	Family Name
1	Alkali Metals
2	Alkaline Earth Metals
13	Boron Family
14	Carbon Family
15	Nitrogen Family or chalcogens
16	Oxygen Family
17	Halogens
18	Noble Gases

Table: 3.1: Some groups with their family name

Interesting Information

In the older fashion the names of the groups are written in Roman letters, and in new fashion the groups are written English numbers.

Activity: 3.1

Point out the similarities in the members of Group-17 (The halogens) of the periodic table.

3.1.3 Blocks in Periodic Table

The name of blocks depends upon the outermost electronic arrangement of the element and is very important in determining the chemical properties of the element.

The periodic table has been divided four blocks, i.e. s-block, p-block, d-block and f-block, which are shown separately in the following figure 3.2.

1 IA 1 H	2 IIA 4 Be	s - block p - block										d - block f - block										13 IIIA 5 B	14 IVA 6 C	15 VA 7 N	16 VIA 8 O	17 VIIA 9 F	18 VIIIA 2 He										
3 Li	4 Be	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII			9 VIII	10 VIII	11 IB	12 IIB	13 IIIA 13 Al	14 IVA 14 Si	15 VA 15 P	16 VIA 16 S	17 VIIA 17 Cl	18 VIIIA 18 Ar																		
11 Na	12 Mg	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	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		
19 K	20 Ca	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	55 Cs	56 Ba	57 La	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			
37 Rb	38 Sr	71 La	72 Ce	73 Pr	74 Nd	75 Pm	76 Sm	77 Eu	78 Gd	79 Tb	80 Dy	81 Ho	82 Er	83 Tm	84 Yb	85 Lu	86 Hf	87 Ta	88 W	89 Re	90 Os	91 Ir	92 Pt	93 Au	94 Hg	95 Tl	96 Pb	97 Bi	98 Po	99 At	100 Rn	101 Fr	102 Ra	103 Ac			
		89 Ac	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	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og	119 Uut	120 Uuq	121 Uup	122 Uuh	123 Uus	124 Uuo

Figure 3.2 Blocks in the modern Periodic table

3.2 Periodicity of Properties

A great advantage of the periodic table is the **periodicity**, i.e. the elements show a gradual increase or decrease in physical and chemical properties, e.g. in atomic radius, ionization energy, electron affinity, and electronegativity, oxides etc.

We will discuss now some of the important of these properties individually in a little detail:

3.2.1 Atomic Size

Atomic size or more properly atomic radius is half of the distance between the nuclei of a diatomic molecule, as shown in figure 3.3.

Although in order to determine the size of an atom the distance between the last electron and nucleus should be in focus but due to fast movement of electrons the exact position of electron is difficult to identify and so for clear reference the distance between to nuclei of a homo nuclear is taken as standard and half of this distance gives the exact of atomic size e.g. for size of 'H' atom 'H₂' is taken.

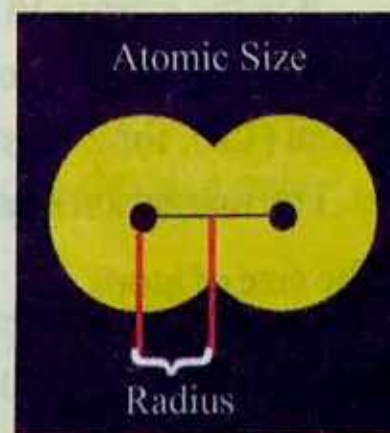
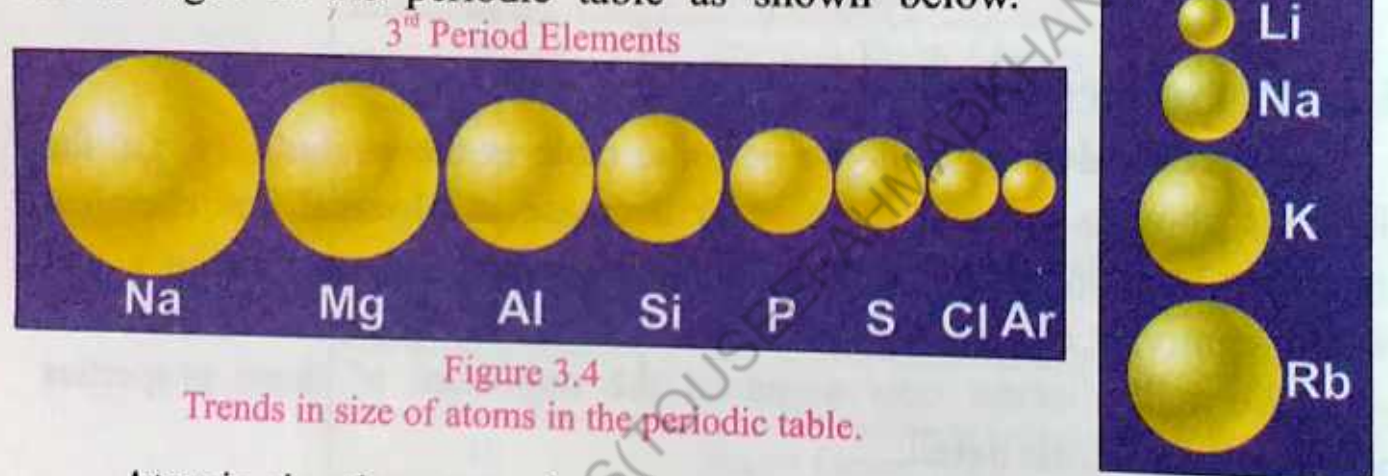


Figure: 3.3:
The Atomic size

The Atomic radius is influenced by 3 main factors:

- Energy level → Higher the energy level greater will be the size.
- Nuclear charge → more is the nuclear charge in a periods lesser will be the size.
- Shielding effect → greater is the shielding effect of electrons i.e. the repulsion between the electrons, higher will be the size.

The Atomic radius increases down a group due to increase in energy level, and decrease along a period from left to right in the periodic table as shown below.



Atomic size decreases along a period, which is due to increase in the nuclear charge, while in a period all atoms have same energy level.

3.2.2 Ionization Potential / Energy

It is the energy to remove the least bonded last electron from an atom. The energy required to remove the last electron is called 1st ionization potential (I.P), for 2nd last electron it is called 2nd ionization potential (I.P) and so on. The ionization energy depends upon:

- The size of atom ⇒ Greater the size less will be the ionization potential.
- The nuclear charge ⇒ higher is the nuclear greater; more will be the ionization potential.
- The Shielding effect ⇒ the electrons in the outermost energy level experience more inter electron

repulsion (shielding), so greater shielding less will be the ionization potential.

- The filled and half filled $\Rightarrow$ filled and half filled orbitals have higher ionization energy, so achieving them is easier.

In the periodic table, the ionization potential decreases down the group because the electron is farther from nucleus and also there is more shielding. Along a period the ionization energy increases because, although all the atoms in a period have same energy level and same shielding but the nuclear charge increases so I.P increases. (Exceptions for filled and half filled sub shells occur). Full energy levels require lot of energy to remove their electrons, so noble gases have very high ionization potentials. The arrows show (Figure 3.5) the total trend changes among elements in the periodic table.

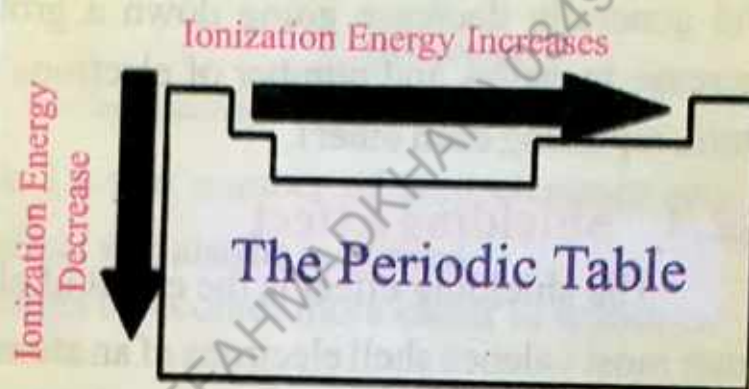


Figure 3.5: Trends in ionization energy of atoms in periodic table.

3.2.3 Electron Affinity (E_{ea})

The electron affinity is the energy change when an electron is added to a gaseous atom. Generally, the element liberates energy when electron is added to them:



The value of energy released greatly depends upon the size and nuclear charge of that element.

Thus the atoms whose anions are relatively more stable than neutral atoms have a smaller E_{ea} . e.g. Chlorine most strongly attracts extra electrons, and its anion is more stable as compared to mercury, which most weakly attracts an extra electron.

E_{ea} of noble gases are close to 0.

E_{ea} generally increases across a period (row) since the radius slightly decreases, because of the increased attraction from the nucleus, and the number of electrons in the top shell increases (helping the atom reach maximum stability) in the periodic table and generally decrease going down a group (family) (because of a large increase in radius and number of electrons that decrease the stability of the atom, repulsing each other).

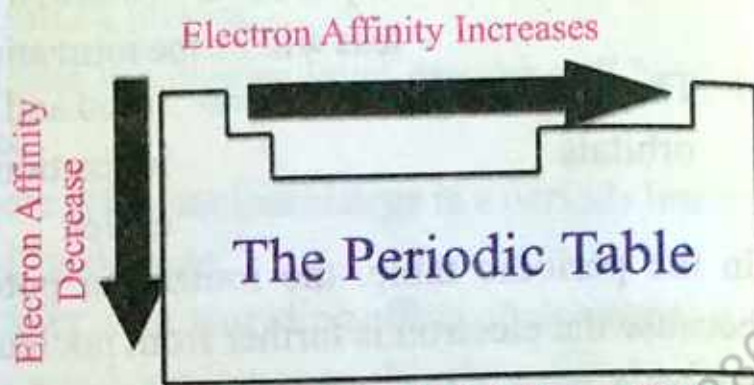


Figure 3.6: Trend of electron affinity in periodic table

3.2.4 Shielding Effect

The shielding effect is the effect of electrons between the nucleus and outer most valence shell electrons of an atom.

This shielding effect causes the inner electrons to block the outer electrons from the nuclear charge of the nucleus so it becomes easier for the valence electrons to be removed. The shielding effect increases down the group in the periodic table, but across a period it remains unchanged. This is further elaborated in the following figure 3.7 and figure 3.8.

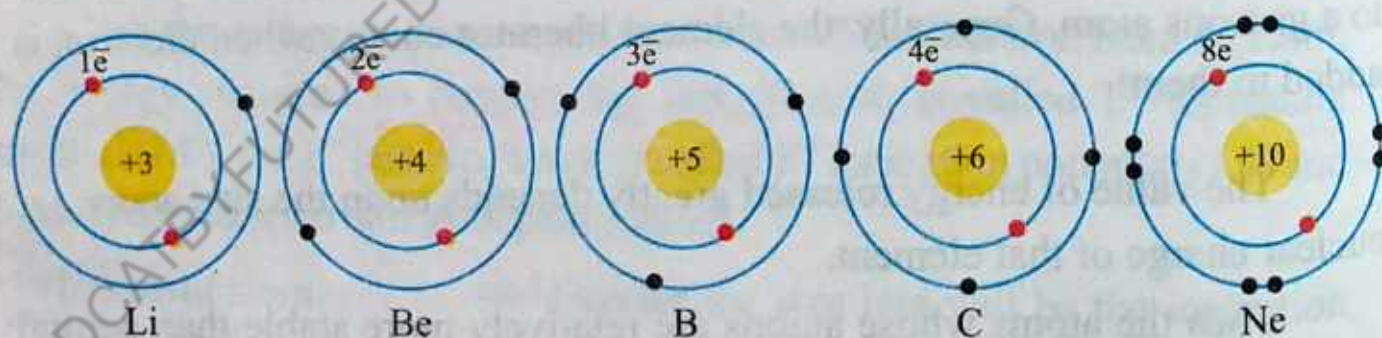


Figure 3.7: Shielding across period 2

• = Core Electron • = Valence Shell Electron

It can be seen that in the above figure all atoms have same number of core electron (2), so their effect will also be same.

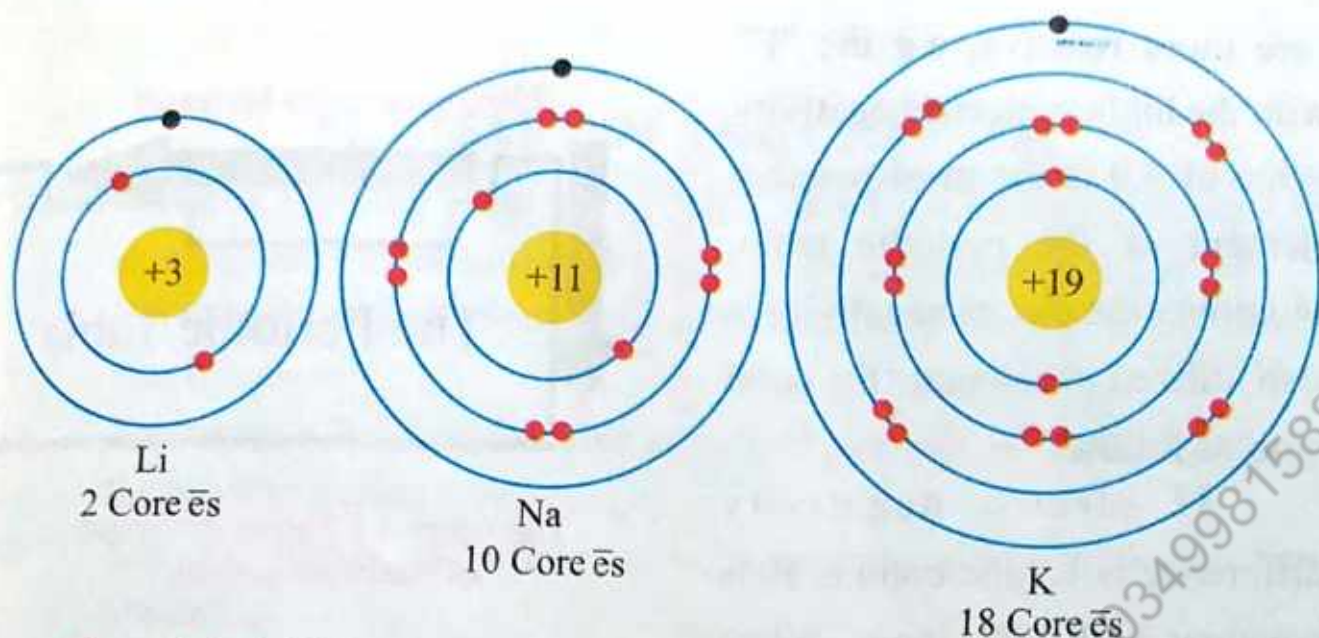


Figure 3.8: Shielding in group 1 ● = Core Electron • = Valence Shell Electron

It can be seen here in figure 3.8 that number of core electrons are increasing down the group, so their effect will also be increased.

Due to increase in shielding effect it becomes more easier to withdraw the valence shell electrons.

3.2.5 Electronegativity

The term electronegativity is defined as the ability of an atom to attract shared electron pair of bond in a compound. The electronegativity of an element is the net effect of electron affinity and first ionization potential.

Electronegativity is the property of non metals because non metals have lower size and more nuclear charge, so they attract the shared electron pair, this ability of attracting electron pair decreases as the size increases down the groups.

In the periodic table, electronegativity values decreases down the group and increases along the period.

Thus 'F' has the greatest electro negativity value of 4.0 and 'Cs' has lower value of '0.7'.

The electronegative values are very often used in prediction of reactivity and the bond type. The elements with high electronegativity value

called the **electron affinity**.

- ❖ The electron affinity increases along a period from left to right and decreases down a group.
- ❖ The inner electron produce a shield around nucleus, and make it easy for the valence shell electron to be removed, this effect is called the **Shielding effect**.
- ❖ The shielding effect increases down a group and decreases along a period.
- ❖ The tendency of an atom to attract shared electron pair is termed as the **Electronegativity** of that atom.
- ❖ The electronegativity increases along a period and decreases down a group.

EXERCISE

Q1. Fill in the blanks with suitable words.

- i) Triads were proposed by
- ii) Mendeleev in classified the all known Elements into groups, and is regarded as of modern periodic table.
- iii) Modern Periodic table has groups and periods.
- iv) A period is a of elements in the periodic table.
- v) Halogen constitutes the group of periodic table.
- vi) The concept of electronegativity was given by
- vii) Smaller atoms have ionization energy.
- viii) d-block elements are also known as
- ix) The size of atoms in the periods from left to right.
- x) Noble gases constitute the group of periodic table.

Q2. Choose the correct answer.

- i) In a group the elements have almost same
 - (a) Physical characters
 - (b) Chemical characters
 - (c) Both physical and chemical characters

- (d) Neither physical nor chemical
- ii) Group 1 members are generally known as:
(a) Alkali metals (b) Alkaline earth metals
(c) transition metals (d) noble gases
- iii) At ordinary temperature and pressure, Hydrogen exists as:
(a) liquid (b) gas (c) solid (d) plasma
- iv) Periodicity is seen in the periodic table in:
(a) groups (b) periods
(c) both in groups and periods
(d) Neither in group nor in period
- v) The atomic size increases in:
(a) group (b) period (c) in both of above (d) none of above

Q3. Answer the following questions in short.

- i) How Dobriener classified elements? Why his classification did not received popularity?
- ii) What is the resemblance between Newland's classification elements and Music?
- iii) Why in the history of periodic table the name of Mandeleev is so important?
- iv) Describe the importance of groups in periodic table.
- v) How many periods are there in Modern periodic table?
- vi) What are Alkali metals?
- vii) Discuss importance of Noble gases.
- viii) What do you understand by the term "blocks of periodic table"? What types of blocks are there in modern periodic table?
- ix) What is ionization energy? How it varies across modern periodic table?
- x) Define periodicity of elements in modern periodic table.

Q4. Answer the following questions with reasoning.

- i) What is the relation between electronic configuration and the positioning of element in period table?

- ii) How atomic size is related with the nuclear charge?
- iii) How electron affinity varies in periodic table?
- iv) How electronegativity is important for chemists?
- v) Discuss inertness of noble gases.

- Q5. Why it is necessary to classify the elements? Explain with reasoning.
- Q6. Discuss the importance of groups and periods in the modern periodic table?
- Q7. What do you understand by periodicity of properties? Explain with examples.
- Q8. Write a brief note on evolution of periodic table?
- Q9. What is modern periodic law? Explain with examples.
- Q10. What do you know about the blocks of the modern periodic table?

STRUCTURE OF MOLECULES

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the,

- Electronegativity concept.
- Polarity of compounds

TIME ALLOCATION

Teaching periods	= 16
Assessment periods	= 04
Weightage	= 15%

LEARNING OUTCOMES

After reading this chapter, the Student will be able to,

- Find the number of valence electrons in an atom using the Periodic table. (Applying)
- Describe the importance of noble gas electronic configuration. (Understanding)
- State the octet and duplet rules. (Remembering)
- Explain how elements attain stability. (Understanding)
- Describe the ways in which bonds may be formed. (Remembering)
- State the importance of noble gas electronic configurations in the formation of ion. (Applying)
- Describe the formation of cations from an atom of a metallic element. (Applying)
- Describe the formation of anions from an atom of a non-metallic element. (Applying)
- Describe the characteristics of an ionic bond. (Understanding)
- Recognize a compound as having ionic bond. (Analyzing)
- Identify characteristics of ionic compounds. (Understanding)
- Describe the formation of a covalent bond between two non metallic elements. (Understanding)
- Describe the characteristics of ionic and covalent compounds. (Understanding)
- Describe with examples single, double and triple covalent bonds. (Understanding)
- Draw electron cross and dot structures for simple covalent molecules containing single, double and triple bonds. (Applying)

4.0 Introduction

In chemistry the structure determination is the basic and most essential element. The properties of matter lie in the structure of compounds, and structure of the compound is largely influenced by the type of bonding in the compound. A chemical bond is the force of attraction between the atoms that constitute the molecule, this attraction results due to the electrostatic forces produced by the charged particles of atoms. The chemical bonding is associated with the sharing or transfer of electrons between the participating atoms. Molecules, crystals, and diatomic gases indeed most of the physical environment around us are held together by chemical bonds, which dictate the structure of matter. In the forth coming topics you will learn the different types of bonds that are found in the chemical compounds, and also their influences upon the nature of these compounds.

4.1 Why Do Atoms Form Chemical Bonds?

It is the basic tendency of matter to be in the lowest energy state, that is why the atoms react to form molecule or form bonds between them.

In the periodic table the most stable electronic structure is of the noble gases, so all the other atoms react to attain the electronic structure of these noble gases.

For example the Na atom has the electronic configuration, $1s^2 2s^2 2p^6 3s^1$. The nearest noble gas of this configuration is Neon with electronic configuration, $1s^2 2s^2 2p^6$.

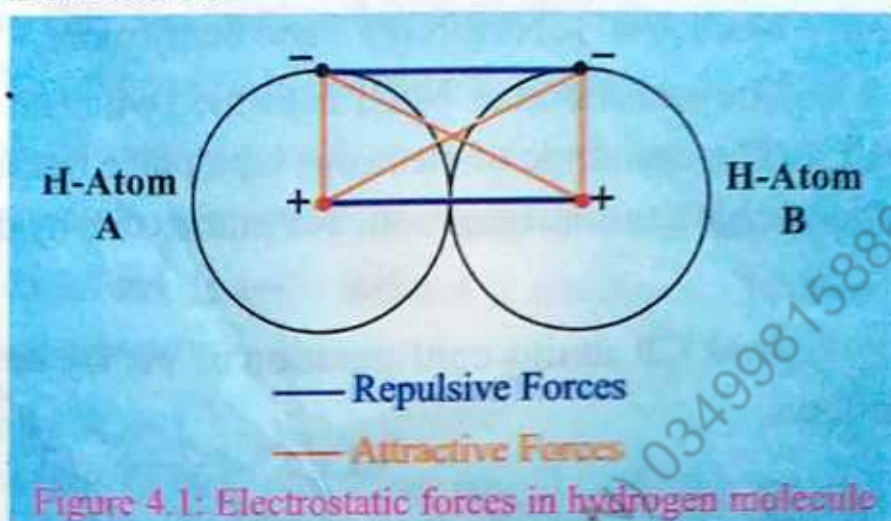
So, the Na tries to attain this configuration when it reacts with the other elements, e.g. in reaction with Cl to form NaCl, the Na losses 1 electron and attains the electronic configuration of Ne, in the same way Cl accepts 1 electron and attains the electronic configuration of his nearest noble gas Ar, during process of formation of NaCl the energy is liberated, and stable NaCl is formed after the spontaneous reaction.

4.2 The Chemical Bond

A Chemical bond is the force that binds the atoms to form the molecules. The chemical bond is the result of attractions among the atoms

caused by the electrostatic forces, (remember that opposite charges attract each other). The electrostatic forces in atoms arise from the charged particles, i.e. the proton and the electron.

These particles form the bond, when the nucleus of one creates attraction with the electrons of the other and vice versa, and hence the atoms are bonded by these forces. Due to the existence of some repulsive forces, i.e. the forces among the nucleus-nucleus and among the electrons-electrons, shown in figure 4.1.



The bond between the atoms forms when the attractive forces are stronger than the repulsive forces.

Although, it seems simple, but in reality this is not the case, and different types of compounds have different situations and the Scientists explain these in totally different ways, i.e. there is no single way to describe all the types of compounds satisfactorily. Some of the important types of bonds are described in coming pages.

4.3 Types of Chemical Bonds

Depending upon their nature there are following 4 types of chemical bonds:

- ♦ The ionic bond
- ♦ The co-valent bond
- ♦ The co-ordinate covalent/dative bond
- ♦ Metallic bond

4.3.1 Ionic Bond

This type of bond is formed when one or more electrons are completely transferred from one atom to the other, as a result of this transfer ions are formed. One of these ions is positively charged and other is negatively ions.

charged. These ions attract each other by strong electrostatic force, and hence ionic bond is formed between these ions. The examples of compounds having ionic bond are:

NaCl, CaCl₂, NaF, NH₄Cl and many more.

The ionic bond in 'NaCl' is formed when an electron transfers from 'Na' to 'Cl'. The exchange of electrons take place because both 'Na' and 'Cl' are to attain noble gas configuration. 'Na' attains this by donating an electron.



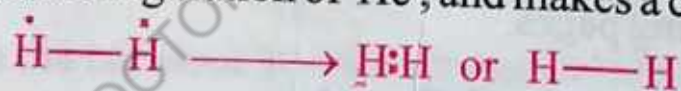
And 'Cl' attains configuration of 'Ar' by obtaining this one electron as follows:



These ions are strongly attracted by electrostatic force forming NaCl. This bond is very strong and compounds having ionic bond have high melting and boiling points.

4.3.2 Covalent Bond

This type of bond is formed when none of atoms are able to donate the electrons completely so they share the electrons, e.g. in case of 'H₂' molecule, each 'H' atom shares an electron with the other hydrogen atom, and hence each 'H' achieves noble gas configuration of 'He', and makes a covalent bond.



This can also be represented as in figure 4.2:

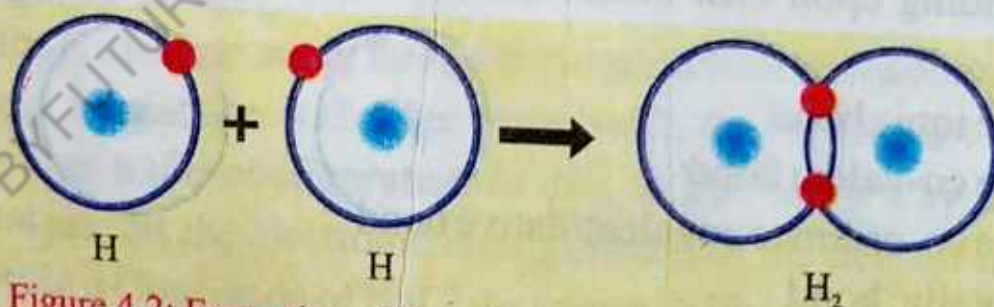


Figure 4.2: Formation of Covalent bond in hydrogen molecule

Other such covalent compounds are O₂, CO₂, H₂O, CCl₄, CH₄, HCl, SiO₂ etc.

It is to be remembered that during the formation of a covalent bond the

atoms tend to attain nearest noble gas configuration and generally they attain such electronic structure that there remains 8 electron in the valence shell of each bonded atom. This is called **octet rule**, but there are also exceptions of this rule e.g., the first period elements i.e. hydrogen gets just 2 electrons in their valence shell during covalent bond formation, this is because of the smaller size of hydrogen and availability of only first shell whose ability to carry the electrons is only limited 2 electrons, for this hydrogen this rule is called **duplet rule**.

4.3.2.1 Types of Covalent Bond

The Covalent bond can be further classified into 3 different types, depending upon the numbers of electron pairs donated. The first type of the covalent bond is the simple covalent bond which is formed by the sharing of single electron pair, this is also known as the single covalent bond. Such covalent bonds are formed between H_2 , F_2 , HCl , CH_4 , Cl_2 etc. (See figure 4.3)

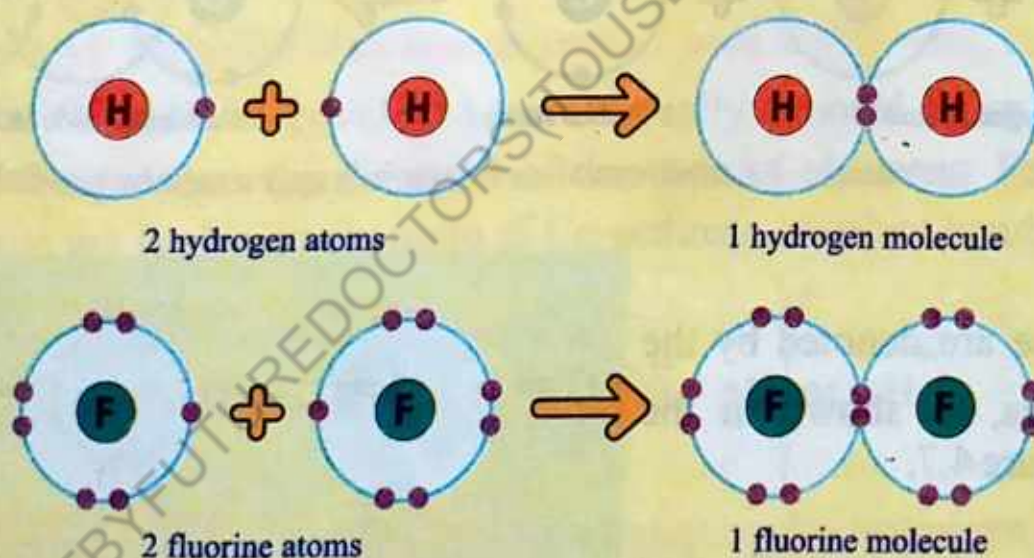


Figure 4.3: Formation of single covalent bonds in hydrogen and fluorine

For ease a single line is written between the two bonded atoms, as in figure 4.4. Or it can also be denoted by the electron cross and dot structures, as in figure 4.5. (Note that the dot and cross are used just to express the electrons of different atoms, otherwise all electrons are alike.) Such cross-dot structure of molecules are called lewis structures.

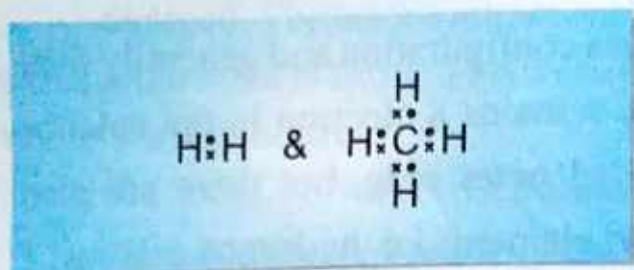


Figure 4.4: Dot and cross representation of hydrogen and methane molecules

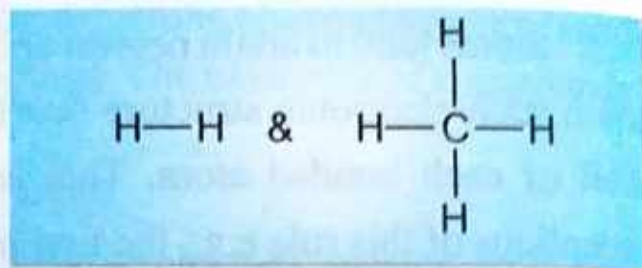


Figure 4.5: Line covalent bonding representation in hydrogen and methane molecules

The second type of covalent bond involves the sharing of two electron pairs, that is why it is called double covalent bond, e.g. in CO_2 , O_2 , C_2H_4 (ethene), see figure 4.6:

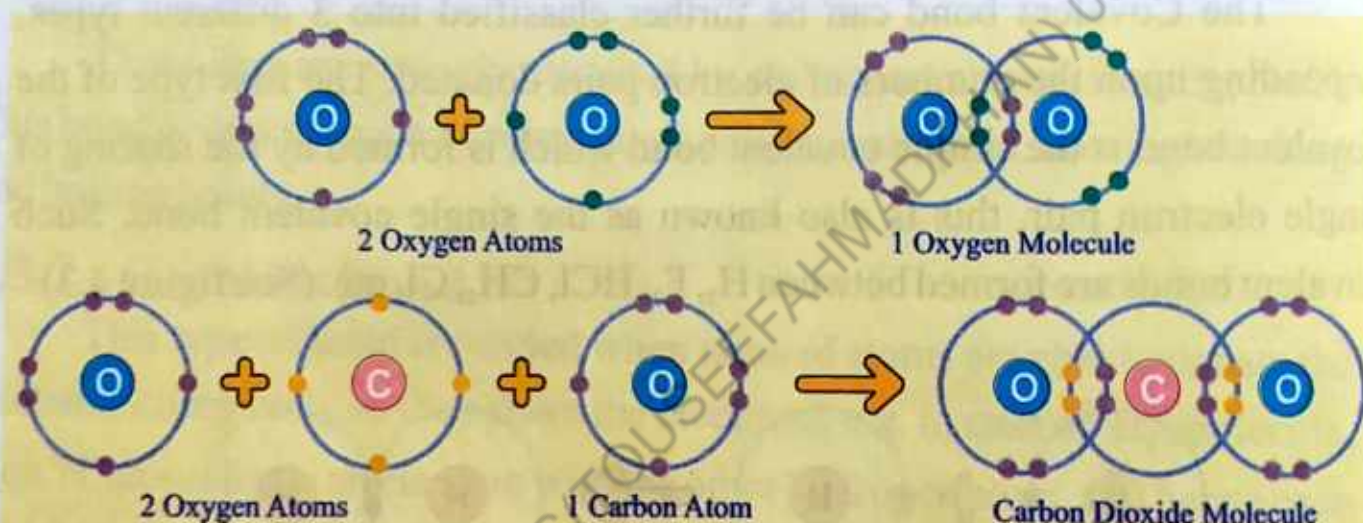
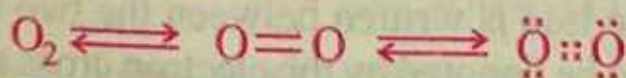
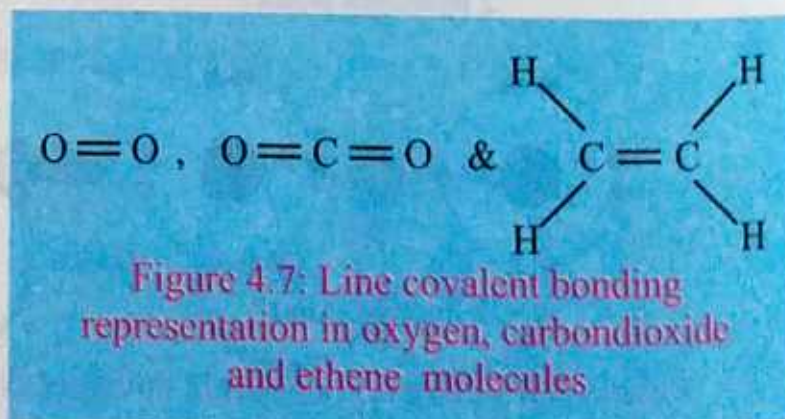


Figure 4.6: Formation of double covalent bonds in oxygen and carbondioxide

These are denoted by the double lines, as shown in the cases in figure 4.7.

The cross and dot structure of O_2 is:



The third type of covalent bond involves 3 pairs of electron sharing, and is known as the Triple covalent bond, e.g. in N_2 and acetylene (C_2H_2).



These are represented by triple lines joining each covalently bonded atoms.

Activity 4.1

Write the electron cross and dot structures for the CO_2 , NH_3 , N_2 and Ethane molecules.

4.3.3 Co-ordinate Covalent/Dative Bond

This is a type of covalent bond in which both electrons of bond (shared electron pair) comes from the same atom, e.g. formation of $\text{BF}_3 - \text{NH}_3$.

Here, 'B' is electron deficient and 'N' has a lone pair of electron, during bond formation this lone pair of electron participates and is shared between 'N' and 'B'.

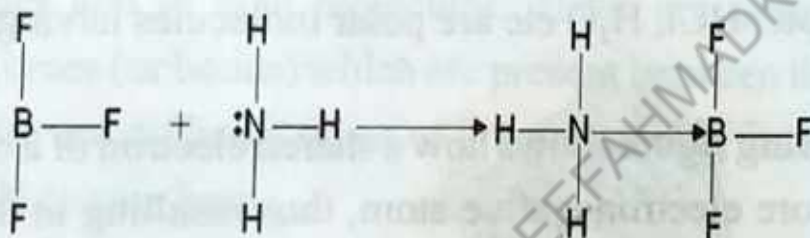


Figure 4.8: Formation of co-ordinate covalent bond in $\text{NH}_3 - \text{BF}_3$ complex

The Co-ordinate covalent bond is usually denoted by an arrow sign which also represents the direction of donation of electrons. Formation of Hydronium ion is also an example of Co-ordinate covalent bond, as seen in figure 4.9.

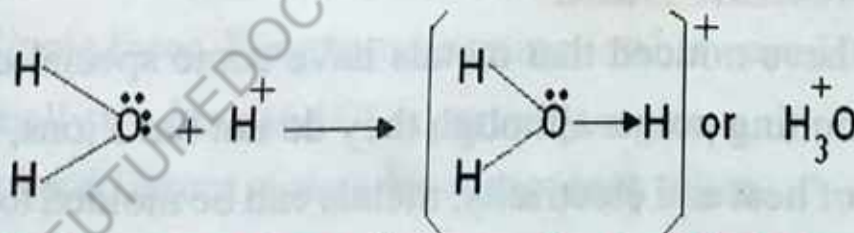


Figure 4.9: Formation of co-ordinate covalent bond in hydronium ion

Another example is the formation of NH_4^+ from NH_3 and H^+ (Figure 4.10)

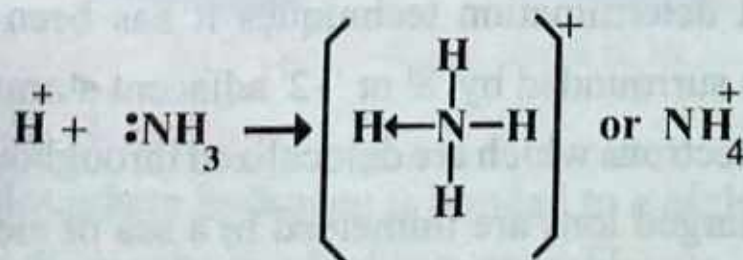


Figure 4.10: Formation of co-ordinate covalent bond in ammonium ion

4.3.4 Polar and Non-polar Bonds

Non-polar Covalent Bond

A covalent bond which is formed between like atoms having same electronegativities is called non-polar covalent bond. In such bond the shared electrons are distributed symmetrically around the bonded atoms.

For example; N_2 , H_2 , O_2 , Cl_2 etc.

Non-polar molecules have non-polar covalent bond.

Polar Covalent Bond

The covalent bond formed between unlike atoms having different electronegativities but the difference in electronegativity would be less than 1.7 but not zero. In such bond the shared electrons are distributed symmetrically around the bond atoms.

For example; HCl , H_2O etc are polar molecules having polar covalent bond.

The following figure shows how a shared electron of a covalent bond is attracted by a more electronegative atom, thus resulting in the formation of polar molecule.



4.3.5 The Metallic Bond

You may have noticed that metals have some special characters e.g. they have high melting points although they do not have ions, the metals are good conductor of heat and electricity, metals can be molded to any shape, all these character show that metals have some special and different type of bonding. Scientists named this bond of metals as metallic bond.

By structural determination techniques it has been revealed that in metals each atom is surrounded by '8' or '12' adjacent atoms. The bonding is due to the valence electrons which are delocalized throughout the entire solid, so the positively charged ions are immersed in a sea of electrons, as can be seen in figure 4.11.

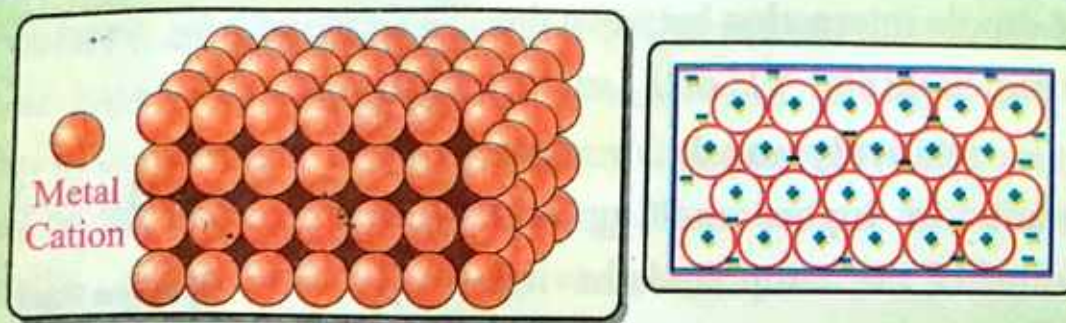


Figure: 4.11: Metallic bonding

Due to these reasons the metals can be drawn into wires, sheets and are very good conductor of heat and electricity. Moreover alloys can be formed easily by them.

4.4 Intermolecular Forces

These are the forces of attraction present between the molecules to bring them closer and to bind molecules. These are quite different from intramolecular forces (or bonds) which are present between the atoms to form a molecule. There are different types of intermolecular forces, of which the important we will discuss here,

1. Dipole-dipole forces
2. Hydrogen bonding

4.4.1 Dipole—dipole Forces

The force of attraction between molecules having permanent dipole is called dipole-dipole force. For example attraction between HCl molecules in which $H^{\delta+}$ is partially positive and $Cl^{\delta-}$ is partially negative.

Due to poles different molecules attract each other.

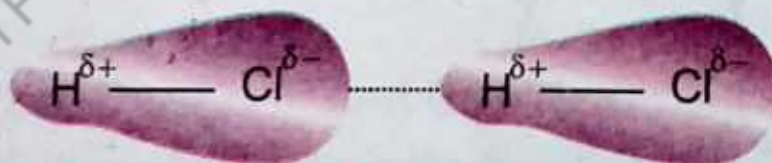


Figure: 4.12: Dipole-dipole forces in HCl

4.4.2 Hydrogen Bonding

The molecules where hydrogen is bonded to a highly electronegative element like 'O' or 'F' results in a highly polar molecule, in such molecules

the dipole-dipole interaction between the 'H' of one molecule and 'O₂' or 'F₂' of 2nd molecule is relatively much stronger, which results in higher melting and boiling points of the compounds having such type of interaction. e.g. water has boiling point '100°C' while H₂S has B.P '-60°C'. Similarly 'HCl' is a gas having B.P '-85°C' as it lack hydrogen bonding while its counterpart 'HF' has B.P='19.5°C'.

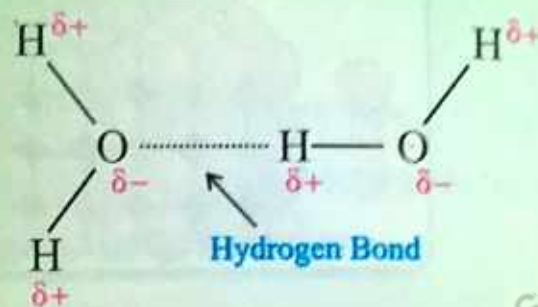
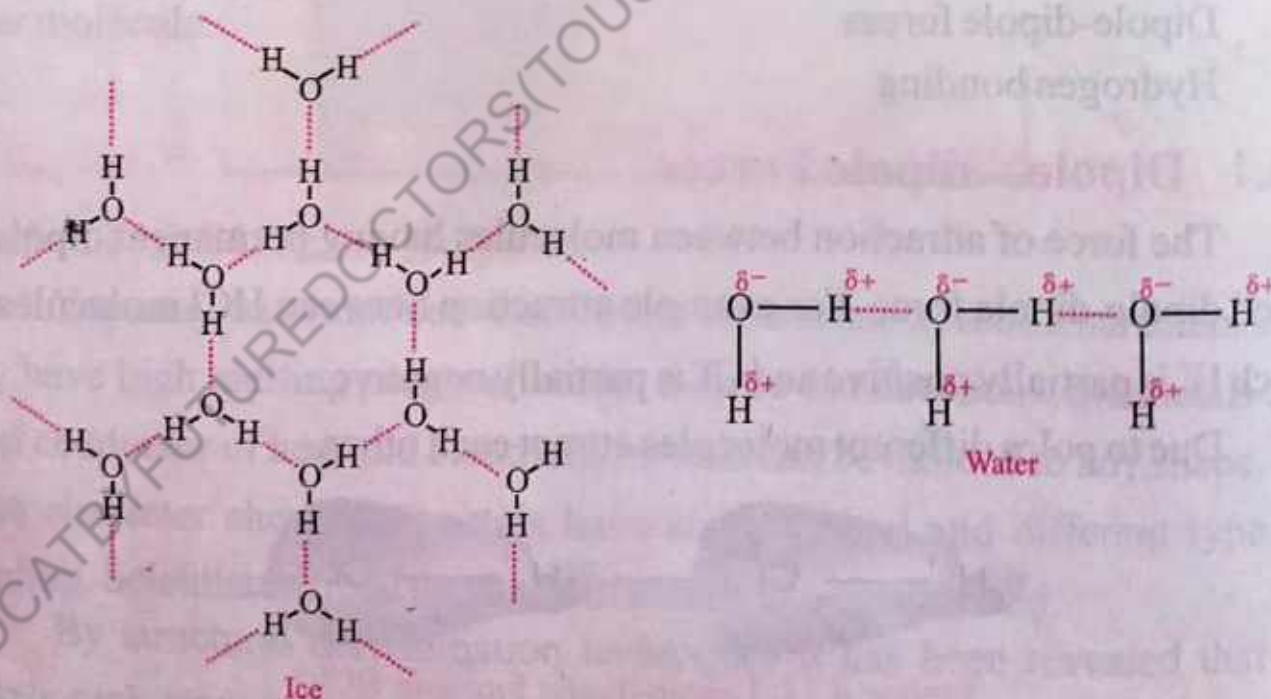


Figure: 4.12: Hydrogen Bonding in Water

The strength of hydrogen bonding changes the physical properties such as boiling point melting point etc.

Interesting Information

This is the hydrogen bonding due to which water expands when freezing, so ice has lower density than water, the picture of arrangement of atoms in ice is shown here.



This property of ice makes the life possible on earth, as ice formed in cold regions of earth floats at surface due to having low density and the life underneath grows without suffering from outer low temperature, (ice is bad conductor of heat too)

4.5 Nature of Bonding and Properties

The type of bonding greatly effects both physical and chemical properties of compounds. It is so important that generally the compounds are classified on the bases of type of bonding. The physical properties e.g M.P, B.P, density , color, surface tension, viscosity, refractive index, solubility, bond energy, vapour pressure, lattice energy etc may be predicted from knowing the type of bonding. Effect of bonding upon various types of compounds is described in following section individually.

4.5.1 Nature of Bonding and Ionic Compounds

The ionic compounds which have a strong type of bonding are hard, solid and have very high melting and boiling points.

The ionic compounds although have ions but they conduct electricity only in molten state, because of the limited mobility in solid state and high mobility in liquid state. The ionic compounds which are strongly polar are generally soluble in the polar solvents like water, for example NaCl is easily soluble in water to large extent.

As concerned with reactions, bonding also plays an important role here. Generally ionic compounds are more reactive and the rate of reaction is very fast.



4.5.2 Nature of Bonding and Covalent Compounds

The co-valent compounds have sharing of electrons and thus there does not exists strong interactions among the constituting particles like that of ionic compounds, hence they have low melting and boiling points and mostly the covalent compounds are gases (unless other factors do not involve). The covalent compounds do not conduct electricity unless they are ionized by

some factor, so the pure water do not conduct the electricity but when some ionizing material such as an acid or a base or some electrolyte is added to the water, which makes it ionized and thus the flow of charged particles becomes possible.

4.5.3 Nature of Bonding and Polar and Non-polar Compounds

The behavior of polar and non-polar compounds is well understood in terms of the ionic and covalent compounds, where the polarity also counts greatly in determining the properties of these compounds. The non-polar compounds are mostly gases and have low melting and boiling points, and do not conduct electricity. On the other side the polar compounds have relatively high melting and boiling points and conduct electricity.

4.5.4 Nature of Bonding and Metals

The metals have a very special type of bonding and thus their characters also are very interesting. In the metals there exists a positive center around which is fig. 4.14 a sea of negative sphere exists.

The negative sphere which consists of free electrons (the electrons which move freely in all the available space in metal piece), this gives the metals a high conductor character for both the electricity and heat. Moreover the metals have lustrous surface due to transition of electron.

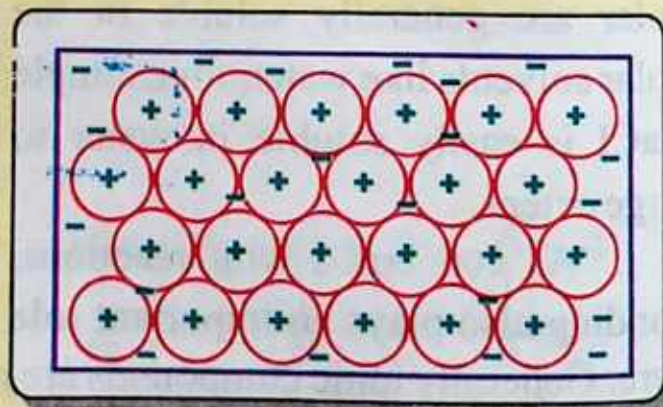


Figure: 4.14: Electron Sea around Cations in Metals

SUMMARY OF THE CHAPTER

- ❖ In order to attain stability the atoms tend to form molecules.
- ❖ The electrostatic force of attraction between atoms in a molecule is termed as the **bond**.
- ❖ There are many types of bonds in the molecules.
- ❖ The bond which is formed by the transfer of electrons between atoms is called the **ionic bond**.
- ❖ The bond which is formed by the sharing of electrons between the atoms is termed as the **Covalent bond**. In a simple covalent bond each atom shares 1 electron.
- ❖ Some times in a covalent bond, the shared electron pair comes from a single atom; such covalent atom is called **Dative covalent bond** or **co-ordinate covalent bond**.
- ❖ When a covalent bond is formed by a single electron pair it is termed as **single covalent bond**.
- ❖ If a covalent bond is formed by two electron pairs it is called **double covalent bond**.
- ❖ When a covalent bond is formed by three electron pairs it is termed as **triple covalent bond**.
- ❖ In a covalent bond when difference of electronegativity between the two atoms is **0.5 or greater** then the bond becomes **polar** and electrons are bent towards more electronegative atoms.
- ❖ Metals have a special type of bonding called the **Metallic bond**, in the metals around the charged positive center an electron sea is present, these electrons are free to move and are called **Free electrons**.
- ❖ The different molecules also show some type of electrostatic attractions, this force between the molecules is called the **intermolecular force**.
- ❖ Different types of intermolecular forces exist, e.g. **dipole-dipole force**,

Hydrogen bonding.

- ❖ The term hydrogen bonding is not like the bond, which joints the atoms in a molecule, but is named such due to its strength. It is present between the hydrogen and some electronegative atom of two different molecules.
- ❖ Both the type of bonding and intermolecular forces affects greatly the properties of different compounds.

EXERCISE

Q1. Fill in the blanks with appropriate words.

- i) A chemical bond is the force which binds atoms to constitute molecules.
- ii) The bond forms when attractive forces the repulsive forces.
- iii) The water molecule has type of bonding.
- iv) The metals have electrons.
- v) The water has bonding, which makes it liquid, although it is covalent.
- vi) The ionic compounds always have physical state.
- vii) Difference in physical states among H_2O and H_2S is because of
- viii) The most electronegative element is
- ix) In the type of bond, both shared electrons comes from the same atom.
- x) Same type of atoms have type of bond.

Q2. Choose the correct answer.

- i) The KCl has the type of bonding
(a) ionic (b) covalent
(c) co-ordinate covalent (d) metallic
- ii) A covalent bond forms by:
(a) transfer of electrons
(b) sharing of electrons

- (c) both by transfer and sharing of electrons
- (d) no transfer and no sharing of electrons
- iii) Generally, the covalent compounds are physically:
 - (a) liquid (b) gas (c) solid (d) plasma
- iv) The metals have characters:
 - (a) solid (b) lustrous (c) conductor (d) all of above
- v) The metals conduct electricity because of:
 - (a) polarity (b) free electrons
 - (c) both by polarity and free electrons
 - (d) Neither polarity nor free electrons are involved

Q3. Answer the following questions in short.

- i) What is a bond? Name different types of bonds formed between atoms.
- ii) How you can differentiate between two compounds having ionic and covalent bonding?
- iii) What type of bond will be formed between following:
 - (a) Na and Br_2 , (b) Ca and Cl_2
 - (c) C and O_2 (d) N_2 and H_2
- iv) What do you know about Lewis structure of a compounds, explain with example.
- v) What different types of covalent bonds are there?
- vi) What is the difference between simple covalent bond and coordinate covalent bond?
- vii) What is meant by polarity of a covalent bond?
- viii) Correlate bonding and special features of chemical compound.
- ix) Heating water would cause breakage of the intermolecular forces or intramolecular forces first.
- x) What types of inter and intra molecular forces are present in H_2O .

Q4. Answer the following questions with reasoning.

- i) Can you explain why similar atoms form covalent bond rather than ionic bond?

- ii) Is hydrogen bonding is a true type of intramolecular attraction? Explain with reasons.
- iii) Explain why water pipes get ruptured during winter in cold regions?
- iv) Why H_2O has boiling point 100°C while H_2S has B.P = -60°C ?
- v) Why NaCl is soluble in cold water but not in boiling oil?

- Q5. What is the difference between the intermolecular force and intramolecular force?
- Q6. Discuss the importance of bonding in establishing the characters of a chemical compound?
- Q7. What is meant by polarity of a covalent bond? Explain with examples.
- Q8. Write a note on types of bonds found in different compounds?.
- Q9. Why the atoms joint together to form a molecule?
- Q10. What are the different types of covalent bonds?
- Q11. Write a note on characteristics of ionic compounds?

PHYSICAL STATES OF MATTER

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the,

- Matter and its states
- Intermolecular attraction
- Simple mathematics methods

TIME ALLOCATION

Teaching periods	= 14
Assessment periods	= 03
Weightage	= 13%

LEARNING OUTCOMES

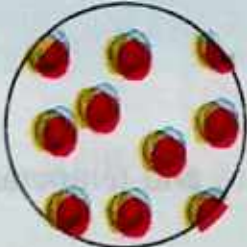
After reading this chapter, the Student will be able to describe the:

- Effect on the volume of a gas by a change in the pressure and temperature. (Understanding)
- Compare the physical states of matter with regard to intermolecular forces present between them. (Analyzing)
- Account for pressure-volume change in gases using Boyle's Law. (Analyzing)
- Account for temperature-volume change in gases using Charle's Law. (Analyzing)
- Explain the basic properties of gases. (Understanding)
- Explain the basic properties of liquids. (Understanding)
- Describe physical properties of solids. (Understanding)
- Explain the effects of temperature and external pressure on vapour pressure and boiling point. (Analyzing)
- Differentiate between amorphous and crystalline solids. (Analyzing)
- Explain the allotropic forms of solids. (Understanding)

5.0 Introduction

Every thing which has a mass and occupy space is called matter. The universe is made up of two entities, matter and the energy. Here in this chapter we will discuss matter and its types in details. Matter exists in four different states, i.e. Gas, Liquid, Solid and Plasma. The plasma state is actually a mixture of gaseous state and ions. Many scientists do not take it as a separate state, moreover in our planet we normally do not notice this plasma, that is why it is not discussed along with other states. Most of the universe is made up of the plasma, as most matter is scattered in stars (or sun).

Following table 5.1 shows the comparison of other three states of matter.

COMPARISON BETWEEN DIFFERENT STATES OF MATTER		
GASES	LIQUIDS	SOLIDS
		
It is the form of matter which do not have definite shape nor definite volume.	They do not have definite shape, but volume is definite.	Both shape and volume are definite
Molecules lie far apart from each other	Molecules lie little close to each together	Molecules are very close to each other.
Molecules are free to move.	Molecules can move in a limited way.	Molecules have only oscillatory movement at their own place.
K.E of molecules is highest	K.E is intermediate.	K.E is lowest.
Intermolecular forces are very weak.	Intermolecular force are fairly stronger.	This force is very strong.
Molecular arrangement is irregular.	Molecular arrangement is some what regular	Molecules are highly arranged.

COMPARISON BETWEEN DIFFERENT STATES OF MATTER

GASES	LIQUIDS	SOLIDS
Molecules can be compressed easily.	Molecular can be compressed up to a few degree.	Molecular can not be compressed.
Rate of diffusion is very high.	Rate of diffusion is lower than gases.	No diffusion.

Table 5.1: Comparison of various states of matter

5.1 The Gaseous State

The gaseous state, as described earlier is the simplest state of matter, having following general properties.

5.1.1 Typical Properties

In gases the molecules are far apart from each other due to weak intermolecular forces. (Figure 5.1)

The gas molecules spread in all available space, i.e they do not have a volume of their own, but the volume of the container where they are kept is counted as their volume.

Usually the gases are colourless and so they are invisible, but some gases have colour too, and they can be seen, e.g. the Cl_2 gas is yellowish green and the NO_2 gas is reddish brown are shown in figure 5.2.

In the gaseous state of matter, the kinetic energy (K.E) of the molecules is highest, (as you have observed many times in the steam which is the gaseous state of water), this is due to the high velocity of gas molecules than compared in liquid or solid states.

Some of these general characters are discussed in detail here in below:

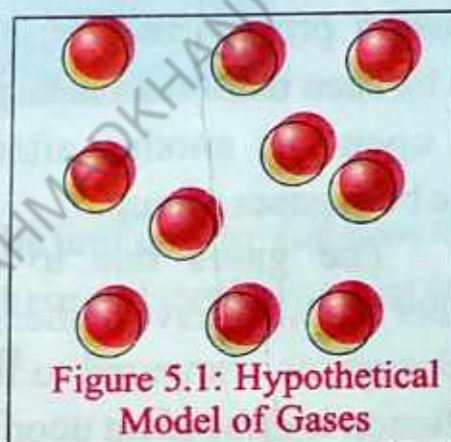


Figure 5.1: Hypothetical Model of Gases



Figure 5.2 Coloured gases

5.1.2 Diffusion

An important character of gases is the high rate of diffusion. Diffusion is defined as the mixing of gas molecules into each other.

You have observed diffusion process of gases many times in your routine life, e.g. when a perfume is sprayed in a room, it readily spreads across the room, and even it leaks out of the room.

The figure 5.3 shows diffusion process in two gases, it can be seen that when both jars are put upon one another after some time both gases mix up.

The gases due to having higher mobility have higher rate of diffusion, as compared to liquids which have slower rate of diffusion. The diffusion is dependent upon the temperature, because at high temperature the speed of molecules is more so the rate of diffusion is also more at higher temperature.

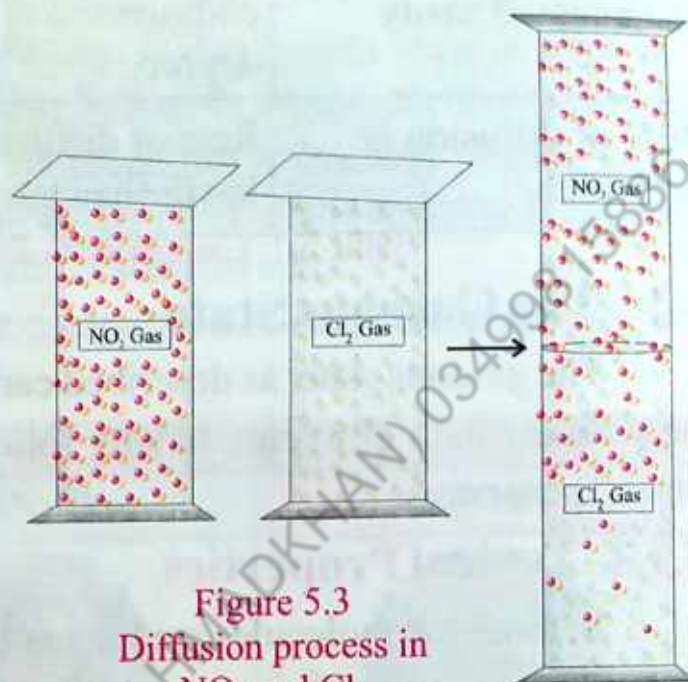


Figure 5.3
Diffusion process in NO_2 and Cl_2

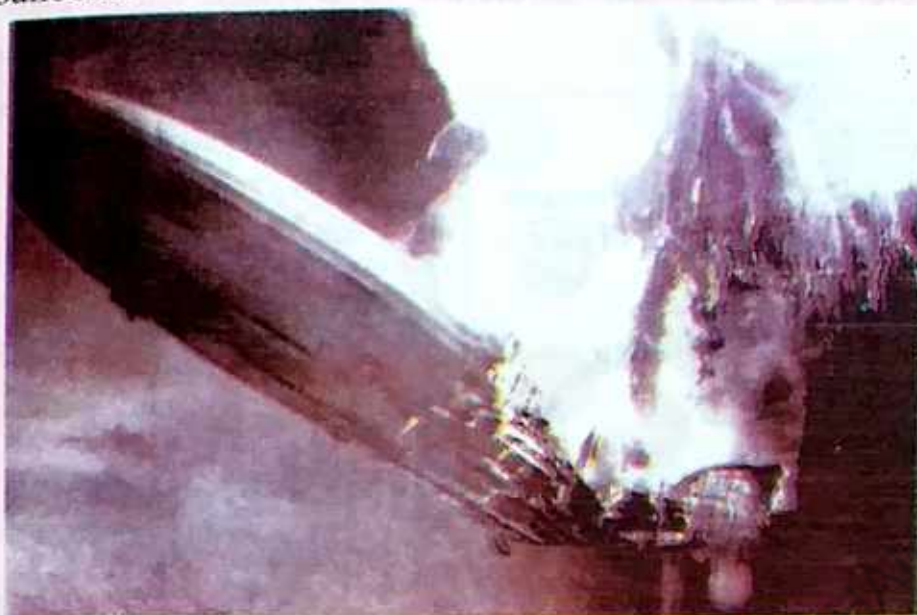
5.1.3 Effusion

This property is similar to diffusion, but the term effusion is specifically used when the gas escapes through small pores of molecular size. You have observed effusion while playing with the hydrogen balloons (gas balloons) in your childhood, when these gas filled balloons are kept over night, they shrink in size overnight.

Interesting Information

Hydrogen is the lightest gas and is also smallest in molecular size as compared to other gases. It is a best fuel which does not produce any harmful products. Once hydrogen gas has been used extensively for filling traveling air balloons and in air ships, but due to its small molecular size it leaks out from molecular sized holes in the containers and this has resulted in accidents many times, a famous air crash which is believed by such incident occurred in 1937 at New Jersey, killing all 36 persons aboard.

This accident stop the use of such air ships for traveling and now a days helium or hot air is filled in air balloons.



5.1.4 Pressure

Pressure is defined as the force per unit area. Its unit is N/m^2 (or Pascal) in S.I system. The commonly used units are atmosphere and mmHg or torri cell.

Like liquids and solids gases also exert pressure, and you may have noticed this pressure many times in your life e.g. while comparing the mass of an empty and filled home gas cylinder and you may have also noticed this pressure while filling air tyres and tubes using pump.

We are living in an atmosphere of air around us and this is obvious that if a little quantity of gas present in the gas cylinder exerts pressure than this air around us must also exert pressure on us much more than the gas of cylinder.

Air which is the mixture of gases also exerts pressure.

The atmospheric pressure is determined by a simple barometer; (figure 5.4) which is a graduated glass tube closed from one end and is

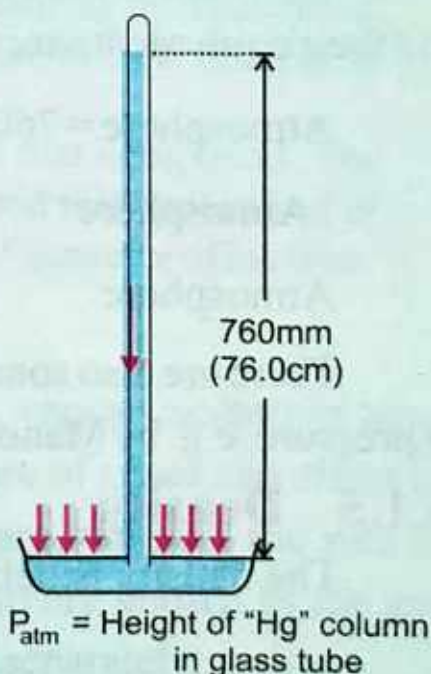


Figure 5.4
Simple Barometer

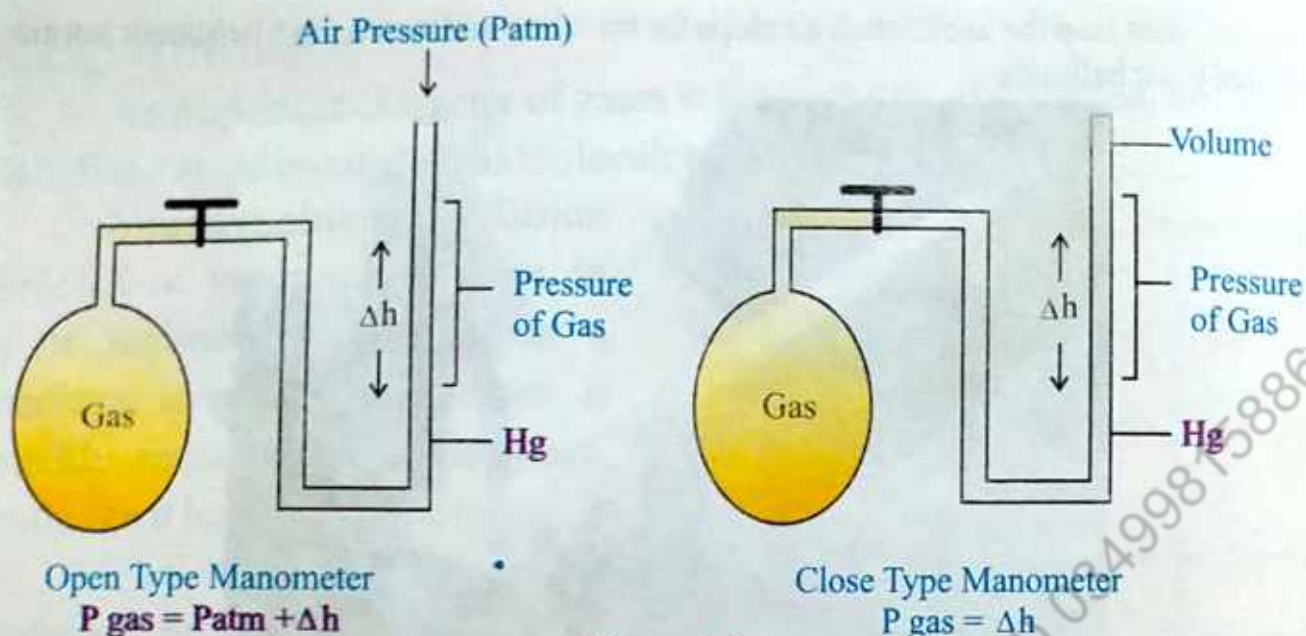


Figure 5.5

filled with mercury (Hg), when inverted in a mercury container the level of (Hg) falls up to the certain height. The height of mercury (Hg) retained in the tube indicates atmospheric pressure, this apparatus was designed by an Italian scientist Torricelli.

There are some other units too which are used for pressure, these units and their equivalents are following:

$$\text{Atmosphere} = 760\text{mm of Hg} = 760\text{Torr}$$

$$1 \text{ Atmosphere} = 101325 \text{ Pascal}$$

$$\text{Atmosphere} = 1.01325 \text{ Bar}$$

There are also some other methods available for accurate measurement of pressure. e.g. by Manometer, which is shown in figure 5.5.

5.1.5 Density

The density is defined as the mass per unit volume of a substance.

$$d = \frac{m}{v}$$

The density of a substance is the measure how the mass is distributed in a unit area, thus the iron has a greater mass per unit area as compared to cotton and therefore cotton has less density than iron.

You can see the difference in the densities of two substances in the given figure 5.6, as the number of particles per unit area increases the density also increases.

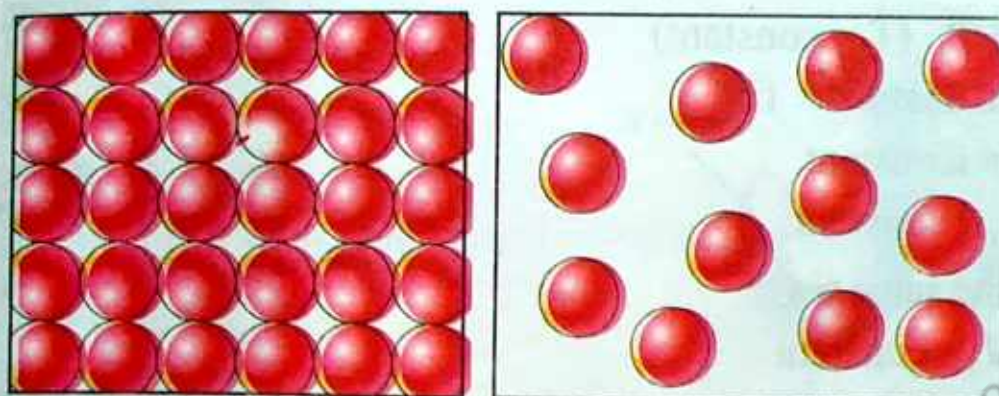


Figure 5.6 Comparison of densities

The density of a substance is dependent upon the temperature too. with the increase of temperature the density decreases.

Interesting Information

The density of a substance in gases state is least while in solid state it has the highest value (water is an exception).

The density also has a relation with the temperature, as the temperature increases density decreases, e.g., all measuring instruments of lab show a temperature (as can be seen in the figure 6.2 of volumetric flask).

Density of a compound is a characteristic of that compound. The famous Greek scientist Archimedes (287-212BC) used this application of density to find out the impurity amount in the crown of Emperor of his time.

5.2 Laws related to Gases

In order to understand the behavior of gases, various scientists have proposed laws, these laws generally explain the nature of gases and effect of variations in the conditions like temperature, pressure, volume, the rate of diffusion etc upon the gases. These laws are collectively known as the gas laws. In the following pages, these laws are described separately.

5.2.1 Boyle's Law (Volume-Pressure Law)

In 1662 Robert Boyle studied the relation between volume and pressure of a gas at a constant temperature.

According to this law the volume of given mass of a gas is inversely proportional to the pressure at constant temperature, i.e.

$$V \propto 1/P \quad (T = \text{constant})$$

$$V = \text{Constant} \times 1/P$$

or $PV = \text{Constant}$

$$P_1 V_1 = \text{constant}$$

In the same way,

$$P_2 V_2 = \text{constant}$$

So,

$$P_1 V_1 = P_2 V_2$$

Where P_1, V_1 are initial pressure and volume while P_2 and V_2 are final pressure and volume.

This is the mathematical representation of Boyle's law.

Experimental Verification

Boyle proved his law experimentally by using a "J" shaped glass tube. He observed that by doubling the height of mercury (Hg) in tube reduced half the volume of gas entrapped,

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

This is visible in figure 5.7

In case 1

Volume (V_1) = 100ml

Pressure (P_1) = 200mm of Hg

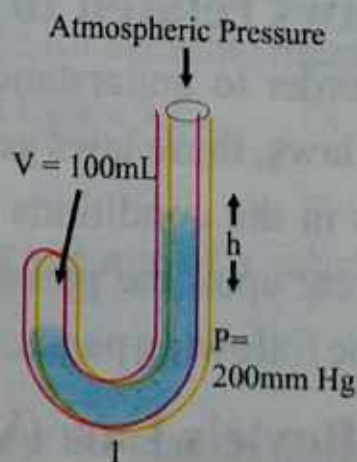
According to Boyle's Law

$$P_1 V_1 = \text{Constant}$$

$$P_1 V_1 = 100 \times 200 = 20000$$



Robert Boyle
(Ireland)
Born 1627—Died 1691



In case 2

(By doubling the pressure by adding mercury, therefore)

$$V_2 = 50\text{ml}$$

$$P_2 = 400\text{mm of Hg}$$

$$P_2 V_2 = \text{Constant}$$

$$50 \times 400 = 20000$$

i.e.

case 1 = case 2

or

$$P_1 V_1 = P_2 V_2,$$

Thus Boyle's law is verified.

So,

$$PV = \text{Constant}$$

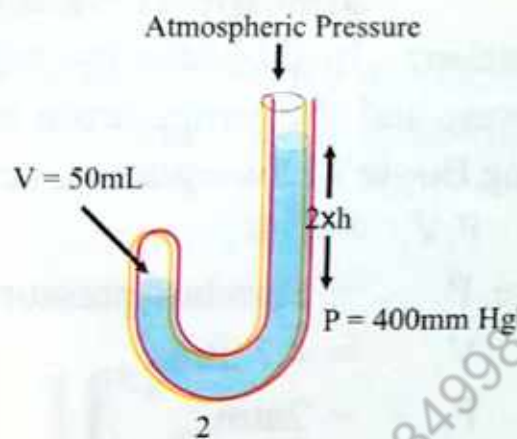


Figure 5.7:
Boyle's Law
experimental Verification

Boyle's Law graphical Representation

By plotting a graph between pressure and volume a curve is obtained indicating that volume decreases with the increase in pressure as shown below in figure 5.8, graph (a).

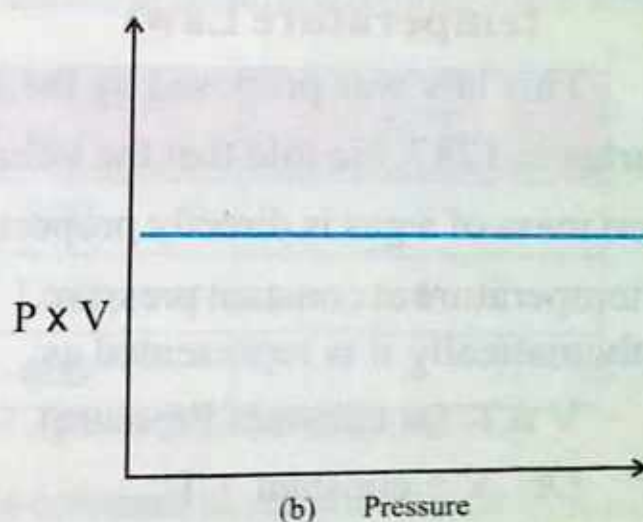
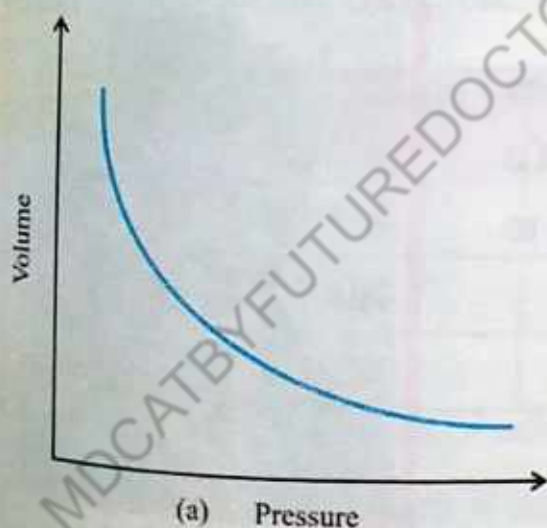


Figure 5.8: Boyle's Law Graph

If a graph is plotted between pressure and $P \times V$, a straight line parallel to pressure axis is obtained, as shown below in figure 5.8(b).

Example 5.1: Volume of a gas at standard pressure is 4.7 litres, what will be the volume of gas when the pressure is increased to 2 atmosphere while the temperature remains unchanged.

Solution: In this case the relation between volume and pressure is in interest and the temperature is constant, so this problem can be solved using Boyle's Law equation, i.e.

$$P_1 V_1 = P_2 V_2 \quad (\text{Temperature} = \text{Constant})$$

Here, P_1 = Standard pressure = 1 atm

$$V_1 = 4.7 \text{ litres}$$

$$P_2 = 2 \text{ atm}$$

$$V_2 = ?$$

$$\text{As } P_1 V_1 = P_2 V_2$$

$$\text{So, } V_2 = \frac{P_1 V_1}{P_2}$$

by putting values:

$$V_2 = \frac{1 \times 4.7}{2} = 2.35 \text{ litres}$$

So, the volume of gas at 2 atmospheric pressure will be 2.35 litres.

5.2.2 Charle's Law (volume-temperature Law)

This law was proposed by the Jacques Charles in 1787. He told that the volume of a given mass of a gas is directly proportional to the temperature at constant pressure. Mathematically it is represented as:

$$V \propto T \quad (\text{at constant Pressure})$$

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

$$\text{Or } V/T = \text{constant}$$

$$\text{Or } V_1/T_1 = V_2/T_2$$

Where V_1, T_1 are initial values of volume



Jacques Charles
(France)
Born 1746—Died 1823

and temperature while V_2 and T_2 are final values of volume and temperature.

Experimental Verification

Charle's used a cylinder fitted with a frictionless piston, as shown in figure 5.9. He observed that the proportion of change in volume at temperature would be the same i.e. by doubling the temperature, volume of gas also doubles.

Mathematically it is represented as:

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

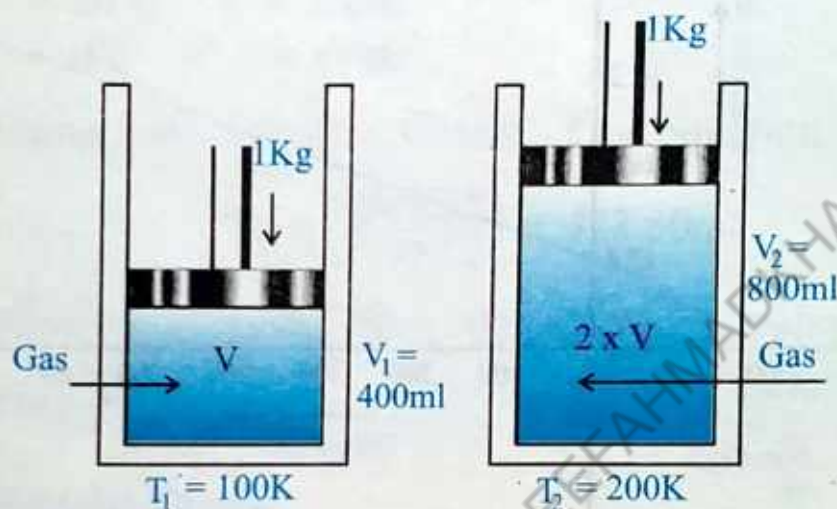


Figure 5.9: Experimental verification of Charle's Law

Consider a given mass of gas in frictionless piston of a cylinder at 100K having volume 400ml, therefore, new volume at different temperature would be:

S.No.	New Temperature (K)	New Volume (ml)	V/T
1	200K	800	4
2	300K	1200	4
3	150K	600	4

Table 5.2: The values of volume of a gas at different temperatures

Thus the ratio between V and T is constant in all conditions.

Graphical Representation

While plotting a graph between ' T ' and ' V ' a straight line is obtained and when this line is increased downwards it touches zero volume line

at -273.15°C , which means that volume at this point is zero which is not possible, so this temperature is not achievable. From here we start a new scale in which zero or negative values are not possible. This new scale is called Kelvin scale, as shown in figure 5.10.

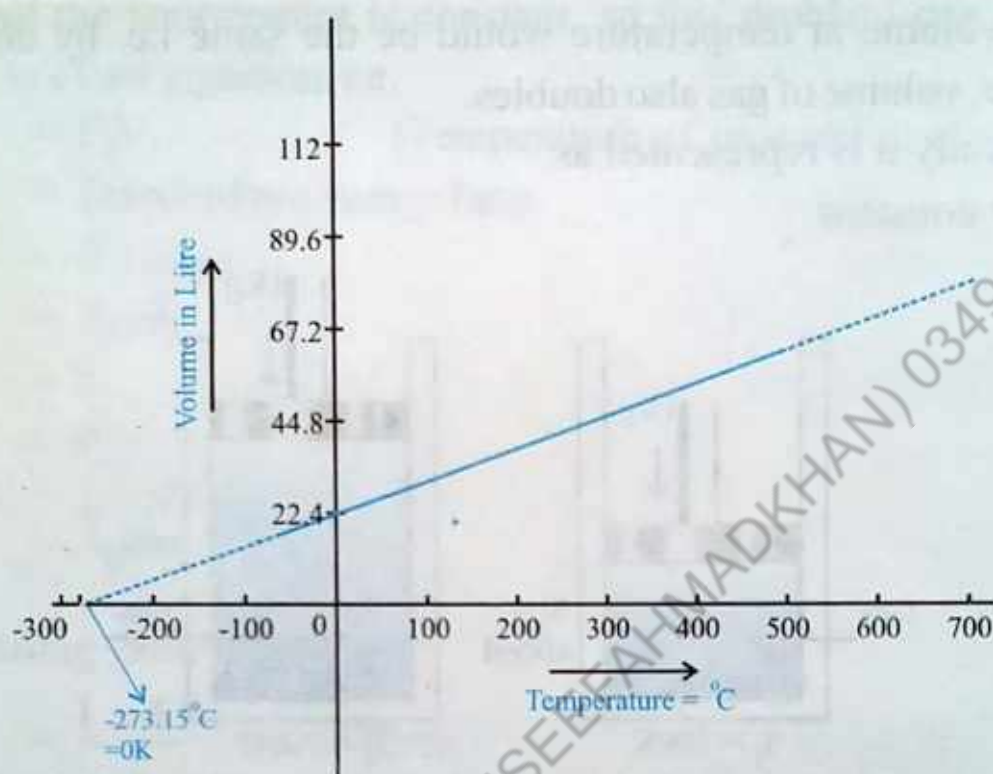


Figure 5.10: Graphical representation of Charles's Law.

The Kelvin scale of temperature is denoted by the symbol 'K', and note that no degree sign is used in this scale.

$$0\text{K} = -273^{\circ}\text{C}$$

The Kelvin scale is needed for calculations specifically, because in this scale the temperature cannot be in negative values or even zero.

Activity: 5.1

Normal human temperature is 37°C express this temperature in Kelvin scale.

In order to change centigrade value of temperature to Kelvin scale just 273 is added.

$$\text{K} = ^{\circ}\text{C} + 273$$

Example 5.2: The volume of a gas at 50°C and standard pressure is 500ml. If the temperature drops to 0°C at same pressure, what will be the new volume?

Solution: In this example the pressure is constant and the relation between volume and temperature is in interest, so here Charles's Law

operates, using Charles's law equation this problem can be solved easily:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad (\text{Pressure} = \text{Constant})$$

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

The given data of example is:

$$V_1 = 500\text{ml}$$

$$T_1 = 50^\circ\text{C} = 323\text{K}$$

$$T_2 = 0^\circ\text{C} = 273\text{K}$$

By putting these values in Charles's Law equation, we get:

$$V_2 = \frac{500 \times 273}{323} = 422.60\text{ml}$$

So, if 500ml gas is cooled from 50°C to 0°C , the volume of gas will be reduced to 422.60ml.

5.3 The Liquid State

The liquids are one of the states of matter which do not have the definite shape but have definite volume.

In the liquid state, the molecules are loosely held, have limited movement in a definite region, so this makes them having a definite volume, but the molecules can not leave the system as in the gaseous state so their volume is definite (See fig. 5.11). The molecules have lesser kinetic energy than the gases due to lesser mobility.

The Liquids have stronger intermolecular forces than the gaseous state, and hence they have comparatively higher melting and boiling points.

Let's discuss some of the basic characteristic properties of liquids in a little detail.



Figure 5.11:
A hypothetical view of
molecules in liquid state

5.3.1 Evaporation

It is the common character of liquids. It is observed that when a liquid like water or petrol is placed in an open container it vanishes away by changing into vapours. This process is called evaporation.

Evaporation is because when molecules of liquid gain energy from surroundings, when the kinetic energy of some molecules becomes enough high to break the intermolecular forces, then molecule enters into gaseous state.

Evaporation occurs at all temperatures, but it has a direct relation with temperature, i.e. when temperature increases the rate of evaporation also increases. Moreover liquids having lower intermolecular attractions are evaporated easily.

Evaporation of a liquid is also increased when the surface area increases, that is why a china dish is used for evaporating rather than a beaker in the labs.

5.3.2 Vapour Pressure

When a liquid is placed in a closed container, some of its molecules vaporize and exert pressure on liquid. Some of these molecules condense after losing their K.E back into liquid state, when rate of evaporation and rate of condensation is equal, the pressure at this stage is called vapour pressure or equilibrium pressure.

The vapour pressure is a characteristic property of

Interesting Information

Due to evaporation process, the vapours from sea comes out and travel hundreds of kilometers away, cool down and fall (condense) in the form of rain or snow, thus life on planet continues.

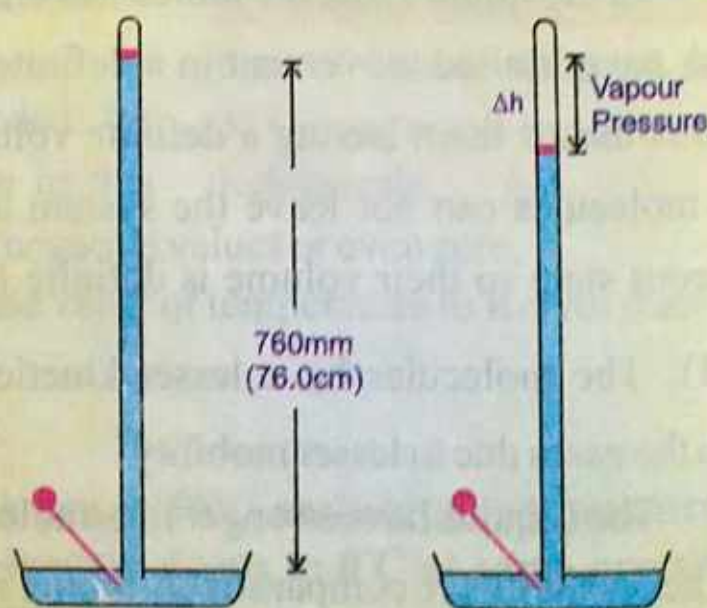


Figure 5.12: A simple mercury Barometer used for vapour pressure determination

liquids but is temperature dependent, and increases with increase in temperature.

The liquids having low intermolecular forces have higher vapour pressure. Vapour pressure can be determined by using simple barometer, as shown in figure 5.12.

The liquid whose vapour pressure is to be determined is inserted by dropper at the bottom of glass tube. The change in height of column gives vapour pressure of liquid (Δh).

5.3.3 Boiling Point

The temperature at which vapour pressure of a liquid and atmospheric pressure becomes equal is called its boiling point. At this temperature the liquid's molecule vaporizes and enters into gaseous state, and its bubbles can be seen arising from bottom moving upward (Figure 5.13).



Figure 5.13: Boiling process in a liquid

The boiling point of a liquid depends upon the atmospheric pressure. If this pressure decreases boiling point also decreases and vice versa.

The following graph shows relation between atmospheric pressure and vapour pressure of different liquids at changing temperatures.

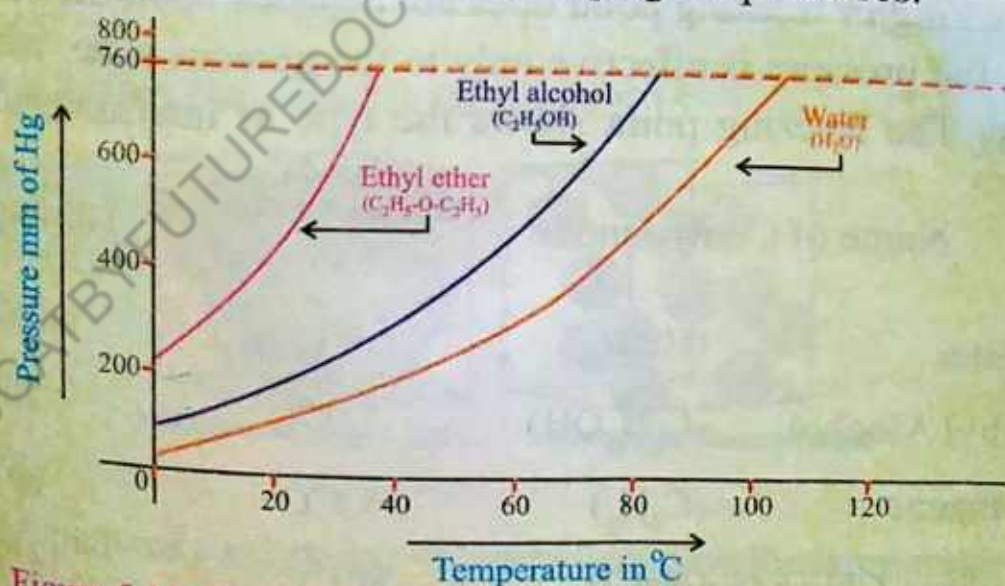


Figure 5.14: Graphical representation of relation between boiling point and atmospheric pressure.

Interesting Information

When the boiling point of a liquid is achieved then further heating does not increase the temperature but only vaporizes the liquid. Amount of heat required to vaporize one mole of a liquid at its boiling point is called molar heat of vaporization, and is characteristic for each liquid e.g. for water it has a value of 40.7KJ/mole.

The boiling point has wide applications of separating mixtures and identification of various chemical compounds e.g. the petroleum in oil refineries is purified to different constituents on the basis of their different boiling points. The table 5.3 shows boiling points of some common liquids.

Interesting Information

Boiling point of a liquid can be decreased or increased by changing external pressure e.g. in pressure cookers boiling point of water is increased above than 100°C and thus at higher temperature cooking time of meat and pulses is reduced.

5.3.4 Freezing Point

The freezing point of a liquid is the temperature at which a liquid freezes to solid state. The freezing point is same as that of the melting point of solid to a liquid. The melting or freezing point is a characteristic property for a compound, e.g. the water has value for both of these "0°C". Unlike the boiling point, the melting or freezing point does not depends upon the atmospheric pressure, as the pressure is effective only to the gaseous state, which is not present here. The freezing point shows the type of intermolecular forces

S/No.	Name of Compound		Melting or Freezing Point	Boiling Point
1	Water	(H ₂ O)	0°C	100°C
2	Ethyl Alcohol	(C ₂ H ₅ OH)	-115°C	78°C
3	Benzene	(C ₆ H ₆)	5.5°C	80°C
4	Sodium Chloride	(NaCl)	801°C	1413°C

Table 5.3: Some common compounds and their melting and boiling points

present in the compound. For the compounds that have stronger intermolecular forces have high values of freezing points as compared to the compounds that have weaker intermolecular forces and have low freezing points.

Table 5.3 shows freezing points of some common compounds along with their boiling points.

5.3.5 Diffusion in Liquids

The diffusion is a property of liquids just like of gases, this refers to the mixing of molecules with one another. The gases have higher mobility and so their rate of diffusion is also higher. Liquids have slower rate of diffusion.

The diffusion is dependent upon the temperature, because at high temperature the speed of molecules is more so the rate of diffusion is also more. The diffusion is easily observed in a glass of water, when a drop of ink is placed in it, and the process when observed at different temperatures can also shows the effect of temperature upon the diffusion.

5.3.6 Density of Liquids

The density which we defined as the mass per unit volume of a substance is also a prominent character of liquids.

The liquids have more density as compared to the gaseous state, which is because of the close distribution of molecules of liquids as compared to the gaseous state.

This is further elaborated in the figure 5.15.

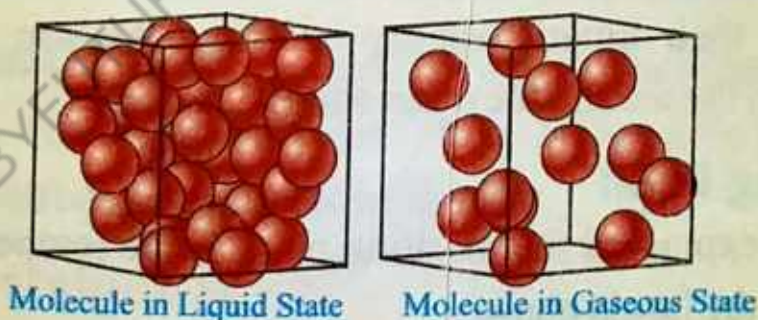


Figure 5.15

The density of a substance is dependent upon the temperature too, with the increase of temperature the density decreases, This is shown by following table 5.2 which shows the density values of water at different temperatures.

SNo.	Temperature °C	Fahrenheit °F	Density (g/cm ³)
1	0.0	32.0	0.9998
2	4.0	39.2	0.9999
3	15.0	59.0	0.9991
4	20.0	68.0	0.9982
5	25.0	77.0	0.9980
6	37.0	98.6	0.9933
7	100	212.0	0.9583

Table 5.2: Density values of water at varying temperatures

5.4 The Solid State

The state of matter that has definite shape and definite volume is called the solid state. Some typical properties of solids are given below.

5.4.1 Typical Properties

The solids are the form of matter having a definite shape and volume. In solids the particles lie very close to each other and have very high intermolecular forces, due to which they cannot change their location so both the shape and the volume here remains unchanged. In the solids the molecules move just vibratory, so their kinetic energy is least. (As you can observe the difference between the ice and the other states of water).

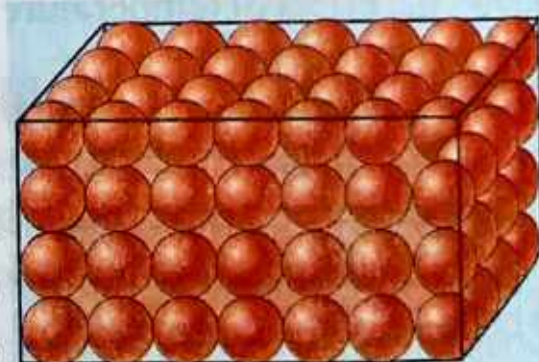


Figure 5.16: A hypothetical view of molecules in solid state

5.4.1.1 Melting Point

As already explained that melting point is the temperature at which a solid changes into the liquid phase.

The molecules in the solid state lie close to each other and thus they have high values of intermolecular attractions, therefore it becomes hard to separate them as compared to the liquids. Thus they have high and sharp melting points.

5.4.1.2 Rigidity

In the solid state, the molecules lie in close association to each other, so deformity in their structure is not easy, that is why they are rigid.

5.4.1.3 Density

The density in the solid phase is also highest because the constituting particles lie very close to each other and more mass is concentrated as compared to the liquid and the gaseous states.

5.5 Types Of Solids

The solids are further classified into two main sub types; these are discussed in the next sections.

5.5.1 Amorphous Solids

The word amorphous has its origin from two parts where 'A' stands for without and morph for shape, thus amorphous means without shape.



Figure 5.17 some amorphous solids

These are the solids which do not have a characteristic shape, e.g Glass, some plastics, wax and gem stones. These are actually not true solids and are sometimes referred as **super cooled liquids**.

5.5.2 Crystalline Solids

The crystalline solids have a characteristic geometrical shape. The crystalline solids are actually the true solids. The crystals have all characters by which a solid is recognized, e.g. the rigidity, elasticity, sharp melting point etc.

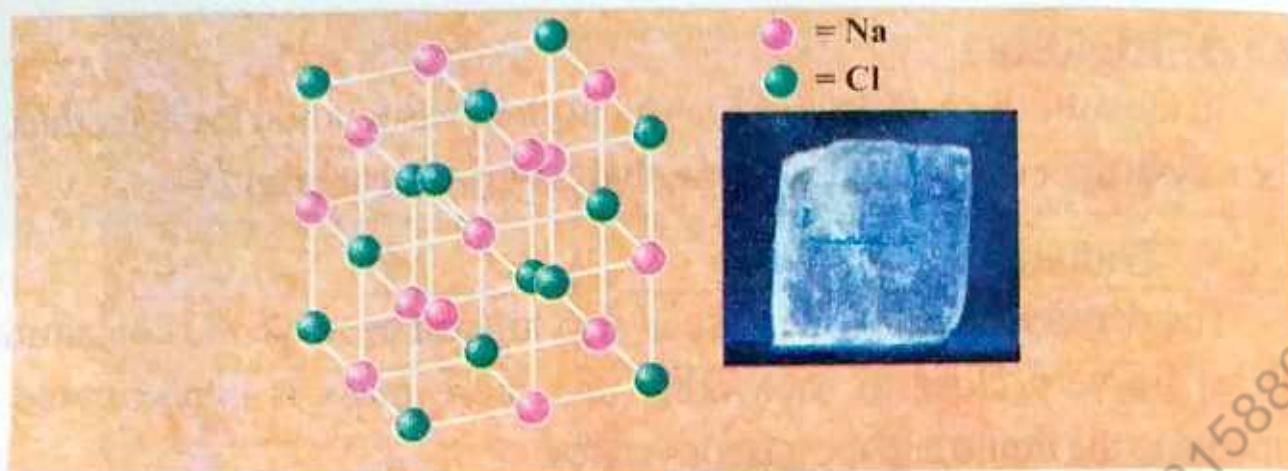


Figure 5.18: The NaCl crystal

5.6 Allotropy

The word allotropy has been derived from two parts, where allos stands for other and tropy (from tropos) means forms, thus allotropy means other forms.

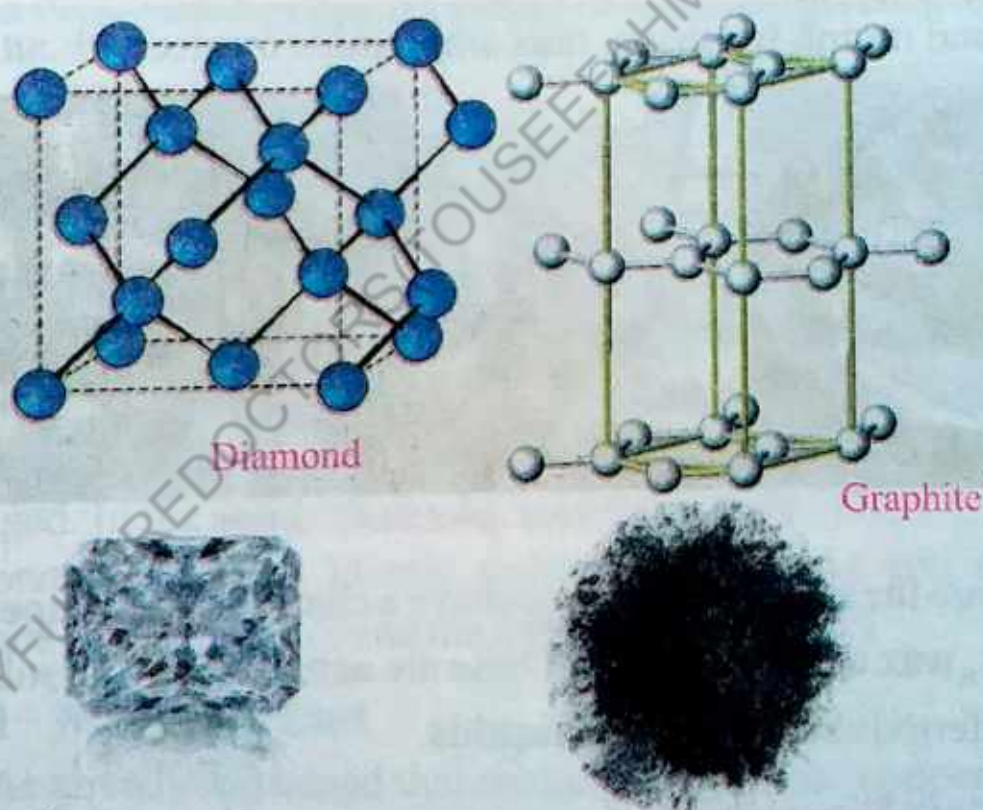


Figure 5.19 Allotropic forms of Carbon with their molecular structures.

The Allotropic forms are the different forms of an element, which have although the same chemical properties, but different structures. For example the famous diamond and graphite are the allotropic forms of carbon.

Other Examples of allotropic forms:

i) Oxygen

- Oxygen O_2 - colourless
- Ozone O_3 - blue

ii) Sulfur

- Plastic (amorphous) sulfur - polymeric solid
- Rhombic sulfur - large crystals composed of S_8 molecules
- Monoclinic sulfur - fine needle-like crystals

iii) Phosphorous

- Red Phosphorous (Polymeric solid)
- White Phosphoric (Crystalline solid)

SUMMARY OF THE CHAPTER

- ❖ The matter exists in **4 different states**, the **gaseous state**, **liquid state**, **solid state**, and the **plasma state**.
- ❖ Generally we deal with the first **3 states of matter**, the plasma is not considered as the true state by some scientists, as this state is not stable in earth on normal conditions.
- ❖ Gases are the simplest state of matter and are defined as the state of matter having neither definite shape nor definite volume.
- ❖ Gases have highest mobility, and kinetic energy is also highest.
- ❖ Gases diffuse easily and can be compressed by pressure.
- ❖ In order to explain the relation between the volume, pressure and temperature, various gas laws have been presented.
- ❖ According to Boyle's law, **the volume of a gas is inversely proportional to the pressure at constant temperature**.
- ❖ The Charle's law states that **volume of a gas is directly proportional to the temperature at constant pressure**.
- ❖ Liquids are the state of matter having definite volume but indefinite shape.
- ❖ Liquids also have the ability of diffusion, but with lesser extent to the gases.
- ❖ Solids are the state of matter having definite volume and definite shape.
- ❖ In solids the particles are held strongly and show only vibratory movement, so diffusion in solids is not possible.
- ❖ The solid can be divided into two categories, **the Amorphous solids** and **the Crystalline solids**.
- ❖ Different types of an element are called **Allotropic forms**. Allotropic forms have same composition but different structures.

EXERCISE

Q1. Fill in the blanks with appropriate words.

- i) Matter in the universe exists in states.
- ii) 1 atmosphere pressure is equal to mmHg.
- iii) The volume of a gas is related with temperature.
- iv) The liquids have volume and variable
- v) The relation between the density and the temperature is
- vi) Mixing of gas molecules is called
- vii) Boyle's law states that volume of a gas is related to the pressure.
- viii) Solids state of matter has density, as compared to liquid and gaseous states.
- ix) Same element with different shapes are called
- x) In Kelvin scale of temperature 300°C is equal to

Q2. Choose the correct answer.

- i) The form of matter which does not have definite shape and volume is called:
(a) liquid (b) gas (c) solid (d) plasma
- ii) Gases can be liquefied by:
(a) simple physical methods (b) simple chemical methods
(c) by using both physical and chemical methods
(d) no physical and chemical method can be used
- iii) Manometer is used for the determination of:
(a) pressure (b) volume (c) length (d) area
- iv) The vapour pressure can be determined by using:
(a) manometer (b) barometer
(c) both manometer and barometer
(d) neither manometer nor barometer
- v) The different forms of an element are called:
(a) isotopes (b) free radicals
(c) molecules (d) allotropes

Q3. Answer the following questions in short.

- i) If you take equal volumes of N_2 and CO_2 in different containers, which sample will expect to have more density?
- ii) Compare different states of matter.
- iii) Why we can't see most of the gases? Name 3 gases which are visible for us.
- iv) Compare diffusion and effusion processes in gases.
- v) Explain the term pressure of gases, also describe its units?
- vi) Why we use Kelvin scale in solving gas laws problem?
- vii) Can you explain why wet skin under the fan feels cool?
- viii) Differentiate between rates of diffusion among gases and liquids with examples?
- ix) Why solids have higher melting points than liquids?
- x) According to their shape what types of solids are found in nature?

Q4. Answer the following questions with reasoning.

- i) Why pressure of a gas sealed in a container when temperature is increased?
- ii) Explain in light of KMT what would happen with respect pressure when more quantity of a gas is added to container of gas?
- iii) Arrange the gases Cl_2 , H_2 , NH_3 in order of increasing speed of their molecules at $25^\circ C$.
- iv) What would you expect to happen in respect of volume when a balloon filled with air is placed in the freezer of a refrigerator?
- v) Name 3 methods by which the volume of a gas can be decreased.

Q5. How can you differentiate between the various forms of matter?

Q6. Derive a relation to show that volume and pressure have indirect relation?

Q7. Discuss the importance of Charles's law?

Q8. Define the term Allotropy? How the allotropes of oxygen effect our life?

Q9. What are the various types of solids? Give daily life examples to

differentiate?

- Q10. An ideal gas is kept in a cylinder of moveable piston of mass 1kg, has a volume of 500cc at the temperature of 25°C . If the temperature rises to 40°C , what will be the new volume of the gas?
- Q11. Suppose a gas is kept in a cylinder of volume 3litres and a pressure of 2 atmospheres, if the gas is shifted into a new cylinder of volume 5 litres. What will be the new pressure if the temperature remains constant?
- Q12. Compare the physical states of matter with regard to intermolecular forces present between them.
- Q13. Differentiate between the amorphous and crystalline solids with examples?
- Q14. Compare the different physical states of matter with regards to their typical characteristics.

SOLUTIONS

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the,

- Type of mixtures
- Matter and its states
- Moles and its determination

TIME ALLOCATION

Teaching periods	= 16
Assessment periods	= 02
Weightage	= 13%

LEARNING OUTCOMES

After reading this chapter, the Student will be able to:

- Define the terms solution, solute and solvent with examples. (Remembering)
- Explain the differences between saturated, unsaturated and super saturated solutions. (Analyzing)
- Explain the formation of solutions (mixing gases into gases, gases into liquids, gases into solids) and give an example of each. (Understanding)
- Explain the formation of solutions (mixing liquids into gases, liquids into liquids, liquids into solids) and give an example of each. (Understanding)
- Explain the formation of solutions (mixing solids into gases, solids into liquids, solids into solids) and give an example of each. (Understanding)
- Explain what is meant by the concentration of a solution. (Understanding)
- Define Molarity. (Remembering)
- Define percentage solution. (Remembering)
- Solve problems involving the Molarity of a solution. (Applying)
- Describe how to prepare a solution of a given Molarity. (Applying)
- Describe how to prepare dilute solutions from concentrated solutions of known Molarity. (Applying)
- Convert between the Molarity of a solution and its concentration in g/dm^3 . (Applying)
- Use the rule that "like is dissolved by like" to predict the solubility of one substance in another. (Understanding)

6.1 Solutions

In our daily life we often deal with a variety of mixtures where the components are mixed for different uses, such combinations are called solutions.

In this chapter we will study such solution mixtures in a little detail.

A solution is defined as the homogenous mixture of two or more substances.

A solution consists of two parts, the solute, and the solvent. The solute is that part which is lesser in quantity, while the solvent is that which is larger in quantity. For example, in the aqueous solution of sugar syrup in water, the sugar is the solute and the water is solvent, similarly when 10g of wax is dissolved in 100ml of benzene, the wax is solute and the benzene is solvent.

Generally in our daily routine we deal with many solutions, mostly prepared in the water solvent, such water solutions are called aqueous solutions. In a solution the components show their own properties, this makes the solution different from compound in which the components lost their properties.

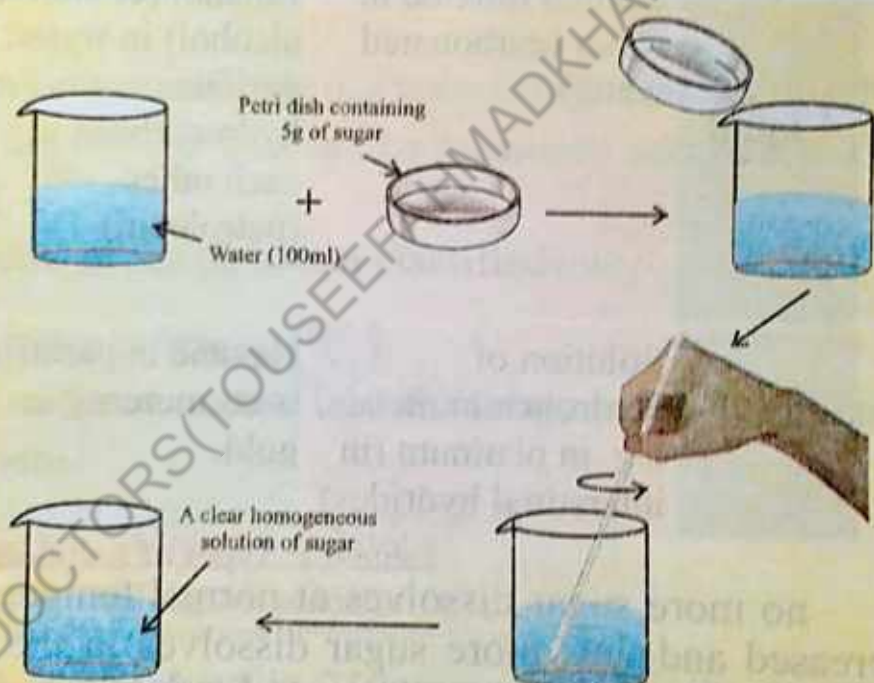


Figure: 6.1: The Process of formation of sugar solution

6.2 Saturated, Unsaturated and Supersaturated Solutions

A solution could be unsaturated, saturated or supersaturated. The unsaturated solution, as the name indicates is that solution which can dissolve further solute on addition at room temperature, e.g. when you put a spoon of sugar in a glass of water and still there is possibility of dissolving more sugar in normal conditions, this solution is the unsaturated solution of sugar. The

saturated solution is that in which no further addition of the solute is possible at normal conditions of temperature, pressure, e.g. in the above example of sugar solution when you mix more sugar and a stage come when no more sugar dissolves further, this is the saturated solution. While the supersaturated solution is the solution in which more of the solute have been dissolved by increase in temperature (or pressure sometimes), e.g. in above

Solvent	Solution		
	Gas	Liquid	Solid
Gas	Oxygen and other gases in nitrogen (air)	Water vapor in air (humidity)	Smoke
Liquid	Carbon dioxide in water (carbonated water)	Ethanol (common alcohol) in water; various hydrocarbons in each other (petroleum)	Sucrose (table sugar) in water; sodium chloride (table salt) in water; gold in mercury, forming an amalgam
Solid	Solution of Hydrogen in metals, e.g. in platinum (in interstitial hydrides)	Hexane in paraffin wax, mercury in gold.	Steel, duralumin, other metal alloys

Table 6.1: Types of Solutions

no more sugar dissolves at normal temperature then temperature is increased and thus more sugar dissolves in the given water, this heated mixture of sugar with water is termed as the supersaturated.

6.3 Types of Solutions

In a solution, the parts may be in the same phase or in different phases. The table 6.1 shows the 9 different types of phases present in a solution, with examples.

6.4 Concentration Units

Concentration of a solution is the amount of the solute in the given quantity of

Practical Activity

Using the help of your teacher in the laboratory prepare unsaturated, saturated and super saturated solutions of table salt (NaCl) in 100ml of water and observe the difference between each.

solvent/solution. It gives us the exact strength of a solution. For representing the concentration of solution following units are used frequently,

6.4.1 Percentage (composition)(%)

The percentage means parts out of hundred. The idea of percentage is very useful in calculating and understanding the constitution of a solution, for example, a 20% solution of sucrose means that the solution contains 20:80 constitution of the solution where the solute is 20 parts and the solvent is 80 parts. (or simply as the 20g sucrose in a solution whose final volume is 100ml).

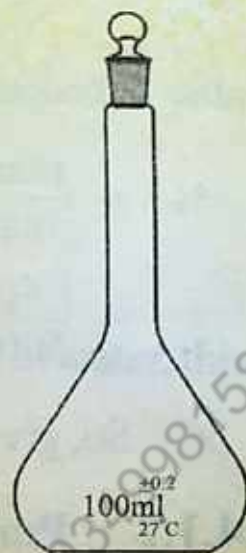


Figure: 6.2
Volumetric flask

For preparing a solution generally we use volumetric flasks, this has a long narrow neck so the volume in it can be measured accurately. This flask is shown in figure 6.2.

The concept of percentage can be further classified into:--

6.4.1.1 Percentage by mass/mass ($\frac{m}{m}$)

In this type of solution the amount of both the solute and the solution is expressed in the units of grams.

$$\% \text{ by } \frac{\text{mass}}{\text{mass}} = \frac{\text{Mass of solute(g)}}{\text{Mass of solution(g)}} \times 100$$

e.g., if 10 grams of NaCl is dissolved in 250grams of H_2O , it represents $\frac{\text{mass}}{\text{mass}}$ relation between the solute and the solvent.

Example 6.1

Find out the % of a solution that has 4g of glucose dissolved in 500g of water:

Solution: Here 4g of glucose has been dissolved in 500mg of water, so $\frac{\text{mass}}{\text{mass}}$ relationship exists;

The % of solution can be determined by applying formula:

$$\begin{aligned}\% &= \frac{\text{mass of glucose}}{\text{mass of solution}} \times 100 \\ &= \frac{4\text{g}}{500\text{g}} \times 100 = 0.8\%\end{aligned}$$

So, given solution has 0.8%

6.4.1.2 Percentage by mass/volume ($\frac{m}{v}$)

This involves the measurement of solute in mass units (grams) and of solution in volume units (cm^3 or ml).

$$\% \text{ by } \frac{\text{mass}}{\text{volume}} = \frac{\text{Mass of solute(g)}}{\text{Volume of solution}(\text{cm}^3)} \times 100$$

For example, if 10g of Glucose is dissolved in 1 litre of water, this relation is $\frac{\text{mass}}{\text{volume}}$.

Example 6.2

Calculate percentage of NaCl in a solution which has been prepared by dissolving 8g of NaCl in 250ml of water.

Solution: In this example the mass/volume relation exists and the percentage of this solution can be determined by applying the formula:

$$\% = \frac{\text{mass of solute}}{\text{volume of solution}} \times 100$$

by putting up the values in the above formula and solving the problem we will get the required result.

$$\begin{aligned}\% &= \frac{\text{mass of solute}}{\text{volume of solution}} \times 100 \\ &= \frac{8\text{ g}}{250\text{ml}} \times 100 = 3.2\%\end{aligned}$$

So, given solution has 3.2% of NaCl

6.4.1.3 Percentage by volume/mass ($\frac{v}{m}$)

In such type of solutions the amount of solute is expressed in volume unit and of solution in mass units.

$$\% \text{ by } \frac{\text{volume}}{\text{mass}} = \frac{\text{Volume of solute (cm}^3\text{)}}{\text{Mass of solution (g)}} \times 100$$

For example, if 10ml of alcohol is dissolved in 100gms of water, this is the $\frac{\text{volume}}{\text{mass}}$ relation.

Example 6.3

In a face cream recipe 6ml of almond oil has been mixed in petroleum jelly to get 50g of total mixture solution. Find out the percentage of almond oil in this cream.

Solution: In this example the solute has been measured in volume unit and solution is in a grams. So the relation here is volume/mass.

the solved this example we will use the following formula:

$$\text{Percentage} = \frac{\text{Volume of solute}}{\text{Mass of solution}} \times 100$$

$$\text{Percentage} = \frac{6\text{ml}}{50\text{g}} \times 100 = 12\%$$

So, given solution mixture has 12% of almond oil.

6.4.1.4 Percentage by volume/volume ($\frac{v}{v}$)

In such type of solutions the amount of both the solute and the solution are expressed in volume units.

$$\% \text{ by } \frac{\text{volume}}{\text{volume}} = \frac{\text{Volume of solute (cm}^3\text{)}}{\text{Volume of solution (cm}^3\text{)}} \times 100$$

Activity 6.1

For the preparation of vinegar 30ml of acetic acid has been dissolved in water to make the total volume of mixture one litre. Calculate the percentage of CH_3COOH (acetic acid) in this vinegar mixture.

For example, this relation is observable in case when 10ml of alcohol is dissolved in 100ml of water.

6.4.2 Molarity (M)

This is the unit which is used frequently for well description of a solution's components. Molarity is the number of moles of a solute dissolved per litre of a solution, e.g. 0.1M NaCl solution means that 0.1 moles of NaCl (5.85g) dissolved in 1 litre of aqueous solution.

$$\text{Molarity (M)} = \frac{\text{number of moles of solute}}{\text{Volume of solution in Litre}}$$

6.4.3 Problems involving the Molarity of a Solution

In order to calculate the molarity of a solution or the amount of solute in molar quantity above formula is used:

$$\text{Molarity} = \frac{\text{Moles}}{\text{Amount of solution in litre}}$$

Or more simplified form of this formula is:

$$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molecular mass of solute} \times \text{Amount of solution in litre}}$$

This is further explained by the help of following example:

Example: 6.4: Calculate amount of NaOH required to prepare its 0.1M solution in 500ml of water.

Solution: As we know that the molarity calculations involve the use of formula: Mass of solute

$$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molecular mass of solute} \times \text{Amount of solution in litre}}$$

Or

$$\text{Mass of Solute} = \text{Molarity} \times \text{Molecular mass of solute} \times \text{Amount of solution in litre}$$

By inserting values in this formula

$$\text{Mass of Solute} = 0.1 \times 40 \times 0.5 \text{ (lit)}$$

$$\text{Mass of Solute} = 2 \text{ gms}$$

So, if 2 grams of NaOH is dissolved in 500ml of water it will give 0.1M solution of NaOH.

Activity: 6.2

How will you prepare a 250 ml of a solution of 0.5M NaHCO_3 ?

6.5 Dilution of Solutions

Although we often make solution for any purpose, but many times we are provided with some prepared solutions and we just have to fix its molarity or percentage as per our requirements, for this purpose we have to dilute the solutions up to the desired level, this process is called dilution of solution.

The dilution of a solution is brought about by using a simple relationship that is present among the required and the given solutions, this relation is given by the expression,

$$M_1 V_1 = M_2 V_2$$

Where M_1 and V_1 are molarity and volume of dilute solutions while M_2 and V_2 are molarity and volume of concentrated solutions.

V = Volume

1 and 2 represents the provided and required solutions.

Example 6.5: What is the volume of HCl required to prepare 500ml of 0.1M solution from the commercially provided 36M HCl acid?

Solution: This process requires the dilution process. The formula that will work is:

$$M_1 V_1 = M_2 V_2$$

Here M_1 = Molarity of commercial HCl = 36M

V_1 = Volume of commercial HCl = ?

M_2 = Molarity of required HCl = 0.1M

V_2 = Volume of required HCl = 500 ml

So, by inserting values in the given formula, we get

$$M_1 V_1 = M_2 V_2$$

$$36 \times V_1 = 0.1 \times 500$$

$$V_1 = 1.39 \text{ ml}$$

Activity: 6.3

Prepare a solution of 0.5M H_2SO_4 from the commercially provided 12M solution in 400 ml quantity.

Activity: 6.4

Prepare a solution of 200ml of 2% glucose from 10%.

It means that when we take 1.39ml of commercially provided HCl, and then make its volume upto 500ml, it will be solution of desired strength and quantity, i.e. 0.1M in 500ml.

Same relation is used for the preparation of solutions from the percentage values.

6.6 Solubility

Solubility is defined as the amount of a substance which is able to dissolve in the given quantity of a solvent at a particular temperature.

Normally in the solubility, the amounts of both the solute and the solvent are taken in the units of grams, thus it can also be defined as the number of grams of a solute which are soluble in 100gm of solvent.

Solubility is related with many factors, of which the nature of solute and solvent are very important, a general phenomenon which is observed during solubility is that like is dissolved by the like, this means that polar solute are soluble in polar solvent and non polar solutes are soluble in non polar solvents, thus NaCl is soluble in water (both are polar) and petroleum jelly is dissolvable in benzene (both are non polar).

Further more temperature also plays an important role in determining solubility; this will be discussed in detail later in this chapter.

6.6.1 Solute-solvent interaction during Solubility

As described earlier that a general phenomenon works in formation of solutions, i.e. like is dissolved by like. This is due to development of attractive forces among the solute and the solvent particles, and the resultant product of the process has lower stabilizing energy, so the solution is formed.

The process of solubility involves ionization of solute (in the polar solvents), in a manner that each cation is captured by negative poles of solvent and anions by positive poles of solvent. The process is known as solvation or solvolysis. This process of solvolysis between NaCl and H_2O is shown in figure 6.3.

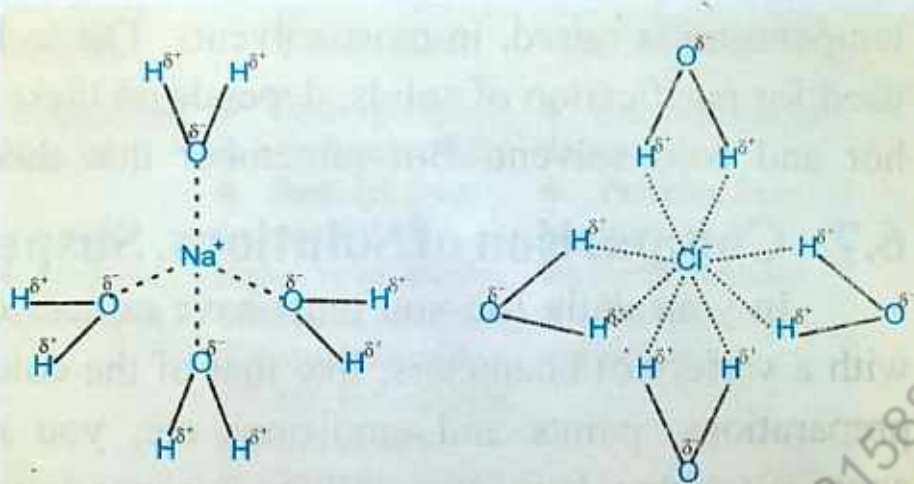


Figure 6.3: Solute – solvent interaction forces in sodium chloride solvolysis

6.6.2 Solubility and the Temperature

The solubility of a given solute in a given solvent generally depends upon the temperature. For around 95% of solid solutes, the solubility increases with temperature, but the gaseous solutes, the temperature increase usually decreases solubility in water, and solubility increases in the organic solvents mostly.

The given graph in figure 6.4 shows solubility curves for some typical inorganic salts (all solids).

It can be seen here that many salts like sodium nitrate and potassium nitrate show a large increase in solubility with temperature. Some solutes (e.g. NaCl in water) are fairly independent of temperature. A few, such as HCl, NH_3 , become less soluble in hot water. Organic compounds nearly always become soluble as the

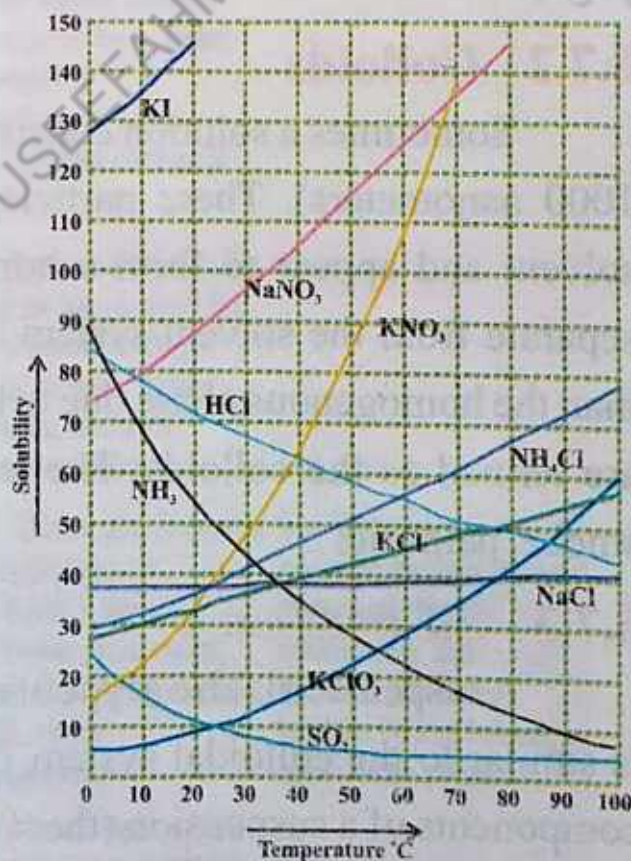


Figure 6.4: Relation between Solubility and Temperature

temperature is raised, in most solvents. The technique of recrystallization, used for purification of solids, depends on these differences in solubility in hot and cold solvent. But remember that there are always exceptions.

6.7 Comparison of Solutions, Suspension and Colloids

In your daily life you must have noticed different types of solutions with a variety of characters, like that of the cold drinks, sugar syrup, drug preparations, paints and emulsions etc, you must have observed their composition too, here in the following pages now we will make you aware of the interesting features and differences among these in the scientific view,

6.7.1 Solution

Solution or a typical completely miscible form of solution is actually the best regarded true form of solution that we have described in the previous pages.

6.7.2 Colloids

Sometimes a solution consists of bigger solute particles, (between 2 to 1000 nanometers). These particles although are evenly distributed in the solvent and appear to form a homogenous mixture, but over a time they separate from the solvent system and form a heterogeneous mixture rather than the homogenous phase due to having a bigger solute size. Such solutions are termed as the colloids. The examples of such colloids are blood, fog, smoke, paint etc.

6.7.3 Suspension

A suspension is also a type of solution which is in heterogeneous phase. It is similar to the colloidal system in many ways, but unlike the colloids the components of a suspension (the solute) seems always separated through the solvent system, whereas in the colloids, the solute gets separated over a time period. The particle size of suspension is greater than 1000nm.

The solution of flour in water (sattu) and mud in water are the common examples of suspensions. Many of the medicinal solutions are marketed as a

	SOLUTIONS	COLLOIDS	SUSPENSIONS																														
COMPOSITION	◆ solute particles have dissolved to the point of ions, atoms, or molecules ◆ solute particles are evenly dispersed throughout the solvent - Particles do not settle down with time.	◆ Solute particles do not dissolve fully ◆ Particles form groups of ions, atoms, or molecules. ◆ Particles are evenly dispersed through solvent, but solution appears cloudy ◆ Particles do not settle out with time	◆ Solute particles do not dissolve ◆ Particles form large groups of insoluble particles. ◆ Particles settle out with time																														
SIZE	◆ Particles are too small to see with naked eye (less than 1 nanometer = 0.000000001 m)	◆ Particles are usually not seen by naked eye (1 - 1000 nanometers)	◆ Particles can be seen with naked eye (larger than 1000 nm)																														
SEPARATION	◆ Solute particles will pass through a paper filter and a semi permeable membrane; cannot be separated except through distillation	◆ Solute particles will pass through a paper filter; can be separated by a semi permeable membrane, i.e., cellophane and cell walls	◆ can be easily separated by filtering																														
OTHER	◆ May be dilute or concentrated; May be unsaturated, saturated, or supersaturated. - Degree of saturation is dependent on temperature; ex, more sugar will dissolve when tea is hot)	◆ Parts of a colloid may be separated by an ultracentrifuge (spin at very high speeds; used to separate blood)																															
TYPES	◆ There are 9 types of solutions, based on the solute (1st) and the solvent (these are also discussed in table 6.1) <table><tr><th>solute</th><th>solvent</th><th>example</th></tr><tr><td>gas</td><td>gas</td><td>air</td></tr><tr><td>gas</td><td>liquid</td><td>soda water</td></tr><tr><td>gas</td><td>solid</td><td>hydrogen in Pt</td></tr><tr><td>liquid</td><td>gas</td><td>water vapor in air</td></tr><tr><td>liquid</td><td>liquid</td><td>alcohol in water</td></tr><tr><td>liquid</td><td>liquid</td><td>silver amalgam</td></tr><tr><td>solid</td><td>gas</td><td>sulfur vapor in air</td></tr><tr><td>solid</td><td>liquid</td><td>sugar in water</td></tr><tr><td>solid</td><td>solid</td><td>brass</td></tr></table>	solute	solvent	example	gas	gas	air	gas	liquid	soda water	gas	solid	hydrogen in Pt	liquid	gas	water vapor in air	liquid	liquid	alcohol in water	liquid	liquid	silver amalgam	solid	gas	sulfur vapor in air	solid	liquid	sugar in water	solid	solid	brass	◆ There are 3 types of colloids: (a) Gels - liquid particles in a solid, flow slowly (gelatin, jelly, stick deodorant) (b) Emulsions - two liquids (mayonnaise, hand cream, milk) (c) Aerosols - solid or liquid in a gas (fog, smoke, paint-spray can)	◆ Examples include Italian salad dressing, liquid medicines that require shaking before being taken; and some paints.
solute	solvent	example																															
gas	gas	air																															
gas	liquid	soda water																															
gas	solid	hydrogen in Pt																															
liquid	gas	water vapor in air																															
liquid	liquid	alcohol in water																															
liquid	liquid	silver amalgam																															
solid	gas	sulfur vapor in air																															
solid	liquid	sugar in water																															
solid	solid	brass																															

suspension, which are also labeled as "suspension" and instructed to be well shaken before use. Ice cream is also a suspension of ice crystals in cream.

The following table 6.2 brings the summary of differences between these three types.

SUMMARY OF THE CHAPTER

- ❖ **Solution** is a homogeneous mixture of two or more substances.
- ❖ Solution is not a pure substance, and its composition varies in different types.
- ❖ The component of solution which is lesser in quantity is called solute.
- ❖ The component of solution which is greater in quantity is called solvent.
- ❖ Solution may be **unsaturated**, **saturated** and **super saturated**, depending upon the amount of solute and temperature.
- ❖ If more solute can be added in a solution then the solution is called **unsaturated solution**.
- ❖ If no more solute can be dissolved further in making of a solution, then the solution is called **saturated solution**.
- ❖ When more solute is added to in a solution by increasing the temperature, then the solution is said to be **super saturated solution**.
- ❖ Solution can be formed in all states of matter, thus there are 9 different types of solutions which are formed.
- ❖ The strength of a solution is represented by various means, like in percentage, molarity etc.
- ❖ The percentage composition of a solution describes the parts of solute per hundred parts of solution.
- ❖ **Molarity** of a solution gives the number of moles of solute dissolved per liter of solvent; it is symbolized as "M".
- ❖ Generally **increase in temperature increases the solubility** of a solute,

but the solutions of gases do not follow this rule.

- ❖ The solution mixture in which size of particles ranges from 1-1000nm is called the **colloid** or **colloidal solution**.
- ❖ The solution mixture in which size of particles is greater than 1000nm is called the **suspension**. Some times the particles in suspension can even be seen with naked eye.

EXERCISE

Q1. Fill in the blanks with appropriate words.

- i) A solution is the type of mixture.
- ii) Strength of a solution in mole/litre is called
- iii) The solute-solvent interaction in dissolution process is termed as
- iv) Generally heating the solution formation.
- v) Milk of magnesia is the example of
- vi) The solute is that part of solution which is in quantity.
- vii) In solution no more quantity of solute can be dissolved at normal temperature.
- viii) In dissolution process, a general rule of solubility is that dissolves like.
- ix) ppm stands for
- x) The solubility of gases with increase in temperature.

Q2. Choose the correct answer.

- i) The solvent is the part of solution which is:
(a) more in quantity (b) less in quantity
(c) equal in quantity (d) does not depend on quantity
- ii) The heterogeneous mixtures where the particle size of solute is larger than 1200nm is called:
(a) solution (b) colloids

- (c) buffers (d) emulsion
- iii) In dissolving table salt, heating accelerates the solvolysis.
 (a) true (b) false
 (c) both statements are false (d) both statements are true
- iv) The air is a solution:
 (a) true (b) false
 (c) both statements are false (d) both statements are true
- v) Steel is the type of solution having:
 (a) liquid-solid composition (b) liquid-liquid composition
 (c) solid-solid composition (d) solid-gas composition

Q3. Answer the following questions in short.

- i) What is the difference between unsaturated, saturated and super saturated solutions?
- ii) Draw a picture to show that what happens when an ionic substance is added to water to dissolve?
- iii) Why NaCl is dissolved easily in cold water but not in heated cooking oil?
- iv) What types of solutions can form? Example of each type.
- v) What is meant by molarity?
- vi) Explain dilution of solutions with an example.
- vii) Explain the term suspension?
- viii) What are colloids?
- ix) Compare advantages and disadvantages of percentage and molarity over each other.
- x) What would happen if solubility of oxygen in water increased one thousand times?

Q4. Answer the following questions with reasoning.

- i) Which of the following expected to dissolve in H_2O and why?
 (a) Alcohol, (b) Hydrocarbon, (c) NaCl, (d) Iron
- ii) Explain why some liquids are miscible in one another which

other liquids are immiscible.

iii) Why it is necessary to bubble air through an aquarium?

iv) Why solubility of gases decreases with an increase in temperature?

v) Explain how can you obtain pure solute from a solution by adding more solute to the solution.

Q5. Define the term solution? How the solutions form.

Q6. Discuss various concentration units?

Q7. Calculate the molarity of a solution of NaHCO_3 that contains 50gm of NaHCO_3 in 250ml of water.

Q8. Find out the amount of glucose required to prepare its 0.5M solution in 500ml of water.

Q9. How the temperature effects the solubility process?, give examples to prove your statements.

Q10. How would you differentiate between the solution, suspension and colloids?

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the:

- Charge concept.
- Basics of Chemical reactions
- Basic structure of periodic table

TIME ALLOCATION

Teaching periods	= 16
Assessment periods	= 03
Weightage	= 14%

**LEARNING OUTCOMES**

After reading this chapter, the Student will be able to:

- Define Oxidation and Reduction in terms of loss or gain of oxygen and hydrogen. (Understanding)
- Define oxidation and reduction in terms of loss or gain of electrons. (Understanding)
- Identify the oxidizing and reducing agents in redox reactions. (Analyzing)
- Identify the oxidizing and reducing agents in redox reactions. (Understanding)
- Define oxidation state. (Remembering)
- State the rules for assigning the oxidation numbers. (Remembering)
- Determine the oxidation numbers of element of any atom in a compound. (Applying)
- Describe the nature of electrochemical processes. (Understanding)
- Sketch an electrolytic cell, with labels. (Understanding)
- Identify the direction of movement of cations and anions towards respective electrodes. (Understanding)
- List the possible uses of an electrolytic cell. (Understanding)
- Sketch a daniell cell, with labeling. (Understanding)
- Describe how battery produces electrical energy. (Understanding)

7.0 Introduction to Electrochemistry

Electrochemistry is a branch of chemistry in which the conversion of chemical energy into electrical energy and electric change produced by an electric current are studied. The electrochemistry also deals with situations where oxidation and reduction reactions takes place, but the direct charge transfer in the form of electrons from one species to another is not the topic of electrochemistry.

Electrochemistry is an important branch of chemistry and many important topics like metallurgy of metals, analysis of chemical compounds, battery operations, electroplating etc are discussed here.

7.1 Oxidation and Reduction

Oxidation and Reduction may be defined in many ways, these definitions can be summarized in the following manner,

Oxidation:

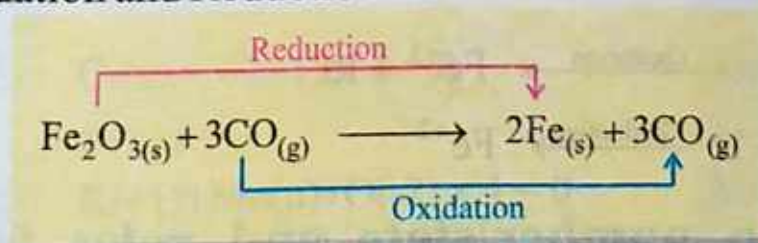
- Oxidation is gain of oxygen
- Oxidation is loss of hydrogen
- Oxidation is loss of electrons
- Oxidation involve an increase in oxidation state

Reduction may be defined as:

- Reduction is loss of oxygen
- Reduction is gain of hydrogen
- Reduction is gain of electron
- Reduction involve a decrease in oxidation state

These point are further explained by the help of following examples.

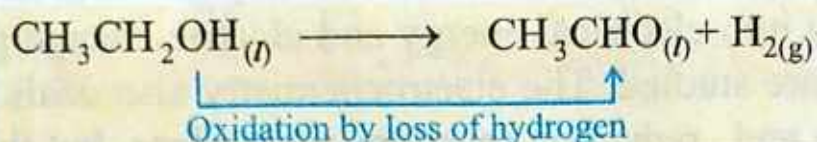
In the extraction of iron (Fe) from its ore Iron oxide (Fe_2O_3), the processes of Oxidation and Reduction are:



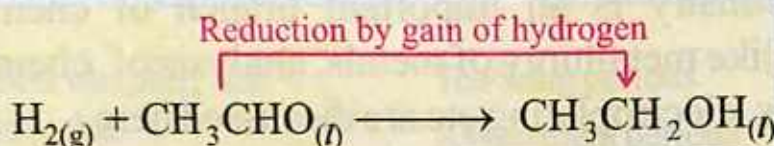
As shown here, ' Fe_2O_3 ' is reduced to Fe, and ' CO ' is oxidized to ' CO_2 ', because here it can be seen that from reducing species ' Fe_2O_3 ' the oxygen is being eliminated, while from the oxidizing species i.e. the ' CO ' the oxygen is being added.

In the following case of ethanol,

The Ethanol (or Ethyl alcohol) is oxidized by loss of hydrogen.

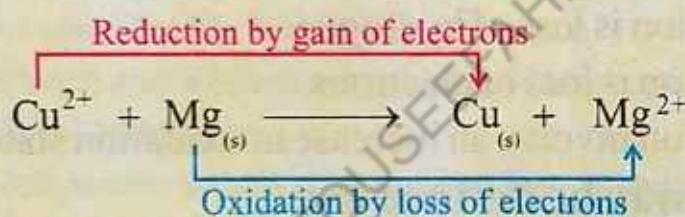


And in this case of Ethanal (also called Acetaldehyde), the reduction is brought about by the gain of hydrogens.



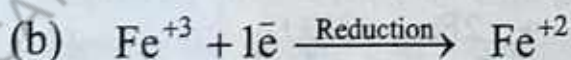
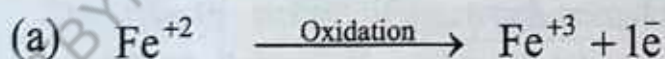
The modern view of explaining the Oxidation and Reduction processes are in terms of electron transfer,

As described above the oxidation is loss of electrons, while Reduction is the gain of electrons.



In this above example, the Copper (Cu) is being reduced to Cu from Cu^{2+} , by gaining electrons, while Magnesium is oxidized to ' Mg^{2+} ' by the loss of electrons.

Similarly in the following case (a), the Fe^{+2} is oxidized to Fe^{+3} by the loss of one electron, while in case (b) the reaction is reversed and the process is the reduction.



7.2 Oxidation number/state and rules for assigning oxidation numbers.

The charge which an atom bears in a compound, molecules or ions is called its oxidation number; the oxidation number may be positive or negative

depending upon the electronegativity of an atom.

Following rules are applied to calculate the oxidation number.

- The oxidation state of a free element is always zero. For example, ' O_2 ', ' Cl_2 ', ' S_8 ', and ' Si ' have zero oxidation state.
- The sum of the oxidation states of all the atoms or ions in a neutral compound is zero.
- The sum of the oxidation states of all the atoms in an ion is equal to the charge on the ion.
- The more electronegative element in a substance is given a negative oxidation state. The less electronegative one is given a positive oxidation state.

In the following table 7.1 common oxidation numbers of some elements are written.

Group	Elements	Oxidation State
1	Na, K etc.	'+1'
2	Mg, Ca, etc.	'+2'
17	F, Cl, etc.	Generally '-1'
16	O	Generally '-2'
1	H	+1 except in metallic hydride where its oxidation number is '-1'.

Table 7.1 Common oxidation states of some Groups in periodic table.

The oxidation state of some elements are calculated in neutral compounds are given below in the table 7.2:

Substance	Oxidation Number Calculation	Oxidation Number of the Elements
$CrCl_3$	$Cr(n) + Cl(-1) \times 3 = 0$ $n = +3$	$Cr = +3$
$KMnO_4$	$K(+1) + Mn(n) + O(-2) \times 4 = 0$ $n = +7$	$Mn = +7$
$K_2Cr_2O_7$	$K(+1) \times 2 + Cr(n) \times 2 + O(-2) \times 7 = 0$ $n = +6$	$Cr = +6$

Table 7.2: Oxidation numbers of selected elements in some compounds.

Calculate the oxidation number of 'N' in following:

- (i) HNO_3 (ii) NO_2 (iii) NOCl

7.3 Oxidizing and Reducing Agents

Substances that bring oxidation process are known as oxidizing agents. In a broad sense, the oxidation, which is defined as the loss of electrons, so any substance which brings loss of electrons or accepts electrons during the chemical reactions is termed as the oxidizing agents, oxidants or oxidizers. The oxidant removes electrons from another substance, and is thus reduced itself. And because it "accepts" electrons it is also called an electron acceptor. Oxidants are usually chemical substances with elements in high oxidation numbers (e.g., H_2O_2 , H_2SO_4 , $\text{K}_2\text{Cr}_2\text{O}_7$, MnO_4^- , CrO_3 , $\text{Cr}_2\text{O}_7^{2-}$, OsO_4) or highly electronegative substances that can gain one or two extra electrons by oxidizing a substance (O, F, Cl, Br).

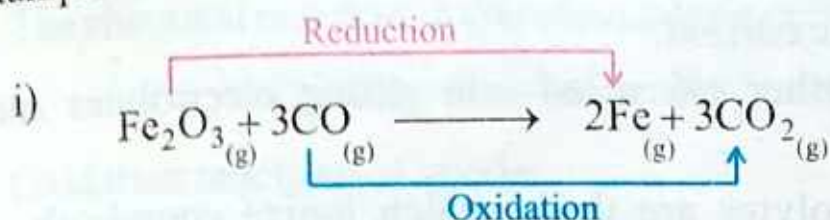
The reducing agents are just opposite to the oxidizing agents. The substances that have the ability to reduce other substances are said to be reductive and are known as reducing agents, reductants, or reducers. The reductant transfers electrons to another substance, and is thus oxidized itself. It "donates" electrons it is also called an electron donor. Metals (Li, Na, Mg, Fe, Zn, Al) can be used as reductants. These metals donate or give away electrons readily. The hydrogen gas (H_2) with a palladium, platinum, or nickel catalyst can also be used as the reducing agent. These catalytic reductions are primarily used in the reduction of carbon-carbon double or triple bonds.

Redox process is that the reductant transfers electrons to the oxidant. Thus, in the reaction, the reductant or reducing agent loses electrons and is oxidized and the oxidant or oxidizing agent gains electrons and is itself reduced.

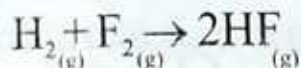
7.4 The Oxidation-reduction Reactions (Redox Reaction)

As described earlier, the oxidation and reduction go side by side in a reaction. A substance which oxidizes the other is itself reduced and vice versa.

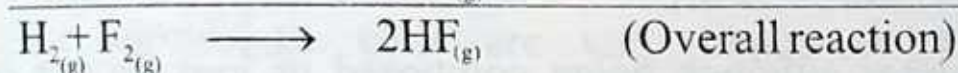
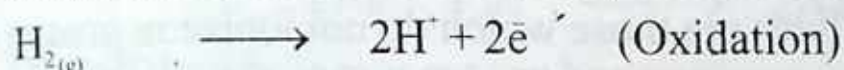
For example:



ii) Another example is the reaction between the H_2 and F_2



Here, The H_2 loses two electrons, which are accepted by the F_2 ,



As seen, the overall reaction results in the formation of hydrogen fluoride (HF).

The redox reactions are very important to understand, both in research and industrial applications, as most of reactions involve the transfer of electrons from one species to the other, and are thus categorized as the redox reactions. There understanding hence is used to increase the desired products by suitable procedures.

7.5 The Electrochemical Cell

An electrochemical cell is a device containing at least two electrodes in the solution of an electrolyte, and here in this cell the oxidation-reduction (Redox) reactions takes place.

The electrochemical cells are of two types, i.e.

i) Electrolytic cell

ii) Galvanic cell

These types are described in detail in the following pages, but let's first understand the term electrolytes.

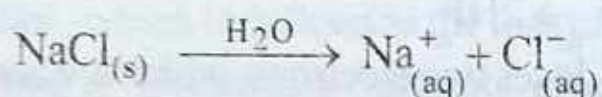
7.5.1 The Concept of Electrolytes

The electrolyte is a substance which is capable of conducting electricity in its solution or molten state. The electrolytes are generally

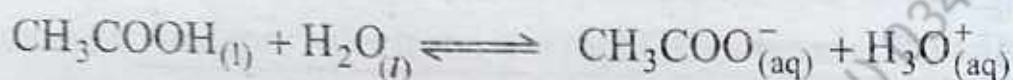
consists of ions which serve as the carrier of electrons and hence conduct easily the electric current.

They may be further classified into strong electrolytes and weak electrolytes.

The strong electrolytes are those which ionize completely in their aqueous solution, e.g. NaCl, HCl, and NaOH etc are some strong electrolytes.



The weak electrolytes are those which do not ionize in greater extent, e.g. acetic acid.



The electrolytes although being composed of ions do not conduct electricity in the solid state because in solid state the mobility of ions is very low.

7.5.2 Electrolytic Cells

An electrolytic cell is the cell that contains a solution of electrolytes and has been fitted with two electrodes, anode and the cathode which are attached with a source of direct current (DC). In this cell a non spontaneous reaction takes place using the electrical energy and some products are formed.

In the electrolytic cell when electricity is passed the ions present in the solution are attracted towards the respective electrode, i.e. anions towards anode and cations toward cathode. The electrode that attracts anions is called anode and the electrode that attract cations is called cathode.

In case of the NaCl, the Cl^{-} ions move towards the anode and loss their electrons there, hence oxidation reaction occurs here, while the Na^{+} ions moves toward cathode where reduction reaction occurs, and Na^{+} gains electron yielding neutral products.

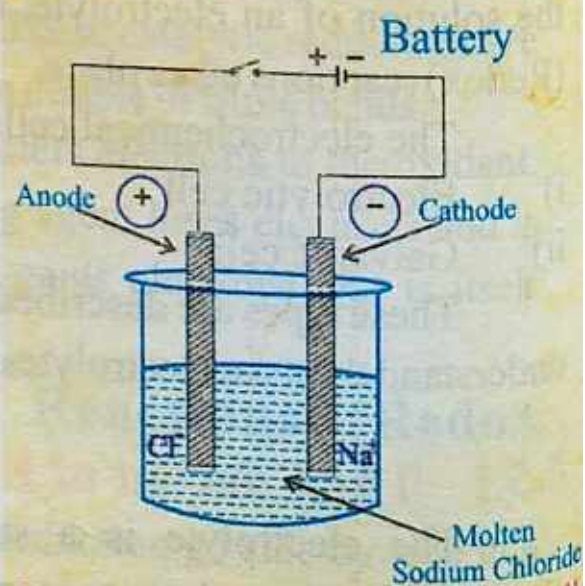


Figure 7.1: A typical electrolytic cell

The chemical reaction in this electrolytic cell are following:



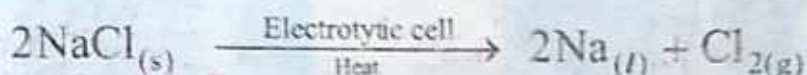
Oxidation reactions at anode:



Reduction reaction at cathode:



Overall the whole process can be written as:



The electrolytic cells are very important. They are used in electrorefining process, through which high grade metals and other substances are produced. These cells are also used in electroplating processes.

7.5.3 Galvanic Cells or the Daniell Cell

The Galvanic cell is also a type of electrochemical cell like the electrolytic cells but differs from that in the sense that it is used for the production of electric current or flow of electron current from the spontaneous redox reaction. This cell is also known as Daniell cell or Voltaic cell, named after Luigi Galvani.

This consists of two different metals connected by a salt bridge or a porous disk between the individual half-cells. A Galvanic cell is shown in figure 7.2.



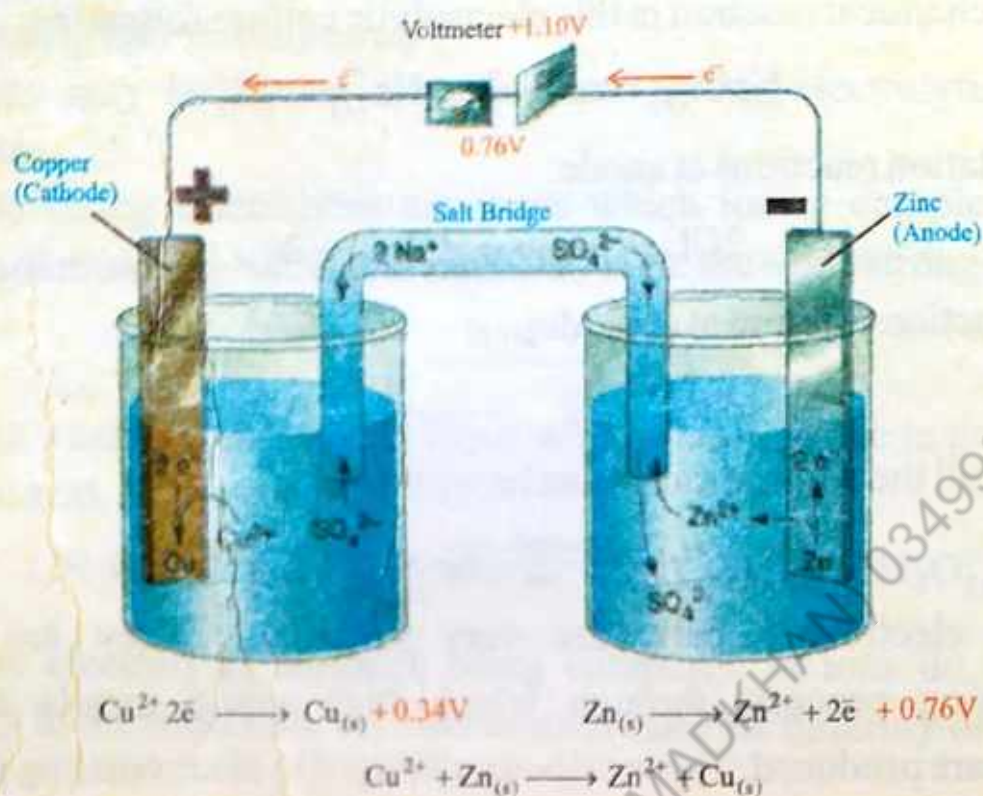
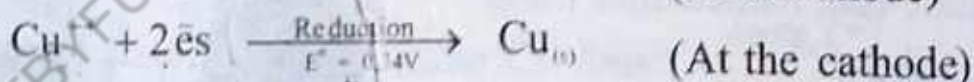
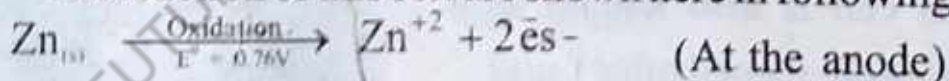


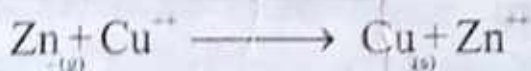
Figure 7.2: A typical Galvanic cell

In the Galvanic cell, the Zn metal has a high tendency to lose electrons than Cu, as a result oxidation reaction takes place at Zn electrode. The electrons flow from Zn electrode through the external wire in the circuit to Cu electrode. These electrons are gained by the Cu ions of the solution and copper atoms deposit at the electrode. As a result of this electric current is produced.

The reaction of this cell are shown here in following.



That the overall reaction could be:



The cell voltage can be calculated as:

$$E_{\text{cell}}^\circ = E_{\text{oxd}}^\circ + E_{\text{red}}^\circ$$

$$E_{\text{cell}}^\circ = 0.76 + 0.34 = 1.10\text{V}$$

The Galvanic cells are extensively used in the production of electricity through the chemical reaction and the lead—acid batteries are based on this cell.

7.6 Electrochemical Industries

The Electrochemistry occupies a great place in the life standards of modern world. Every institution, whether it is a home or a factory, utilizes the applications based upon the electrochemical instrumentations. For example, the batteries, whether they are the conventional Lead-acid batteries (without

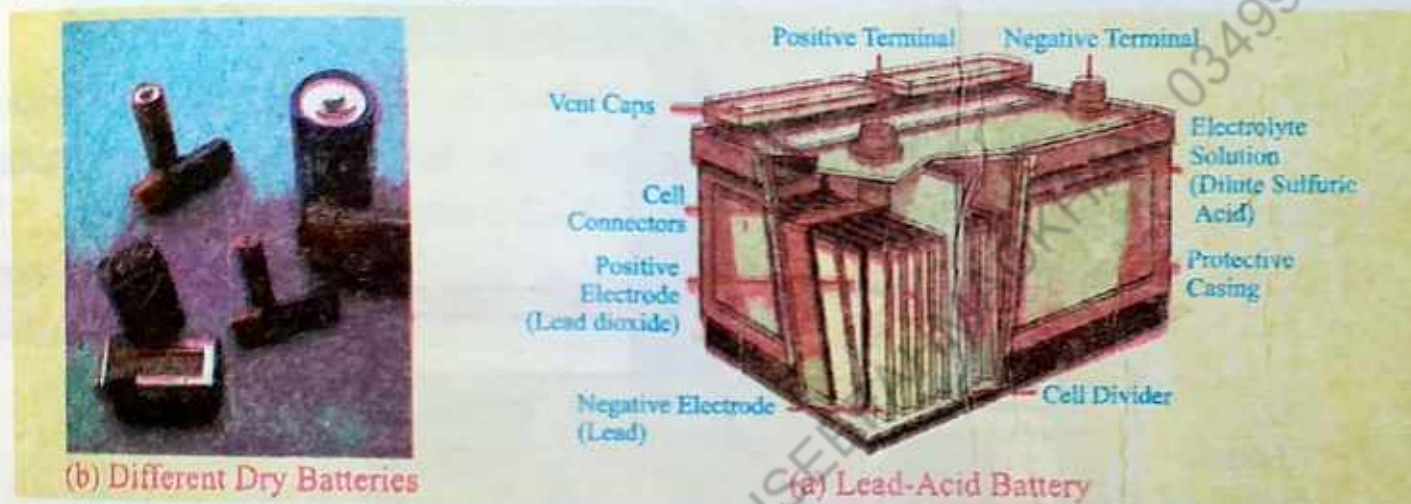


Figure: 7.3 (Different Types of Batteries)

which no vehicle can run) or dry cells batteries or the lithium batteries, (which provides the source of electricity for the most of electronic equipments such as the wrist watches, radios, and even the computers) or the most modern used the sealed batteries (to which you are familiar in your rechargeable emergency lights), all these are the direct result of the electrochemical applications.

The electroplating methods, which are used widely for the prevention of corrosion, and also for the fine looking of the objects (e.g. the gold and other electroplating) also use the electrochemical principles.

The process of electrolysis is also the important example of the electrochemical application. This



Figure 7.4
A Gold plated ornament

process is frequently used in the preparation of many important compounds and for the decomposition of many substances into simpler fragments.

7.6.1 Manufacture Of Sodium Metal from fused NaCl

The Sodium metal which holds an important place is obtained from its common salt, the NaCl. This process is commercially performed in a specialized electrolytic cell called Down's cell.

The Down's cell is shown in figure 7.5.

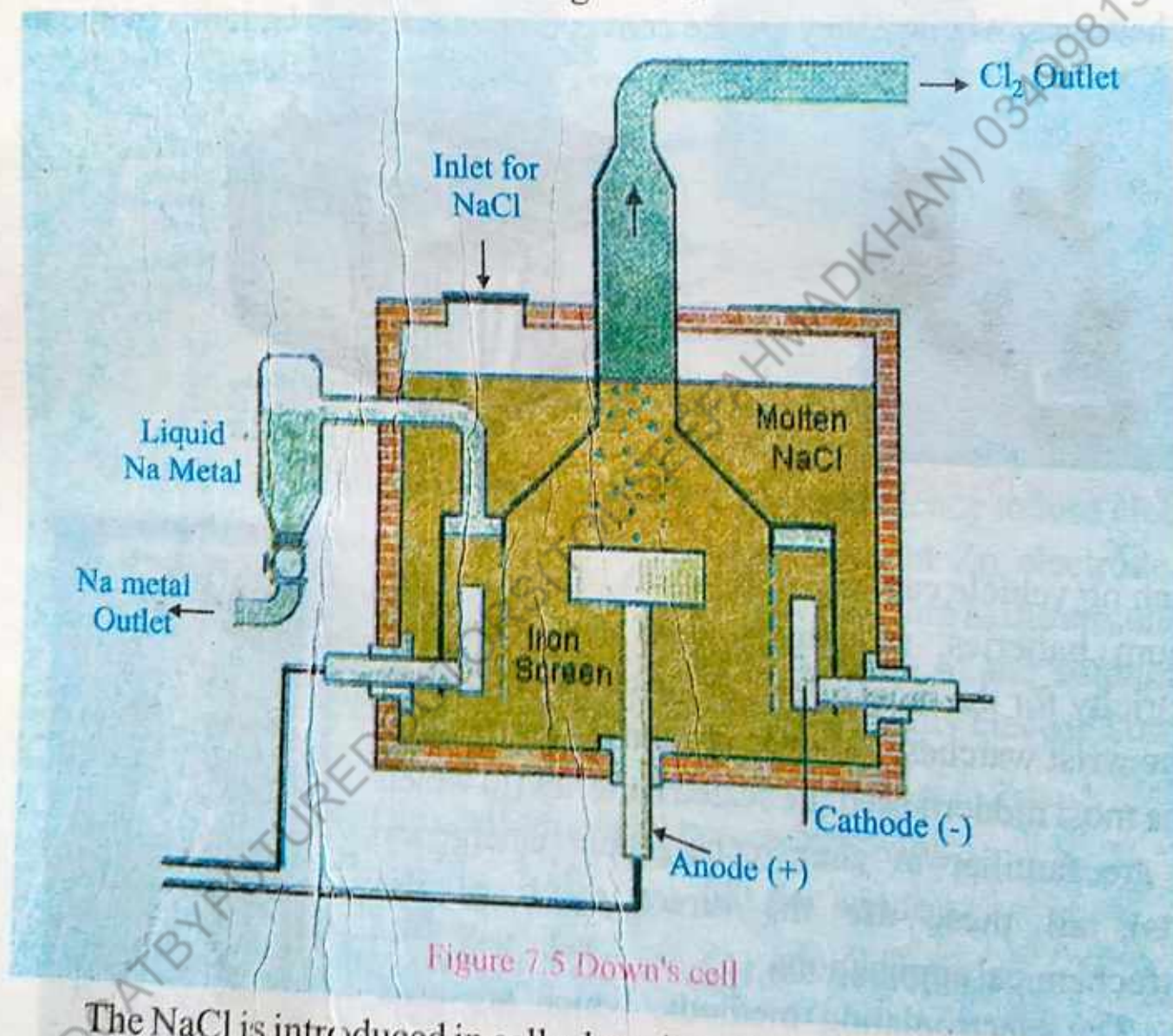
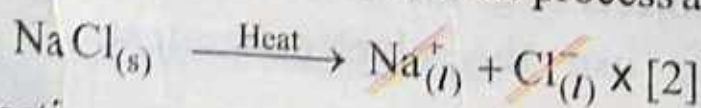


Figure 7.5 Down's cell

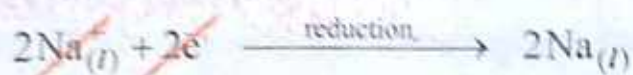
The NaCl is introduced in cell where it melted. The electricity is passed through the electrodes. Different reactions of the process are:



Oxidation reactions at anode.



Reduction reaction at cathode.



Overall the whole process can be written as:



As a result Cl_2 is given out from the anode and molten sodium is collected at the cathode.

7.6.2 Manufacture of NaOH from Brine

The NaOH or Caustic soda is a very important chemical and is used in countless industries like in leather industry, in fabric industry, in soap industry, paper industry and many more.

It is commonly known as Caustic soda and is industrially prepared from the electrolysis of brine (concentrated aqueous solution of NaCl) in Nelson's cell which is a specialized electrolytic cell.

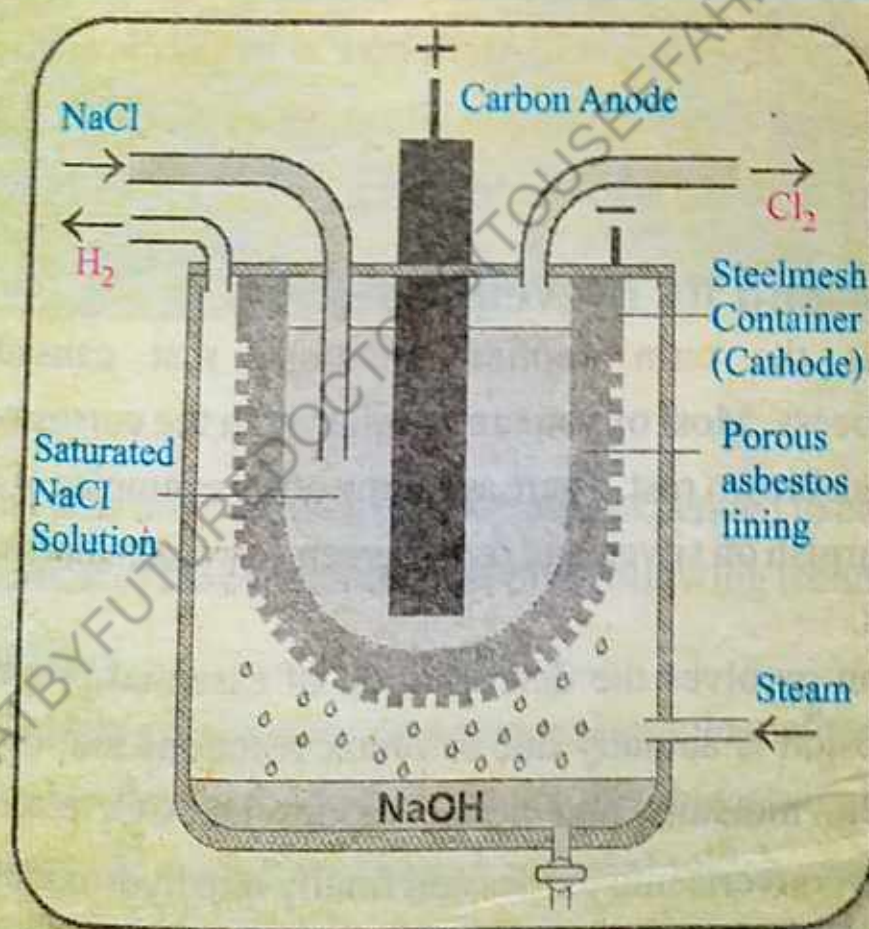


Figure 7.6 Nelson's cell

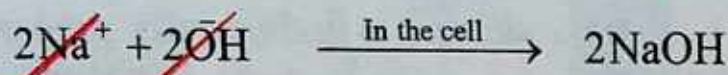
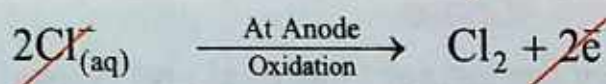
The process differs from the Down's method in the sense that in brine

some H_2O is also present which is reduced during the process and form OH^- ions. Na^+ and OH^- combine and give NaOH .

The Nelson cell is shown in the figure 7.7.

During the electrolysis Cl_2 is given out at the anode, H_2 at the cathode and NaOH is collected from the base of the cell.

Following reaction occur in the cell.



The chlorine and Hydrogen produced here as by product and are filled in cylinders, transported for marketing, while aqueous NaOH is dehydrated and dried up, filled in bags and is sold in market.

7.7 Corrosion and its Prevention

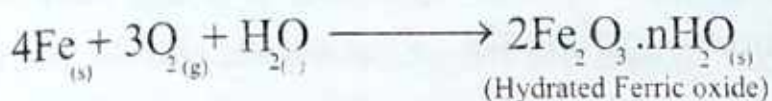
Corrosion is the term applied to metal rust caused by an electrochemical process. Most of you are familiar with the corrosion of iron, in the form of reddish brown rust. There are some other examples of corrosion too, like the black tarnish on silver, and red or green corrosion that may appear on copper and brass.

The corrosion involves the deterioration of essential properties in a material. The corrosion is actually due to redox reactions that occur in the presence of oxygen, moisture and acidic oxides i.e. CO_2 etc. corrosion damages metals and converts into oxides and finally into hydroxides e.g. Fe is converted into $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$ (rust). The corrosion causes a great damage to the metal and other objects in the world costing billions of dollars every year, and also causes accidents resulting in the loss of precious loss of lives.

7.7.1 Rusting Of Iron

The rusting of iron chemically means the electrochemical reaction of iron with oxygen and water to form $\text{Fe}(\text{OH})_3$, $\text{Fe}(\text{OH})_2$, or even $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$.

Chemically the corrosion of Iron is written as:



You may have observed this rusting around yourself at many places, the rusting of Iron which causes great loss of money and efforts can be reduced by applying following operations.

1. By painting the surface.
2. By coating the surface with a thin film of oil or grease.
3. By electroplating iron with some non-corrosive metal such as nickel, chromium, copper, etc.



Figure 7.7 Rust in Iron

7.7.2 Prevention of Corrosion

Each year billions of dollars are also expended to control the corrosion processes, most of these processes involve following techniques.

i) Coating

The coating is the most common and simple method for the prevention of corrosion. In this method, the metal or the object which is to be prevented from corrosion is coated with some material which seals it from the outer atmosphere; the most common coatings are done with the **paints or grease**, this cut off the metal from atmosphere and reaction between the O_2 and the metal is prevented.

ii) Electroplating

The electroplating, in which the metal which is being protected by corrosion is electroplated by some substance which is resistant to corrosion, e.g the ornaments are electroplated with the gold, or some objects with the copper. This reduces the corrosion significantly.

iii) Alloys Forming

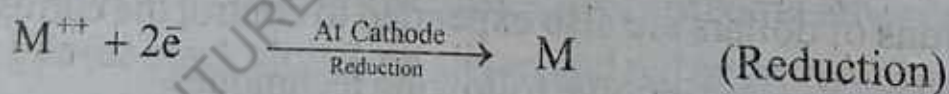
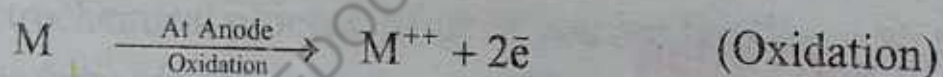
In this method, two or some time more metals are mixed to obtain a mixture of these metals, called **alloy**, which is very resistant to corrosion, e.g. the steel is an alloy of iron, chromium and nickel, Brass is an alloy of copper and zinc. These alloys are very resistant to corrosion.

7.8 The Electroplating Process

The Electroplating is the process of coating a metal object with a thin layer of another metal using the process of electrolysis. Usually the electroplated coating is about 0.002 inch (0.05 mm) thick.

Electroplating is carried out by constructing an electroplating cell in which the metal to be plated acts as a cathode and the metal to be deposited acts as anode, which is immersed in an aqueous solution of its salt.

During electrolysis following reaction occur.



Electroplating is used to give metal objects a better appearance or to protect them from corrosion, wear, or rust. Tableware, trays, decorative pieces, and jewelry are plated with gold or silver to make them more attractive.

There are several types of electroplating processes, classified on the basis of type of compound which is used for electroplating coatings; some of these electroplating processes are described below:

7.8.1 Electroplating of Tin

Tin is a malleable, ductile, highly crystalline, silvery-white metal. This metal resists corrosion from distilled, sea, or soft tap water, but can be attacked by strong acids, alkalis, and by acid salts.

Tin Plating is used extensively to protect both ferrous and nonferrous surfaces. Metals plated using tin are useful for the food processing industry since they are non-toxic, ductile and corrosion resistant. The excellent ductility of tin allows a coated base metal sheet to be formed into a variety of shapes without damage to the surface tin plating layer. Plating with Tin provides sacrificial protection for copper, nickel and other non-ferrous metals, but not for steel.

During electroplating, a mixture of tin salts is dissolved in a water. When an electric current is passed through water, the tin ions will be attracted to a metal object placed in the water, creating a permanent coating of tin.

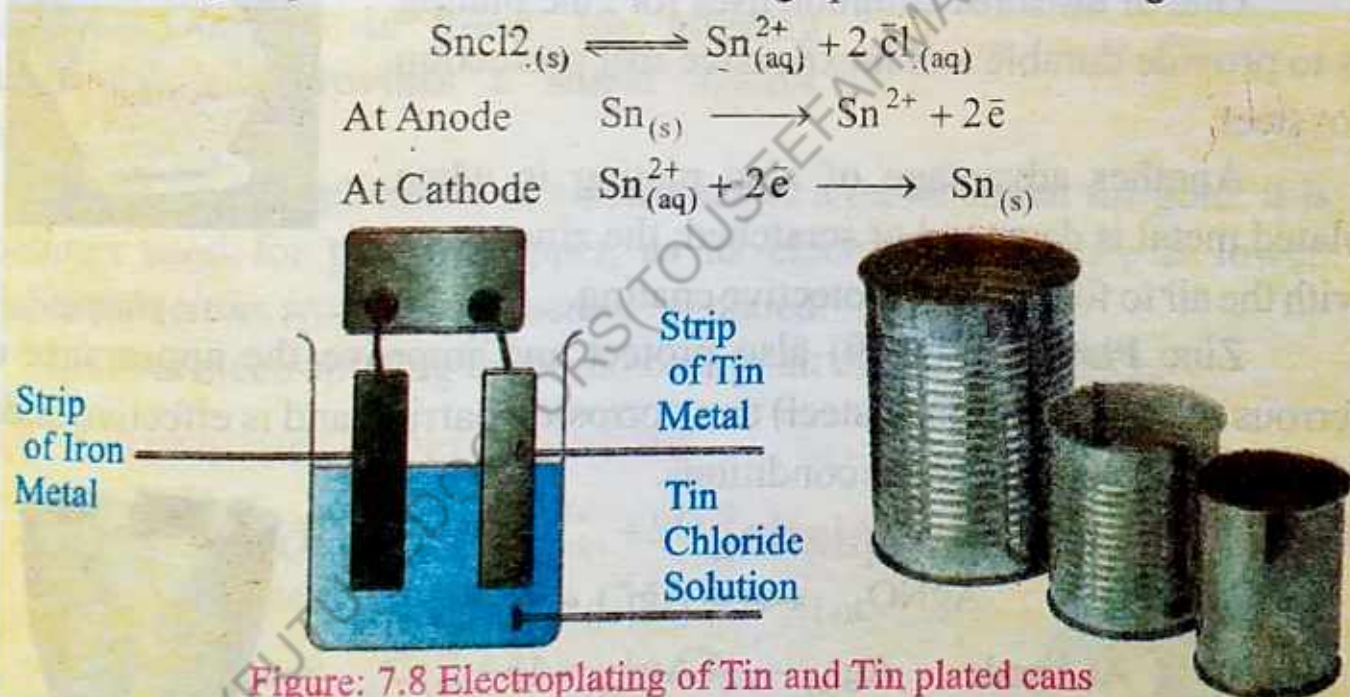


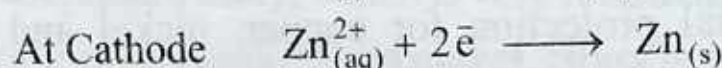
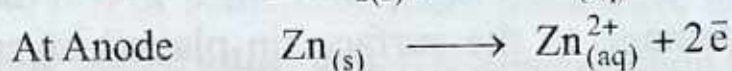
Figure: 7.8 Electroplating of Tin and Tin plated cans

Tin plating is an effective method of protecting base metals like iron or steel from rust and corrosion. While zinc plating can also be used to improve corrosion resistance, zinc is too toxic for food or medical applications, so tin plating must be used for these types of projects. Tin plating also increases the ductility of a metal object, making it easier to weld to another surface.

One potential drawback to tin plating is that it is not as effective as zinc plating, or galvanizing, when it comes to protecting materials from rust and

corrosion. When galvanized metal is damaged or scratched, the zinc reacts with the air to form a new protective coating. When tin plated metal is damaged or scratched, the underlying metal is left exposed to rust and corrosion. Tin is also a relatively soft metal, making it highly susceptible to damage. This material cannot hold up under extremely cold temperatures, and should not be used in freezing conditions.

7.8.2 Zinc Plating



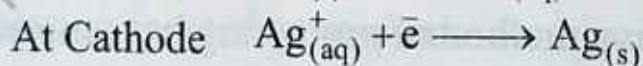
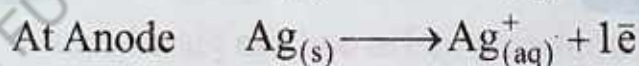
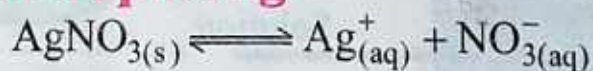
Zinc plating is also an important plating process which involves less hazardous chemicals than most other forms of plating, and is successfully done in small shops and even at homes.

One of the most common uses for zinc plating is to provide durable and inexpensive rust protection for steel.

Another advantage of zinc plating is when plated metal is damaged or scratched; the zinc reacts with the air to form a new protective coating.

Zinc Plating (Fig.7.9) also protect and improve the appearance of ferrous metals, (i.e. iron & steel) as a corrosion barrier, and is effective under high humidity and moisture conditions.

7.8.3 Silver Electroplating



Silver is a white colored semi precious metal with an oxidizing property. Due to its property of high electrical conductivity it is used in electronics and semiconductor industries. Silver is also an excellent conductor of heat and offers great solderability. It is malleable and has a high degree of ductibility so it can be



Figure 7.9:
Zinc plating



Figure 7.10:
Silver plated utensil

bent into any shape and is hence preferred by jewelry and flatware industries for its decorative appeal.

The silver also acts as an economically beneficial metal that may be put to good use for a relatively cheaper price as compared to other precious metals like gold or platinum. (Figure:7.10)

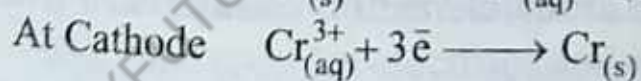
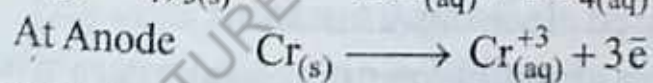
The most common usage of silver plating is for ornamentation of jewelry, plates, cups, trophies and medals. The process adds a perfect polish to the surface on which the plating is done.

Silver plating improves solderability by hardening the surface and increasing wearability. It is also used to inhibit corrosion. The plating helps in reducing friction and to improve paint adhesion. It is good for altering conductivity and provides a shield against radiation.

Silver finds application in electronics as a replacement for gold, it is sometimes used for plating copper, as its electrical resistance is lower. Variable capacitors work best when silver plated.

In silver electroplating the common Ag salt used is the AgNO_3 .

7.8.4 Chromium Electroplating



The chrome plating involves the electroplating of a thin veneer of chromium onto an underlying metal. Sometimes used for decorating purposes, chrome plating can also serve as a layer of protection as well. Often, chrome plating is applied to metal objects when there is an increased opportunity for corrosion to occur.

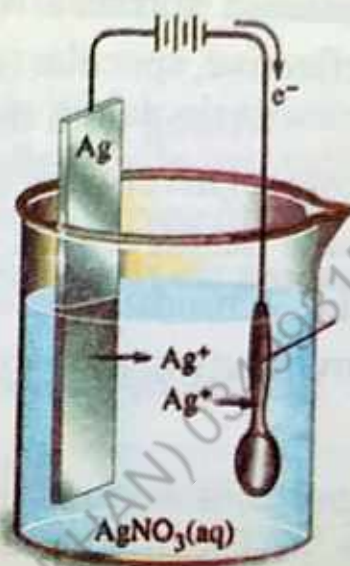


Figure 7.11:
Silver Plating Process



Figure 7.12:
Chromium plated tap

This plating provides corrosion resistance, eases cleaning procedures, increases the surface hardness and also offers a decorative shine.

Though differentiating between chrome plated surface and other polished surfaces is tough; yet it is possible as chrome plating is more reflective, specular (gives mirror like reflection) and bluer than others.

Along with this, chrome plating also have some disadvantages e.g. it costs relatively much higher than the zinc and tin plating, moreover it also has some health effects, as Chromium VI is poisonous and causes even cancer. The Chrome plating also produces some compounds which create environment problems.

SUMMARY OF THE CHAPTER

- ❖ Electrochemistry is a branch of chemistry that studies the reactions which take place at the interface of an electronic conductor called electrode.
- ❖ In the electrochemistry study of oxidation and reduction processes is carried out.
- ❖ Oxidation is the process of adding oxygen, removing hydrogen or removing electrons.
- ❖ Reduction is a process of adding hydrogen, removing oxygen or adding electrons.
- ❖ The apparent charge upon an atom in a compound is called its **Oxidation state** or **Oxidation number**. Remember it is **quite different from the valency** which is the combining capacity of an atom or the number of unpaired electrons in an atom.
- ❖ Oxidation number of an atom /element is always zero when it is in free state.
- ❖ An **Oxidizing agent** is the species that causes oxidation of another species.
A reducing agent is the species that reduces another species.
- ❖ Oxidation and reduction reactions involve migration of electrons from one species to another species, these reactions always occur side by side.
- ❖ Collectively the oxidation and reduction reactions are called **Redox reactions**.

- ❖ The cells where the chemical reaction uses or produces electricity are called **Electrolytic cells**.
- ❖ The electrochemical cells are of two types, one that uses electricity is called **electrolytic cells**, in such cells electricity is used to breakup a compound into its constituents.
- ❖ Second type of electrochemical cells are those in which electricity is produced by chemical reactions, such cells are called **Galvanic cells** or **Daniell cells**, the Lead-acid battery (car battery) is such type of cell.
- ❖ Many important industries e.g electroplating, electrolysis etc are based on electrochemical cells.
- ❖ Many important compounds like NaOH, Sodium metal etc are obtained by using electrolytic cells.
- ❖ **Corrosion** of metals is a redox reaction, which involves reaction of metals with oxygen, moisture of air.
- ❖ Corrosion can be reduced by or prevented by making alloys, by painting or by electroplating processes.

EXERCISE

Q1. Fill in the blanks with appropriate words.

- i) Reduction is the process which involves of electrons.
- ii) In the free state of an element, the oxidation state has always value.
- iii) The Galvanic cell produces using redox reaction.
- iv) Coating with paints prevents
- v) Dry cells are also a type of
- vi) Electroplating brings both the and resistance to
- vii) The food cans are electroplated with
- viii) Caustic soda is industrially prepared by cell.
- ix) $K_2Cr_2O_7$ is a strong agent.
- x) Collectively the oxidation and reduction processes are called

Q2. Choose the correct answer.

- i) The Semiconductors are the species that:
 - (a) conduct electricity easily

- (b) do not conduct electricity easily
- (c) conduct electricity but not heat
- (d) conduct heat but not electricity
- ii) Oxidation means:
 - (a) addition of electrons
 - (b) removal of electrons
 - (c) addition of protons
 - (d) removal of protons
- iii) Oxidation state is the:
 - (a) apparent +ve charge
 - (b) apparent -ve charge
 - (c) apparent charge
 - (d) not related with charge
- iv) The electroplating prevents:
 - (a) oxidation
 - (b) reduction
 - (c) both oxidation and reduction
 - (d) neither oxidation nor reduction
- v) Steel is the alloy of:
 - (a) iron and mercury
 - (b) mercury and copper
 - (c) iron and copper
 - (d) iron, chromium and nickel

Q3. Answer the following questions in short.

- i) Define electrochemistry. Also state its importance.
- ii) What is oxidation? Illustrate with an example.
- iii) Explain the term reduction with an example.
- iv) Draw a sketch of an electrolytic cell, showing its different parts.
- v) What are oxidizing and reducing agents? Explain with example.
- vi) What is the difference between weak and strong electrolytes explain with examples.
- vii) What is electroplating? Also describe its importance.
- viii) What is electrolysis?
- ix) Describe importance of Alloy formation in industries.
- x) Explain why the food articles are packed in tin plated cans rather than chrome plated cans.

Q4. Answer the following questions with reasoning.

- i) Differentiate between oxidation number and valency.
- ii) How electricity is produced from chemical reactions?
- iii) Explain how oxidation and reduction goes side by side in chemical processes?

- iv) Explain how NaCl yields two different products i.e. Na-metal and NaOH by the same process of electrolysis.
- v) How you can prevent a piece of iron from corrosion?
- Q5. Distinguish between the oxidation and reduction processes?
- Q6. What is the oxidation number? Describe the rules to calculate the oxidation number.
- Q7. Calculate oxidation number of Mn in KMnO_4 .
- Q8. What do you know about the electrochemical industries?
- Q9. What is corrosion? How can you prevent the corrosion process?
- Q10. Describe the working of an electrolytic cell, How Sodium metal can be purified by using electrolytic cell?
- Q11. What is the electrochemical cell? Discuss the different types of electrochemical cells?
- Q12. Discuss the importance of Caustic soda? How it is prepared on industrial scale?
- Q13. How the electricity is prepared by chemical means?
- Q14. Write a note on electroplating process.

CHEMICAL REACTIVITY

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the:

- Symbols of Elements
- Writing equations

TIME ALLOCATION

Teaching periods	= 07
Assessment periods	= 02
Weightage	= 06%

LEARNING OUTCOMES

After reading this chapter, the Student will be able to:

- Show how cation and anions are related to the terms metals and non-metals. (Applying)
- Explain why alkali metals are not found in free state in nature. (Analyzing)
- Identify elements as an alkali metals or an alkaline earth metal. (Applying)
- Explain the difference in ionization energies of alkali and alkaline earth metals. (Understanding)
- Describe the position of sodium and magnesium in periodic table, their simple uses and properties. (Understanding)
- Differentiate between soft and hard metals. (Analyzing)
- Describe the inertness of noble metals. (Understanding)
- Identify the commercial value of Silver, Gold and Platinum. (Analyzing)
- Compile some important reactions of halogens. (Applying)
- Name some elements, which are found in uncombined state in nature. (Understanding)

8.0 Introduction to Chemical Reactivity

The chemical reactivity of a substance is the ability of that substance to undergo chemical changes. During chemical reactions, the molecules or atoms come close to each other and linked through bonds to form a new compound, (Fig 8.1).

The Chemical reactions are in process all around us, the process of cooking food, digestion of this food stuff in our

body, our clothing, medicines, burning of gasoline in our vehicles all these processes are the Chemical reactions. In this chapter we will discuss this process of chemical reactivity.

The reactivity of a substance is due to the tendency of donating or accepting or sharing the electrons, determined by the electropositive or electronegative characters. As the values of these characters increases, the ability to react also increases.

It is to be noted that the substances react to gain the lowest energy state, where they expend less energy to exist. The elements in this regard attain the electronic configuration of the nearest noble gas, which is the least energy state.

The chemical reactions on the basis of their speed are divided into 3 categories, the slow reactions, moderate reactions and fast reactions. Rusting of iron is the example of slow reaction, the organic reactions mostly are of

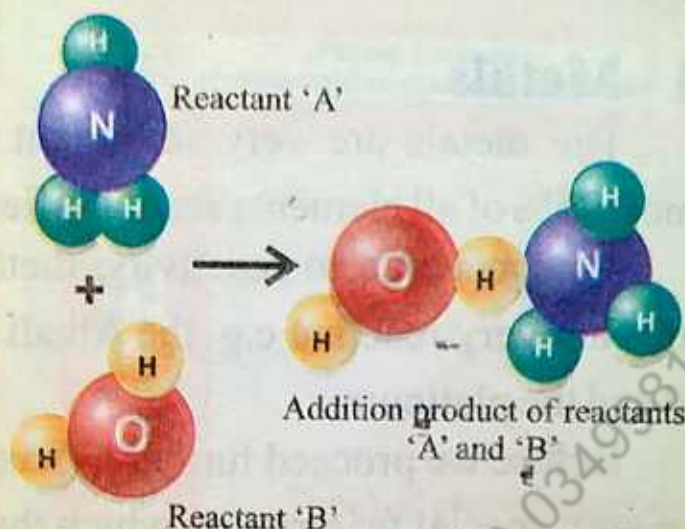


Figure: 8.1: Molecular view of an addition product by the reaction of two reactants 'A' and 'B'

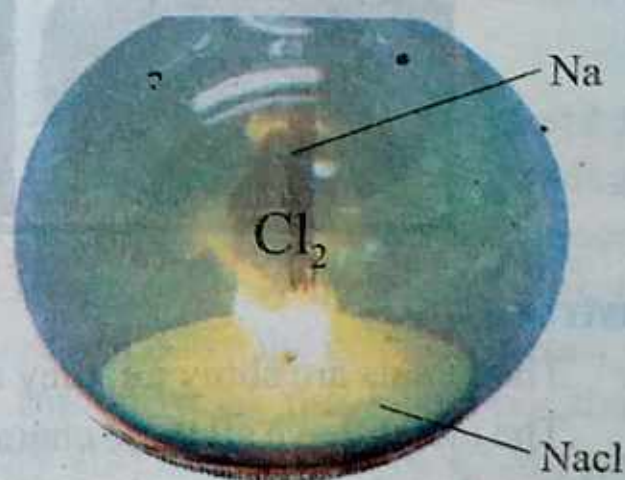


Figure: 8.2 Reaction between Sodium Metal and Chlorine gas is very fast and exothermic, as seen here in this picture.

moderate speed, and reactions between the ionic compounds are very fast (fig. 8.2).

8.1 Metals

The metals are very important and vital of our use in this world. Almost 75% of all elements are classified as metals.

As concerned to reactivity, metals show mix behavior, some of the metals are very reactive e.g. the Alkali metals, and some are quite inert like gold and the platinum.

Before we proceed further let's understand the nature of metals, which have very special features, by which they are recognized and separated from the other elements. These characteristics are:

Conduction

Metals are good at conducting electricity and heat, so they are used in almost all applications of electricity.

Ductile and Malleable

The metals have very special character, that they can be transformed in any desired shape. Moreover metals can be drawn into wires of small diameters, (ductile). (Fig 8.3)



Figure 8.3: Copper made coin and wire

Lustrous

The metals are shiny i.e. they are lustrous.

The metals have all these characters, and a single of these characters is if lacking, then the compound is not categorized as a metal. e.g. the plastics, although are malleable and ductile, and may have luster too, but they are not conductors, hence they are not metals.

8.1.1 Electropositive Character of Metals

The metals are electropositive in nature, i.e. they tend to lose electrons during the chemical reactions, and hence gain positive charge. The relative reactivity of elements depends upon the electropositive or the electronegative character of the elements, this is because, the chemical reactions involve the transfer of electrons from

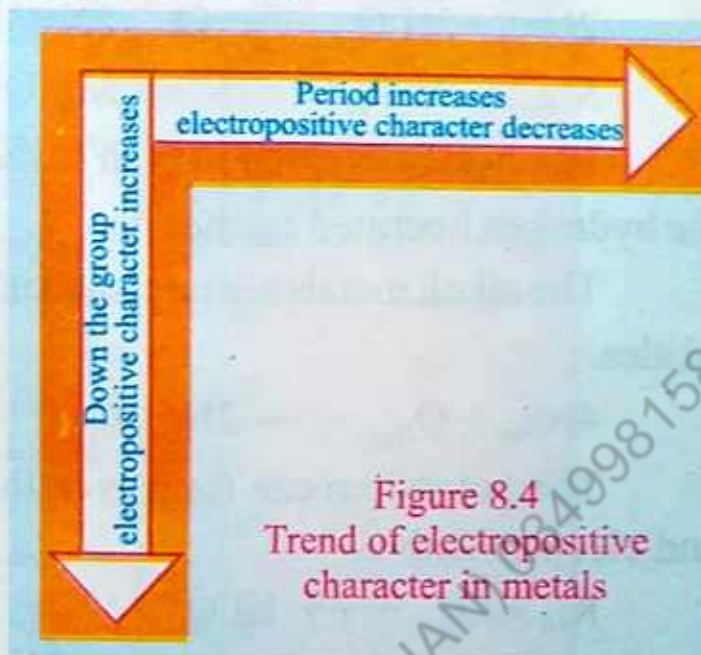


Figure 8.4
Trend of electropositive character in metals

one to another atom and hence new compounds are formed. The elements that have greater tendency to donate or accept the electrons are more reactive. In the periodic table the electropositivity of metals increases down the groups and decreases along the period.

Hence the reactivity of the metals is also increases down the groups, and reactivity of metals decreases along the periods.

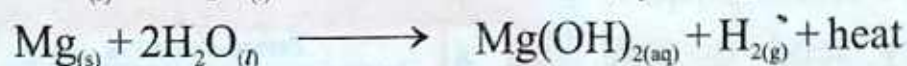
But remember that this is not the case with the non-metals, where the electronegative character counts, rather than the electropositive character, which is just opposite to the electropositive character.

8.1.2 Comparison of Reactivity between Alkali Metals and Alkaline Earth Metals

The alkali metals comprise the Gr-1 of the alkali metals, and the alkaline earth metals comprise Gr-2 of the periodic table. The alkali metals are soft, very reactive having a high value of electropositivity. They do not occur in free state. As described earlier, the electropositivity increases down the group of metals, so the reactivity also increases down the group in metals, so "K" is much reactive than the "Na". Similarly, as the electropositivity decreases along a period, so the reactivity also decreases, hence "Mg" and "Ca" are less reactive than their counterparts "Na" and "K" respectively.

For example, the alkali metals react vigorously with water to form the product hydroxides, while the alkaline earth metals react comparatively

slowly.

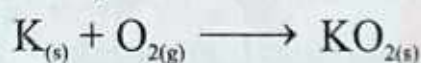


The heat is evolved in both cases, but in the first case it is so large that the hydrogen liberated catches fire.

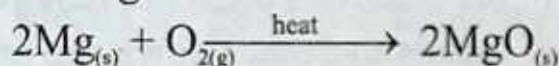
The alkali metals are tarnished in air quite easily to form corresponding oxides.



The potassium can form even the superoxide, being more reactive and electropositive.



In comparison to this, the alkaline earth metals react with oxygen of air by heating.



8.1.3 Inertness of Noble Metals

The noble metals are the metals that are resistant to corrosion and oxidation in moist air. They are usually precious and rare in the Earth's crust palladium, silver, osmium, iridium, platinum, and gold are the common examples of the noble metals, of these the silver, gold and copper are also called coinage metals.

Most of the noble metals are used in jewelry making.

The noble metals tend to be inert toward the common reagents, and do not dissolve in common substances, that's why the jewelry made up of these metals do not tarnishes after long periods.

The important reason for the inertness of the noble metals is their

Interesting Information

Due to inertness of noble metals they are used for making jewelry articles, as well as for making some other products, e.g., silver is used in mirror making in photographic films, in dental fillings with other compounds. Platinum is used as catalyst, in LCD screen, in computer CD and hard disk. Gold is used in electronics industry for making many electronic parts.

Group	Period	1A	2A	3A	4A	5A	6A	7A	8A	9A	10A	11A	12A	13A	14A	15A	16A	17A	18A
1	1	H																	
2	2	Li	Be	B	C	N	O	F	Ne										
3	3	Na	Mg	Al	Si	P	S	Cl	Ar										
4	4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	6	Cs	Ba	*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	7	Fr	Ra	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo
8	8																		
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114	114																		
115	115																		
116	116																		
117	117																		
118	118																		

Figure 8.5: Periodic table of the elements showing metals, non-metals

electronic arrangement. They have the stable electronic configuration and hence do not need to accept or donate electrons from others, so they tend to be inert at normal conditions.

8.2 Non-metals

There is no rigorous definition for the term "nonmetal", it covers a general spectrum of behavior. The common properties considered characteristic of a nonmetal include:

- Poor conductors of heat and electricity when compared to metals.
- They form acidic oxides (whereas metals generally form basic oxides).
- In solid form, they are dull and brittle, rather than metals which are lustrous, ductile or malleable.
- Usually have lower densities than metals.
- They have significantly lower melting points and boiling points than metals.
- Non-metals have high electronegativity.

Only seventeen elements in the periodic table are generally considered nonmetals, compared to over eighty metals, but nonmetals make up most of the crust, atmosphere and oceans of the earth. Bulk tissues of living organisms are composed almost entirely of nonmetals. Many nonmetals (hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine) are diatomic while rest are polyatomic. In the periodic table the non-metal elements are located at the right corner, as shown in figure 8.5.

In the non-metals, the electronegative character decreases down the group, and hence the reactivity also decreases. So the "F" is more reactive than the "Cl", which is more reactive than the "Br".

8.2.1 Electronegative Character

As already explained in section 3.2.5 that electronegativity is the ability of an atom to attract shared electron pair, this character is much more effect in non metals rather than metals because of the high nuclear charge and smaller size of non metals as compare to metals.

The electronegative character increases along the period and decreases down the group, (noble gases are the exception). Thus fluorine is the element which has highest electronegativity value of 4.0.

The electronegative character and reactivity of elements are directly related with each other, so fluorine is also the most reactive

element in non metals and reactivity of all non metals decreases down the group and increases along the period.

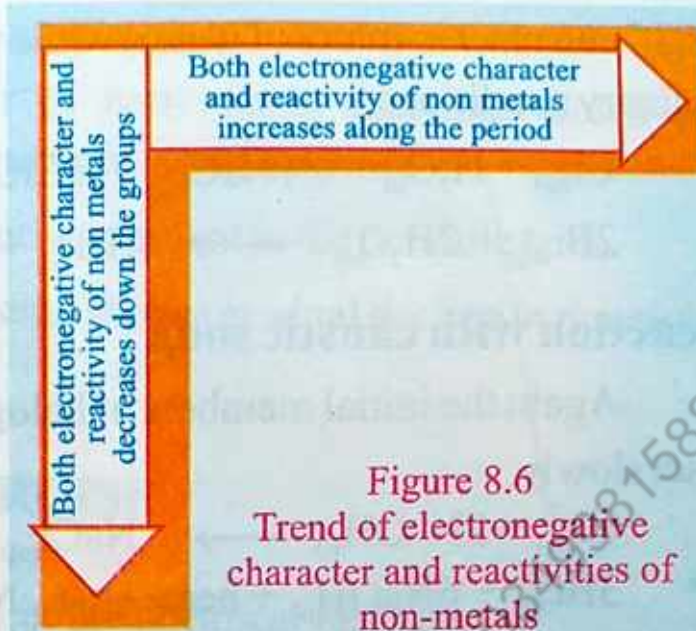


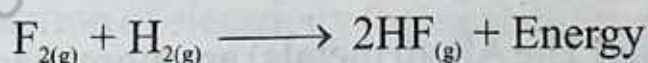
Figure 8.6
Trend of electronegative character and reactivities of non-metals

8.2.2 Comparison of Reactivity of the Halogens

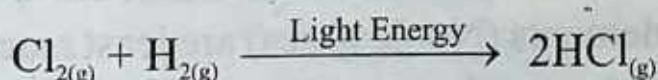
In the periodic table, the halogens are the most electronegative in their respective period, and so they are also most reactive among the non-metals of their periods. Among themselves, the halogens show a decline in the electronegativity down the group, so their reactivity also decreases down the group. This is further explained by the following reactions.

Reaction with hydrogen

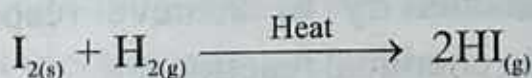
All halogens react with hydrogen, but under different conditions, for example, the "F₂" reacts violently,



Cl₂ reacts in a little presence of energy (light energy is sufficient),



And Iodine shows sluggish reaction.



Reaction with water

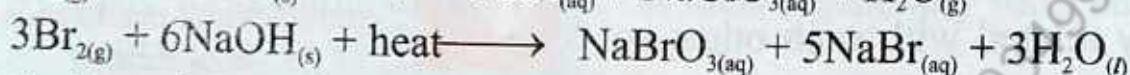
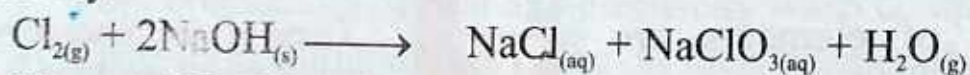
The "F₂" reacts explosively with water, "Cl₂" and "Br₂" react readily,

("Cl₂" can react in traces of water) while the "I₂" do not reacts with water under ordinary conditions.



Reaction with caustic soda

Again the initial members of halogens reacts fast, and the last members react slowly.



As it can be seen in equations that Cl₂ is reacting directly with caustic soda, while Br₂ reacts on heating, this proves the decline in reactivity down the group.

SUMMARY OF THE CHAPTER

- ❖ Chemical reactivity of a substance is the ability of a substance to undergo chemical reaction, this is because of the tendency of every substance to attain the lowest energy state.
- ❖ Chemical reactions involve migration of electrons.
- ❖ Metals are electropositive in nature, i.e they tend to donate electrons during the chemical reactions.
- ❖ The electropositive character of metals increases down the group and decreases along the periods.
- ❖ Group 1 elements (The Alkali metals) are more reactive than the Group 2 (alkaline earth metals) elements.
- ❖ The Group 18 elements (Noble gases) are least reactive of all elements in the periodic table, and they are also called inert gases.
- ❖ All the other elements try to achieve respective the noble gas configuration during chemical reactions.
- ❖ The non-metals which are situated at right corner of the periodic table are electronegative in nature, i.e. they accept electrons during chemical

reactions.

- ❖ The electronegative character of non metals decreases down the groups, and increases from left to right along a period.
- ❖ Fluorine is the most electronegative element in the periodic table.
- ❖ The Group 17 elements (Halogens) show a gradual decline in reactivity down the group.

EXERCISE

Q1. Fill in the blanks with appropriate words.

- i) are involved in the chemical reaction.
- ii) Metals are in nature while are electronegative in nature.
- iii) The are most reactive metals.
- iv) The noble gases are least of all elements.
- v) When difference between electronegativities is more than 2, the nature of bond is
- vi) The alkali metals found only in state.
- vii) Iron is a hard metal while Sodium is metal.
- viii) Among halogens is most reactive.
- ix) The chemical reactions involve a change.
- x) NaOH is also known as

Q2. Choose the correct answer.

- i) The metals are characterized by:
(a) conduction of electricity (b) malleability and ductile nature
(c) lustrous nature (d) all of above
- ii) In a group the reactivity of metals:
(a) increases (b) decreases
(c) no effect at all (d) increases first and then decreases
- iii) The electronegativity is the tendency of an atom:
(a) to attract electron pair of bond
(b) to repel electron pair of bond
(c) to attract another atom (d) to repel another atom

- iv) The halogen means:
(a) water producer (b) salt producer
(c) taste producer (d) steam producer
- v) The noble gases do not react due to:
(a) smaller size (b) less charge
(c) stable electronic structure (d) high electronegativity

Q3. Answer the following short questions.

- Based on the speed how many types of chemical reactions exists.
- Define term metals.
- How electropositive character of metals show trend in periodic table?
- What is electronegativity? How electronegativity shows variation in periodic table?
- Write down three most abundant elements of human and of earth crust?
- Write chemical equations of three reactions of water.
- What is caustic soda? Write three uses of caustic soda
- Differentiate between noble gases and noble metals?
- Write uses of halogens.
- Differentiate between soft and hard metals.

Q4. Answer the following questions with reasoning.

- What is the relation between chemical reactivity and the energy of a system?
- Explain what electronic configuration Mg and Br_2 achieve after they react.
- Why K is more reactive than Mg?
- Explain why group 18 elements do not react?
- Among halogen which member is most reactive and why?

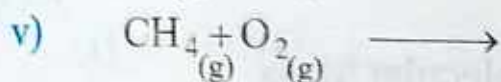
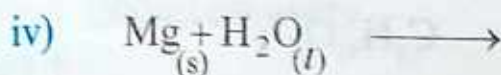
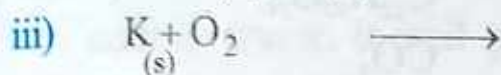
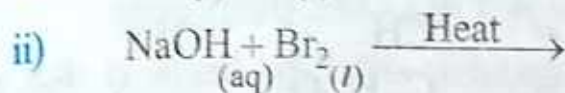
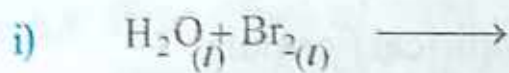
Q5. What is the metal? How they differ from non metals?

Q6. What is the difference between reactivity of metals and non metals?

Q7. What is electronegativity? How it effects the nature of the chemical bonding and bond polarity?

Q8. Compare the reactivity of alkali metals and alkaline earth metals?

- Q9. What are halogens? Discuss their chemical reactivity.
- Q10. Give an account on the properties of non metals? Also describe their importance.
- Q11. Complete and balance the following chemical reactions.



Appendix "A"

Answers to Activities:

Activity 1.2: Compound

Benzene

Sucrose

CO₂

(C₆H₁₀O₅)_n

Empirical Formula

= CH

= C₁₂H₂₂O₁₁

= CO₂

= C₆H₁₀O₅

Activity 1.3: Compound

NaCl

CO₂

Glucose

Al(OH)₃

Molecular Mass

= 58.5

= 44

= 180

= 78

Activity 1.4:

100gms of CO₂ has 4.101×10^{24} atoms

100gms of Glucose has 8.027×10^{24} atoms

Activity 2.1:

N = $1s^2, 2s^2, 2p^3$

P = $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

Activity 4.1

O₂ = $\begin{array}{c} \times \times \\ \text{O} \text{---} \text{O} \\ \times \times \end{array}$

CO₂ = $\begin{array}{c} \times \times \quad \times \times \\ \text{O} \text{---} \text{C} \text{---} \text{O} \\ \times \times \quad \times \times \end{array}$

NH₃ = $\begin{array}{c} \times \times \\ \text{H} \text{---} \text{N} \text{---} \text{H} \\ \times \times \\ \text{H} \end{array}$

N₂ = $\times \text{N} \text{---} \text{N} \times$

C₂H₆ = $\begin{array}{c} \text{H} \quad \text{H} \\ \times \quad \times \\ \text{H} \text{---} \text{C} \text{---} \text{C} \text{---} \text{H} \\ \times \quad \times \\ \text{H} \quad \text{H} \end{array}$

Activity 5.1: 310K

Activity 6.1: 3%

Activity 6.2: If 10.5g NaHCO_3 is dissolved in 250ml water, it gives 0.5M sodium bicarbonate solution.

Activity 6.3: If 16.67ml of 12M H_2SO_4 is dissolved in 400ml, it yields required 0.5M H_2SO_4 solution.

Activity 6.4: If 40ml of 10% solution of Glucose is diluted to 200ml by adding water, it will give required 2% solution of Glucose in 200ml quantity.

Activity 7.1

i) HNO_3 = +5

ii) NO_2 = +4

iii) NOCl = +3

Appendix "B"

Answer of numerical problems in exercise.

Chapter No.1

Q5. 1mg of Nicotine (in an average cigarette) has 3.716×10^{18} molecules

Q6. Compound Molar Mass

K_2SO_4 = 174

MnCl_2 = 126

$\text{Al}(\text{OH})_3$ = 78

$\text{Fe}(\text{NO}_3)_3$ = 242

$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ = 249.5

Q8. 16.33g of 'S' = 0.51 moles of 'S'

Chapter No.5

Q10. 525.17cc

Q11. 1.2 atm

Chapter No.6

Q7. 2.38M

Q8. 45gms

Chapter No.7

Q7. In KMnO_4 'Mn' has +7 Oxidation Number

Glossary of Chemical terms

Absolute zero: The lowest temperature that can be obtained, equal to -273.16°C or 0 K .

Accuracy: The closeness of a measurement to an accepted value.

Acid: A substance that when dissolved in water dissociates and can donate a hydrogen (proton) to another molecule. Examples include sulphuric acid and nitric acid. Under these conditions nitric acid dissociates according to the following equation:



Activation energy, (Ea): It is the minimum energy that is required for a successful reaction between colliding molecules.

Addition reaction: A reaction whereby two or more substances react together to form a product whose molecular mass is the sum of the molecular masses of the reactants.

Adsorption: If any compound, solid, liquid or gas, is loosely held by weak attraction to the surface of a solid it is said to have undergone adsorption. This process is much weaker and less permanent than absorption.

Alcohol: An organic compound that contains the -OH functional group.

Aldehyde: An organic compound containing the -CHO group.

Alkali: Water soluble hydroxide of one of the alkali metals.

Alkali metals: Group IA elements.

Alkaline earths: Group IIA elements.

Allotrope: Different physical forms of an element are known as allotropes. For example, diamond and graphite are allotropes of carbon.

Alloy: An alloy consists of an intimate mixture of two elements, usually metals to give a metal compound or solid solution. Bronze is an alloy of copper and tin and brass of copper and zinc. By varying the composition it is possible to alter the chemical and physical properties of the alloy such as conductivity, ductility and corrosion resistance.

Alpha particles: These are particles possessing a positive charge, and are emitted from radioactive elements. The particles are actually helium nuclei, i.e. helium atoms that had lost their outer electrons. The study of the alpha particle played a major part in the development of modern atomic theory. Rutherford bombarded other elements with alpha particles to produce nuclear disintegration.

Amalgam: Liquid mercury has the ability to dissolve other metals to produce alloys and these alloys are called amalgams. Amalgam examples are gold/mercury, silver/mercury and copper/mercury all of which are used in dentistry.

Amorphous Solids: Solids that have indistinct shapes, because their particle arrangements lack order.

Amphoteric substance: It is the substance that act as a proton donor (an acid) in one reaction and a proton acceptor (a base) in a different reaction.

Anion: An atom, or group of atoms, carrying one or more negative charges. Cl^- and SO_4^{2-} are anions.

Anode: In an electrochemical cell or electrolytic cell, the anode is the electrode where oxidation takes place.

- Atmosphere:** Atmosphere is the cloud of gases around us, the main constituents of atmosphere are 78% Nitrogen, 21% Oxygen, 0.93% Argon gas and 0.03% Carbondioxide.
- Atom:** The smallest particle of an element which can take part in chemical reactions and display the chemical properties of that element.
- Atomic mass (Relative):** The atomic mass of an element (more correctly the relative atomic mass) is the mass of one atom of that element, divided by $1/12$ th of the mass of one atom of the carbon isotope ^{12}C . In older textbooks, the relative atomic mass was called the atomic weight.
- Atomic Mass Number (mass number, A):** The total number of protons and neutrons in the nucleus of an atom.
- Atomic mass unit:** The atomic mass unit is $1/12$ th of the mass of one atom of the carbon isotope ^{12}C .
- Atomic number:** The atomic number of an element is the number of protons in its nucleus.
- Aufbau principle:** The system of rules whereby electrons are accommodated in atomic orbitals.
- Avogadro number (N_A):** The number of atoms, molecules or ions, 6.02×10^{23} , contained in one mole of a given substance.
- Base:** A substance which can accept a proton, H^+ from an acid.
- Base unit:** One of the set of seven units on which the SI system is based.
- Beta particles:** Beta particles are very fast moving electrons. The particles have a velocity of between 30-99% speed of light, and it is this velocity which gives the beta particles greater penetrating power than alpha particles.
- Boiling point:** The temperature at which the vapour pressure of a liquid is equal to the atmospheric pressure. A liquid is said to "boil" when its temperature is raised to its boiling point.
- Bond:** A chemical link between two atoms. In an ionic bond the attraction is between opposite charges on two neighboring ions. In a covalent bond the two atoms share a pair of electrons.
- Boyle's law:** At constant temperature, the volume of a given mass of gas is inversely proportional to its pressure.
- Buffer:** A substance that resists changes in pH when moderate amounts of acid or base are added.
- Calorie:** An obsolete non-SI unit of heat energy. $1 \text{ calorie} = 4.184 \text{ J}$.
- Calorimeter:** A device that is used to measure enthalpy changes for chemical and physical processes.
- Catalyst:** A substance which speeds up the rates of certain chemical reactions, without being permanently used up in the process.
- Cathode:** In an electrochemical or electrolytic cell, the electrode where reduction takes place.
- Cation:** An atom or group of atoms bearing one or more positive charges. Cu^{2+} and NH_4^+ are cations.
- Caustic soda:** Sodium hydroxide, NaOH .
- Chalcogen:** Group VIA elements.

- Chemical bond:** The force that holds atoms together in the form of a compound.
- Combustion reaction:** The reaction of a substance with oxygen, producing oxides, heat, and light; burning.
- Compound:** A pure substance that consists of two or more elements chemically combined in fixed proportions.
- Concentration:** An expression of the amount of solute that is present in a given amount of solution.
- Constant-pressure calorimeter:** A calorimeter (such as a coffee-cup calorimeter) that is open to the atmosphere.
- Critical temperature:** The temperature above which a real gas will not liquefy, no matter how high its pressure.
- Covalent bond:** A chemical bond between two atoms formed by sharing an electron pair between these atoms.
- Crystalline solids:** Solids that have organized particle arrangements and, therefore, have distinct shapes.
- Daniell cell:** An electrochemical cell made up of a zinc anode and a copper cathode.
- Dilution:** The addition of water or another appropriate solvent to a solution in order to reduce the concentration of solute.
- Dipole:** A separation of negative and positive charge by a short distance. in molecular compound called a dipole.
- Dipole-dipole force:** An intermolecular attraction between opposite charges of polar molecules.
- Distillation:** A technique of purification based on the difference in boiling points of constituents.
- Dissociation:** The process in which a bigger molecule breaks into two or more simpler units.
- Double bond:** A covalent bond between two atoms resulting from the sharing of two pairs of electrons. Normally written as $C=C$, $C=O$, $C=N$, $N=O$.
- Electrochemical cell:** A device which can generate an electric current by exploiting the energy of a redox reaction between two electrodes (usually metals) having different electrode potentials.
- Electrode:** A conductor (usually a metal or graphite) in contact with an electrolyte, such as a salt solution of molten salt.
- Electrolysis:** The conduction of an electric current from an outside source through an electrolyte, whereby chemical reactions take place at the electrodes.
- Electrolyte:** A solution capable of conducting electricity, usually a solution of a salt in water. Alternately, a substance capable of conducting an electric current, such as a molten salt.
- Electron:** A particle of mass 9.11×10^{-31} kg, carrying the fundamental negative electric charge, -1.60×10^{-19} C.
- Electron affinity:** The energy liberated when an element acquires an electron to form an anion.
- Electronegativity:** A measure of the ability of an element to polarize a covalent bond in

which it takes part.

Electronic Configuration: Shorthand notation that shows the number and arrangement of electrons in an atom's orbitals.

Element: A substance which cannot be decomposed by chemical means into a simpler substance.

Empirical formula: A formula for a substance which expresses the simplest relationship between its constituent elements. For example, the empirical formula for ethanoic acid is CH_2O , although its molecular formula is $\text{C}_2\text{H}_4\text{O}_2$.

Endothermic reaction: A reaction that results in a net absorption of energy.

Enthalpy: The total kinetic energy and potential energy of a substance.

Entropy: A measure of the disorder in a system.

Enzyme: An enormous protein molecule that acts as a biological catalyst.

Evaporation: A physical process whereby a liquid passes into the gas phase, not necessarily at the boiling point of that liquid.

Excited state: An atom is said to be in an excited state if an electron has absorbed energy sufficient to promote that electron to an energy level which is higher than that in which it finds itself in the ground state. For example, the single electron of hydrogen in the ground state is in the level $n = 1$. If the electron is in the level $n = 2$ or higher, the atom is in an excited state.

Exothermic reaction: A reaction which takes place with a release of energy to the environment. In an exothermic reaction, the potential chemical energy of the products is less than that of the reactants.

Free Electron: An electron that is not bonded to an atom or molecule and so is free to move under external electric or magnetic fields.

Galvanic cell: See Daniell cell.

Gamma rays: Electromagnetic radiation with very high energy and frequency in the range 10^{20} Hz. These are the three types of rays produced by radioactive disintegration, but gamma rays are the only one comprised of electromagnetic radiation. Gamma rays have a very short wavelength, and have the greatest penetrating power of all radioactive emissions.

Greenhouse effect: Warming of the Earth's surface as a result of atmospheric pollution by gases. It is now feared that the warming effects are being undesirably increased, causing climate changes and melting polar icecaps.

Ground state: The state of an atom where all electrons have the minimum energy.

Halogens: This important group of elements occupies Group 17 of the periodic table.

Heavy water: A molecule of water contains two atoms of hydrogen and one of oxygen. Most water comprises of hydrogen/oxygen but a small percentage is composed of another hydrogen isotope, deuterium and oxygen. Deuterium differs from hydrogen by having one neutron in the nucleus of each atom. Deuterium water is given the formula D_2O . There is approx. 1 part in 5000 D_2O in normal water and it can be concentrated by electrolysis. D_2O had a higher boiling point (101.4°C) and melts at 3.6°C .

Homogeneous mixture: A mixture in which the different components are mixed so that they appear to be a single substance; a solution.

Heterogeneous mixture: A mixture in which the different components can be seen distinctly

Hydrides: Binary compound of hydrogen with other elements.

Hydrogen bonding: A special type of dipole-dipole interaction involving hydrogen atoms and electronegative elements such as oxygen and nitrogen.

Ideal gas: A gas which obeys the gas laws perfectly.

Indicator: A substance which shows a colour change after completion of a chemical reaction.

Inert gas: Group 18 elements.

Intermolecular forces: A force of attraction between molecules to hold them close.

Intramolecular forces: Forces that are exerted within a molecule or polyatomic ion; an example is covalent bonds.

Ion: An atom, or group of atoms, bearing an electrical charge. Ions may be monoatomic (Na^+) or polyatomic (SO_4^{2-}).

Isomer: Two compounds are said to be isomers if they have the same molecular formula, but different structural formulae.

Isotope: Two or more different forms of an element differing only in their mass numbers.

Kinetic energy: The energy of a body acquired by virtue of its motion.

Kinetic theory of gases: The concept that gases consist of perfectly elastic molecules of negligible volume, moving randomly, and creating pressure by collisions with the walls of the container.

Lattice energy: The energy liberated when a crystal lattice is formed by the electrostatic binding of its constituent cations and anions.

Limiting reactant: The reactant that is completely consumed during a chemical reaction, limiting the amount of product.

Litmus: An indicator of vegetable origin, which turns red in acids and blue in alkalis.

Mass number: The sum of the protons and neutrons in the nucleus of an atom.

Matter: Anything that has mass and takes up space.

Melting point: The temperature at which a transition from the solid to the liquid phase takes place. A solid is said to "melt" when this transition takes place.

Metal: The elements that have lustre, are good conductors and malleable.

Metallic bond: Interatomic forces holding metal atoms in a solid sample of a metal.

Mixture: A sample of matter which consists of two or more pure substances.

Molarity: Concentration of a solute expressed in units of moles dm^{-3} . A solution is said to be molar if it has a concentration of 1.0 mole dm^{-3} .

Mole: Atomic or molecular mass in grams.

SI unit for the amount of matter. One mole of a substance contains 6.023×10^{23} particles (atoms, molecules or ions) representative of that substance.

Molecular formula: A formula which reflects the actual number of each constituent atom in a molecule of a substance.

Monomers: Small molecules that are combined in long chains to produce very large molecules called polymers.

Neutron: An atomic particle having a relative atomic mass of 1, and zero charge.

Nucleus: That central part of the atom where its mass is concentrated, consisting of protons and neutrons.

Orbital: A region where the probability of finding of electron is maximum is called orbitals.

A region in space in an atom where probability of finding the electron is most. Orbitals are designated with symbols, s, p, d and f.

Organic chemistry: The chemistry of the hydrocarbons and their derivatives.

Oxidation: The process whereby a substance loses electrons.

Oxidizing agent: A substance which can accept electrons from another substance.

Oxide: Substance in which some element is combined with oxygen only.

Ozone: An allotrope of oxygen, O_3 .

p-block elements: Elements in Group 13 to Group 18, where the last electrons of an atom lies in p orbitals.

Periodic table: A table in which all the known elements are arranged in increasing values of their atomic numbers.

pH: A function that expresses the acidity or alkalinity of an aqueous solution. pH is related to the hydrogen ion concentration: $pH = -\log[H^+]$.

Phenolphthalein: An indicator which is colourless in acids and pink or purple in alkalis.

Precipitate: An insoluble solid that is formed by an aqueous chemical reaction between two soluble compounds.

Precision: The closeness of a measurement to other measurements of the same object or phenomena.

Pressure: Force exerted on a unit area. Pressure is measured in Pascal, (P), where $1 \text{ Pascal (P)} = 1 \text{ N.m}^{-2}$.

Proton: An atomic particle with relative atomic mass of 1, and unit positive charge, found in the nucleus of the atom.

Qualitative analysis: The branch of analytical chemistry that involves identifying elements, compounds, and ions in samples of unknown or uncertain composition.

Quantitative analysis: The branch of analytical chemistry that involves determining how much of a compound, element, or ion is in a sample.

Quantum mechanical model of the atom: A model that describes atoms as having certain allowed quantities of energy because of the wave-like properties of their electrons.

Radical: Atom or group of atoms having positive charge or negative charge or odd number are called radical.

Reactant: A starting substance in a chemical reaction.

Reaction mechanism: As series of steps that make up an overall reaction.

Reaction Rate: Speed of a chemical reaction.

Real gas: A gas which Exists in nature.

Reducing agent: A reducing agent (reductant) is a substance which can give up electrons to another substance, the oxidizing agent, or oxidant.

Reduction: A reaction in which electrons are taken up by a reactant. It is also defined as the addition of hydrogen or the removal of oxygen.

Redox Reactions: The oxidation and reduction reactions are collectively called redox

reactions.

Relative Atomic Mass: The mass of an atom or molecule compared with $C^{12.0000}$

Saturated solution: A solution in which no more solute will dissolve under the existing conditions of temperature and pressure.

s-block elements: Elements in Group 1 and Group 2, where the last electrons of an atom lie in s-orbital.

Single bond: a covalent bond in which one pair of electrons is shared between two atoms.

SI units: An international system of coherent units, formed by simple multiplication and/or division within a set of seven base units.

Solute: The substance dissolved in a solvent to form a solution, its quantity in the solution is always lesser.

Solution: A homogeneous mixture of variable composition consisting of two or more substances, one of which is usually a liquid.

Solvent: A liquid which dissolves a solute to form a solution, its quantity in the solution is always greater.

Solubility: The amount of solute that dissolves in a given quantity of solvent (normally 100ml) at a specific temperature.

Standard pressure: A pressure of 1 atmosphere = 760 mmHg.

Standard temperature: A temperature of 0.0°C , or 273.16 K.

S.T.P.: S.T.P. is the abbreviation of standard temperature pressure. (273.16 K and a pressure of 1 atmosphere).

Structural formula: A formula which defines the bonding and arrangement of the atoms present in a compound.

Sublimation: A phase change whereby a solid passes directly into the gas phase without melting.

Supersaturated solution: A solution which contains a higher concentration of solute than would be present if the solution had been saturated at the same conditions of temperature and pressure.

Temperature: The temperature of a sample of matter is a measure of the average kinetic energy of the molecules in that sample.

Thermodynamics: The study of energy and energy transfer

Unsaturated compound: A compound in which double or triple bonds are present.

Valence electrons: Electrons in the outer electronic shell of an atom, and which are involved in chemical reactions.

Valence shell: The outermost shell of electrons.

Vapour pressure: The pressure exerted by a vapour. At the boiling point, the vapour pressure of a liquid is equal to the pressure exerted on the liquid by the atmosphere.

Vander Waals forces: Weak forces between molecules of molecular solids and liquids.

These include dipole-dipole forces, dispersion forces and hydrogen bonds.

Voltaic cell:

A device which can generate an electric current by exploiting the energy of a redox reaction between two electrodes (usually metals) having different electrode potentials.

Water of crystallization:

Water molecules which form an integral part of the lattice of ionic crystals. The water of crystallization is normally included in the formula for the substance, i.e., $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

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TABLE OF ELEMENTS

Element	Symbol	At.No	At.mass	Element	Symbol	At.No	At.mass
Hydrogen	H	1	1.0079	Selenium	Se	34	78.96
Helium	He	2	4.0026	Bromine	Br	35	79.904
Lithium	Li	3	6.941	Krypton	Kr	36	83.798
Beryllium	Be	4	9.0122	Rubidium	Rb	37	85.4678
Boron	B	5	10.811	Strontium	Sr	38	87.62
Carbon	C	6	12.01	Yttrium	Y	39	88.9058
Nitrogen	N	7	14.0067	Zirconium	Zr	40	91.224
Oxygen	O	8	15.9994	Niobium	Nb	41	92.906
Fluorine	Fr	9	18.9984	Molybdenum	Mo	42	95.94
Neon	Ne	10	20.1797	Technetium	Tc	43	98.9063
Sodium	Na	11	22.9898	Ruthenium	Ru	44	101.07
Magnesium	Mg	12	24.305	Rhodium	Rh	45	102.9055
Aluminium	Al	13	26.9815	Palladium	Pd	46	106.42
Silicon	Si	14	28.0855	Silver	Ag	47	107.8682
Phosphorus	P	15	30.9738	Cadmium	Cd	48	112.411
Sulphur	S	16	32.065	Indium	In	49	114.818
Chlorine	Cl	17	35.453	Tin	Sn	50	118.71
Argon	Ar	18	39.948	Antimony	Sb	51	121.76
Potassium	K	19	39.0983	Tellurium	Te	52	127.6
Calcium	Ca	20	40.078	Iodine	I	53	126.9045
Scandium	Sc	21	44.9559	Xenon	Xe	54	131.293
Titanium	Ti	22	47.867	Caesium	Cs	55	132.9054
Vanadium	V	23	50.9415	Barium	Ba	56	137.327
Chromium	Cr	24	51.9961	Lanthanum	La	57	138.9055
Manganese	Mn	25	54.938	Cerium	Ce	58	140.116
Iron	Fe	26	55.845	Praseodymium	Pr	59	140.9076
Cobalt	Co	27	58.933	Neodymium	Nd	60	144.242
Nickel	Ni	28	58.6934	Promethium	Pm	61	146.9151
Copper	Cu	29	63.546	Samarium	Sm	62	150.36
Zinc	Zn	30	65.409	Europium	Eu	63	151.964
Gallium	Ga	31	69.723	Gadolinium	Gd	64	157.25
Germanium	Ge	32	72.64	Terbium	Tb	65	158.925
Arsenic	As	33	74.9216	Dysprosium	Dy	66	162.5

Element	Symbol	At.No	At.mass	Element	Symbol	At.No	At.mass
Holmium	Ho	67	164.93	Neptunium	Np	93	237.0482
Erbium	Er	68	167.259	Plutonium	Pu	94	244.0642
Thulium	Tm	69	168.93	Americium	Am	95	243.0614
Ytterbium	Yb	70	173.04	Curium	Cm	96	247.0703
Lutetium	Lu	71	174.967	Berkelium	Bk	97	247.0703
Hafnium	Hf	72	178.49	Californium	Cf	98	251.0796
Tantalum	Ta	73	180.9479	Einsteinium	Es	99	252.0829
Tungsten	W	74	183.84	Fermium	Fm	100	257.0951
Rhenium	Re	75	186.207	Mendelevium	Md	101	258.0986
Osmium	Os	76	190.23	Nobelium	No	102	259.1009
Iridium	Ir	77	192.217	Lawrencium	Lr	103	260.1053
Platinum	Pt	78	195.084	Rutherfordium	Rf	104	261.1087
Gold	Au	79	196.9665	Dubnium	Db	105	262.1138
Mercury	Hg	80	200.59	Seaborgium	Sg	106	263.1182
Thallium	Tl	81	204.283	Bohrium	Bh	107	262.1229
Lead	Pb	82	207.2	Hassium	Hs	108	265
Bismuth	Bi	83	208.98	Meitnerium	Mt	109	266
Polonium	Po	84	208.98	Darmstadtium	Ds	110	269
Astatine	At	85	209.98	Roentgenium	Rg	111	272
Radon	Rn	86	222.176	Copernicium	Cn	112	285
Francium	Fr	87	223.0197	Ununtrium	Uut	113	284
Radium	Ra	88	226.0254	Ununquadium	Uuq	114	289
Actinium	Ac	89	227.278	Ununpentium	Uup	115	288
Thorium	Th	90	232.0381	Ununhexium	Uuh	116	292
Protactinium	Pa	91	231.0359	Ununseptium	Uus	117	unknown
Uranium	U	92	238.0289	Ununoctium	Uuo	118	294

Table of Basic S.I units

S/No	Measurement	Unit	Symbol
1	Length	Metre or Meter	m
2	Mass	Kilogram	Kg
3	Time	Second	s
4	Electric current	Ampere	A
5	Temperature	Kelvin	K
6	Amount of substance	Mole	mol
7	Luminous intensity	Candela	cd

Some important derived units with their common names

S/No	Measurement	S.I unit	Commonly used Unit
1	Area	Square metre, (m^2)	m^2 and cm^2
2	Volume	Metre cube, (m^3)	m^3 and cm^3
3	density	Kilogram per metre cube (Kg/m^3)	g/cm^3
4	Energy	Kilogram metre square per second per second ($Kg.m^2.s^{-2}$)	Joule or J
5	Pressure	Newton per square metre (N/m^2)	Atmosphere or mmHg

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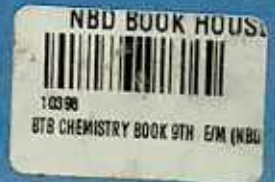
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