Test Edition

A Textbook of Chemistry

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Grade

IX

Khyber Pakhtunkhwa Textbook Board
Peshawar

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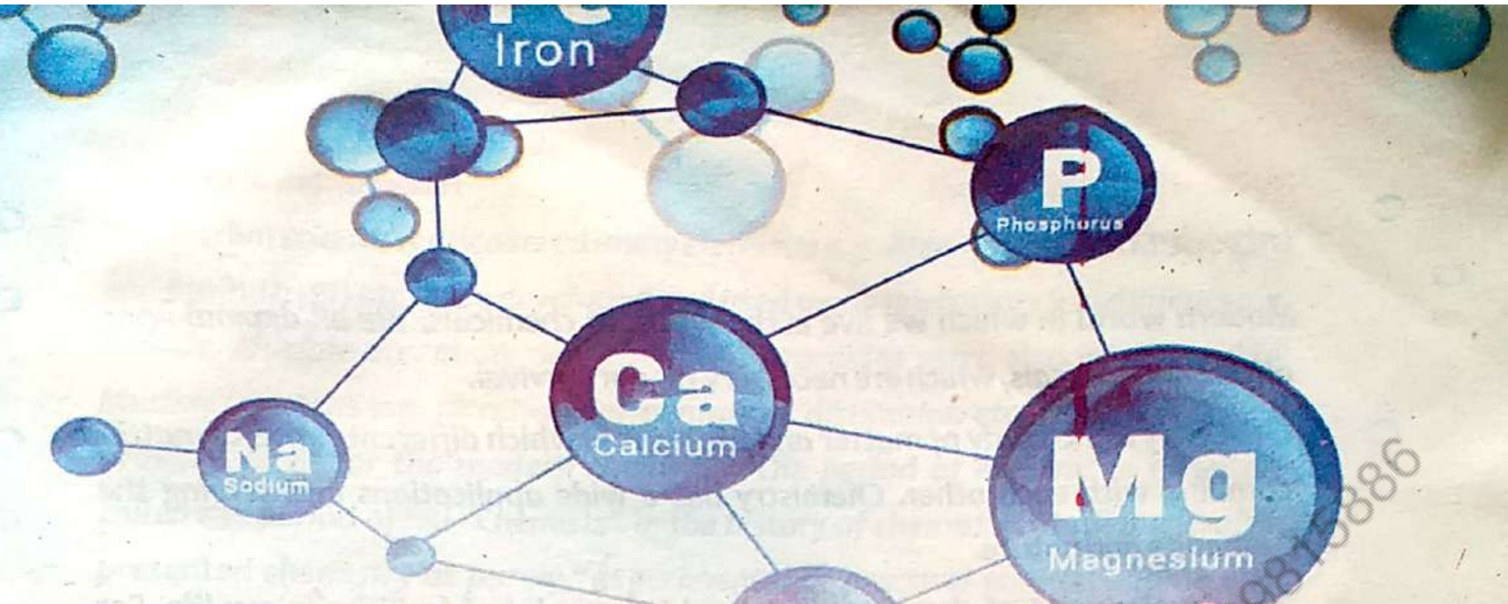


Unit 1 Fundamentals of Chemistry

After studying this unit, the students will be able to;

- *Identify and provide examples of different branches of chemistry.*
- *Differentiate between branches of chemistry.*
- *Distinguish between matter and a substance.*
- *Define ions, molecular ions, formula units and free radicals.*
- *Define atomic number, atomic mass and atomic mass unit.*
- *Differentiate among elements, compounds and mixtures.*
- *Define relative atomic mass based on C-12 scale.*
- *Differentiate between empirical and molecular formula.*
- *Distinguish between atoms and ions.*
- *Differentiate between molecules and molecular ions.*
- *Distinguish between ion and free radical.*
- *Classify the chemical species from given examples.*
- *Identify the representative particles of elements and compounds.*
- *Relate gram atomic mass, gram molecular mass and gram formula mass to mole.*
- *Describe how Avogadro's number is related to a mole of any substance.*
- *Distinguish among the terms gram atomic mass, gram molecular mass and gram formula mass.*
- *Change atomic mass, molecular mass and formula mass into gram atomic mass, gram molecular mass and gram formula mass.*

Introduction

Modern world in which we live is the world of chemicals. We all depend upon different chemicals, which are necessary for our survival.

Chemistry is the study of matter and the ways in which different forms of matter combine with each other. Chemistry has a wide applications and serving the humanity and nature.

The development of chemistry has provided us a lot of facilities in our life, For example, Petrochemicals, medicines and drugs, papers, plastics, paints, Color pigments, soap, detergents etc. The study of chemistry also provides knowledge and techniques to improve our health and environment and to explore and conserve the natural resources.

In previous grades, you have learned about the matter, element, compound, mixtures etc. In this unit, you will study, different branches of chemistry, basic definitions and concepts of chemistry.

Muslim Period 600 – 1600 AD

The science of chemistry grew and flourished in early civilization of the world. The Egyptians, the Greeks, the Romans and the Muslims contributed much to the science of chemistry. The Muslims made a rich contribution to the knowledge in chemistry. They spread like shining stars on the horizon of the world of science. They made effective and invaluable services in the field of chemistry. This period of Muslims is almost 1000 years long (600 – 1600AD).

The principal goals of the Muslim chemists were,

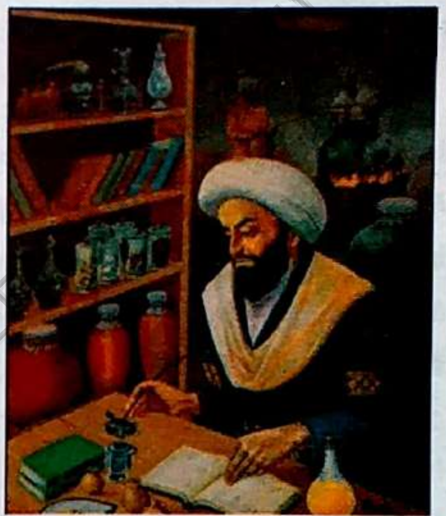
1. to find out methods to prolong life,
2. looking for ways to change base metals (such as Lead) into Gold,
3. to find physical evidences to support religious and philosophical beliefs.

Though they failed in doing so but they laid the foundation of the laboratory methods. These methods are still used in the modern chemistry.

The Muslim scientists discovered many elements e.g. Arsenic (As), Antimony (Sb) and Bismuth (Bi) etc. They developed and used many laboratory instruments e.g. funnels, crucible etc. Many new chemicals processes were also introduced by Muslim scientists e.g. filtration, fermentation, distillation etc. Thus, this period provides basis for the modern chemistry. The period of Muslims is generally called the period of "Al - Chemists" in the history of chemistry. Muslim scientists presented chemistry as purely "experimental or practical science". Some well known Muslim scientists and their achievements are mentioned here.

i) Jabber Ibn Haiyan (721-803 A.D)

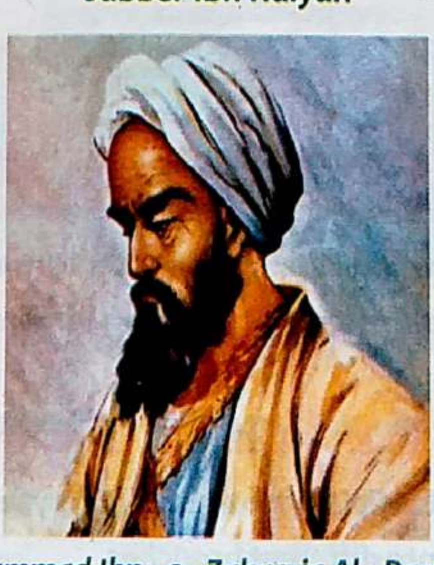
Jabber Ibn Haiyan is generally known as the father of the chemistry. He was probably the first scientist who had a well established laboratory. He invented experimental methods such as distillation , sublimation, filtration, extraction of metals etc. He prepared Hydrochloric acid, Nitric acid and white lead.



Jabber Ibn Haiyan

ii) Muhammad Ibn–e–Zakaryia Al–Razi (864-930)

Al –Razi was a physician, chemist and a philosopher. He wrote 26 books but the most famous book was "Al-Asrar". In this book, he discussed the different processes of Chemistry. He was the first chemist to divide the chemical compounds into four types and also divides the substances into living and non-living origin. He prepared alcohol by fermentation.



Muhammad Ibn –e– Zakaryia Al –Razi

iii) Al – Beruni (973 – 1048 A.D)

Al – Beruni

He made greater contribution in the field of Physics, Meta-physics, Mathematics, Geography, astronomy and History. In the field of Chemistry, he understood the different chemical procedures and chemical combinations. He determined the densities of different substances.

iv) Ibne – Sina (980 – 1037 A.D)

Ibne – Sina

Ibne – sina is generally known as the Aristotle of the Muslim's world. He is famous for his contributions in the fields of Medicine, Mathematics, Astronomy, Medicinal chemistry and philosophy. He is the first scientist who rejected the idea, that base metals can be converted into Gold. He wrote more than hundred books. These books were taught in the Europe for centuries.

Chemistry

Chemistry is a branch of physical science; it includes the study of materials and substances of the world in which we live. Chemistry deals with the composition, structure and properties of matter, the changes occurring in matter and the laws and principles under which these changes occur.

1.1 BRANCHES OF CHEMISTRY

The field of chemistry is very vast. To cover all the areas of chemistry, it is divided into different branches.

BRANCHES OF CHEMISTRY

Physical Chemistry

Concerned with the relationship between the physical properties of substances along with the chemical changes in them.

Organic Chemistry

Study of those compounds which contains Carbon and hydrogen called hydrocarbons and their derivatives.

Inorganic Chemistry

Deals with the study of all the elements and compounds, except hydrocarbons (organic compounds) and their derivatives.

Analytical Chemistry

Deals with the study of qualitative (type / kind) and quantitative (amount / nature) analysis of matter.

Industrial Chemistry

Deals with the study of techniques and chemical processes for the preparation of different industrial products e.g. cement, glass, fertilizers etc.

Nuclear Chemistry

Concerned with the study of nucleus, changes occurring in the nucleus, properties of the particles present in the nucleus and the emission or absorption of radiation from the nucleus.

Biochemistry

Concerned with the study of synthesis, composition, decomposition and chemical reaction of substances, which take place in the living organisms.

Environmental Chemistry

Concerned with the interaction of chemical substances / processes with the environment and their impact on it.

Test Yourself.

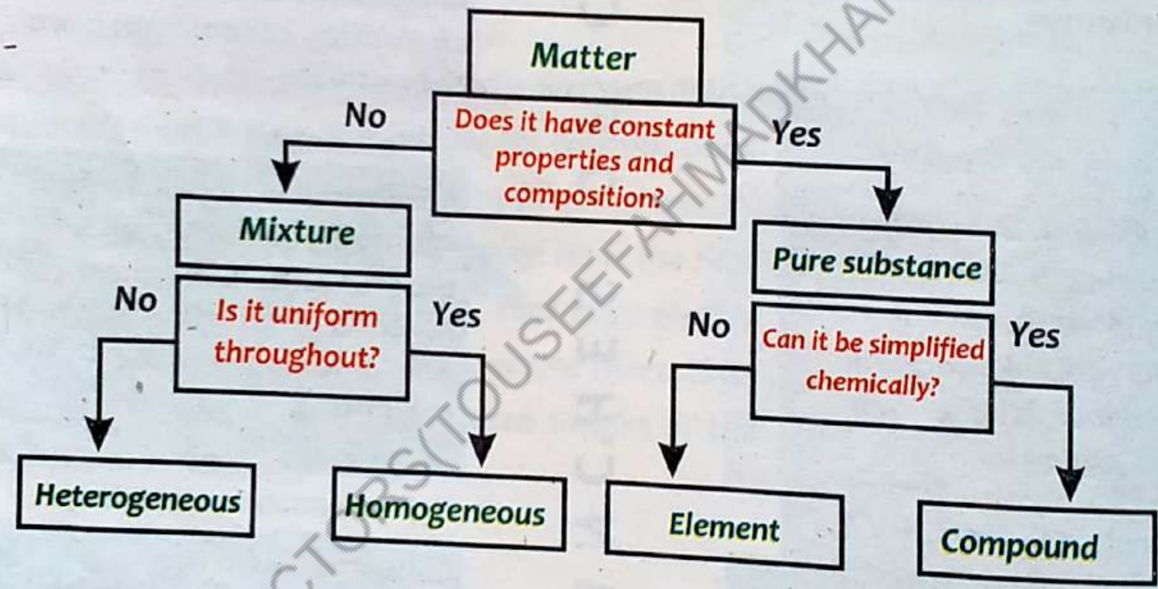
(a) A chemist mostly use which branch of chemistry to determine the amount of an impurity in drinking water?

(b) A chemist in which branch of chemistry determines the composition of substance and determine their masses?

(c) A chemist develops a medicine to help patients with thyroid problems. In which branch of chemistry is this chemist working?

1.2 BASIC DEFINITIONS

All the things present around us, in the universe, are made up of matter. Matter is defined as anything that has mass and occupies space. The quantity of matter in an object is its mass. We can see that almost everything in the universe is made up of matter because they have masses and occupy volumes. A few examples of matter are the food we eat, the water we drink and the air we breathe. Man himself is matter. A piece of matter in pure form is called substance. Pure water is an example of substance, regardless of its sources, have exactly the same composition and properties. All materials may be classified into pure substances (elements, compounds) and impure substances (mixtures).



Society, Technology and Science

In 1924 de Broglie put forward the theory of dual nature of matter i.e. matter has both the properties of particles as well as waves. He explained the background of two ideas. He advocated that these two systems could not remain detached from each other. By mathematical evidences, he proved that every moving object is attached with waves and every wave has corpuscular nature as well. It formulated a basis to understand corpuscular nature of matter.

1.2.1 Elements, Compounds and Mixtures

(i) Elements

An Element is a pure substance which cannot be split up into other simpler substances by any ordinary physical or chemical processes. For example, Gold is an element and it can never be converted into any other simpler substance by ordinary physical or chemical reactions. The fundamental unit of an element is the atom. Scientists have discovered approximately 118 elements, out of which 92 are naturally occurring elements, while the rest have been artificially prepared in the laboratory.

Elements are made up of particles. An atom is the smallest particle of an element that can take part in a chemical reaction. Each element consists of only one kind of atoms. Atoms of a particular element are same to each other but different from the atoms of other elements.

Symbols of Elements

In 1814, Berzelius suggested the system for representing elements with symbols. The shortest name of an element is called symbol. In most cases, the first letter of the name of an element is taken in capital letter as the symbol. In some cases, where the first letter has already been used, then the initial letters in capital together with a small any other letter from its name are used. A list of names and symbols of some important elements is given in the table 1.1.

Table 1.1: Names and Symbols of Some Common Elements

S. No.	Name of element	Symbol
1.	Boron	B
2.	Carbon	C
3.	Fluorine	F
4.	Hydrogen	H
5.	Oxygen	O
6.	Nitrogen	N
7.	Sulphur	S
8.	Aluminum	Al
9.	Calcium	Ca
10	Magnesium	Mg

3rd 3rd home work

Some element's symbol starts with their Latin or Greek language name. A list of names and symbols are given in table 1.2.

Table 1.2 Elements Whose Names and Symbols Begin with Different letters

home work

S. No.	Name of element		Symbol
1.	Antimony	Stibium	Sb
2.	Copper	Cuprum	Cu
3.	Gold	Aurum	Au
4.	Iron	Ferrum	Fe
5.	Lead	Plumbum	Pb
6.	Mercury	Hydragyrum	Hg
7.	Potassium	Kalium	K
8.	Silver	Argentum	Ag
9.	Sodium	Natrium	Na
10	Tin	Stannum	Sn
11	Tungsten	wolfram	W

Activity:1.1

Make a list of Ten elements other than the above one, with their names and symbols. Also write down uses of five elements in daily life.

Society, Technology and Science

Elements are the building blocks of all the substances that make up all living and non-living things. A careful observation of the physical world reveals that matter usually occurs as mixtures. Most of the components of these mixtures are elements and compounds that exist as molecules. Air, water, clay, petroleum, coal, rock, earth etc. are mixtures of numerous compounds. Living things contain thousands of different substances such as carbohydrates, proteins, fats, lipids, DNA, RNA etc. All these substances are molecular in nature.

Scientific Information

Gallium symbol Ga, atomic number of 31 and atomic mass of 69.723. It is silvery soft metal. In nature, gallium is never found in free state. Gallium has some very unique properties. For example, although it is a solid at room temperature, it is still so soft that it can be cut with a knife. It has melting point of 29.76 °C. So a lump of gallium, it would melt from the warmth of your hand.

(ii) Compound

Compound is a pure substance consist of two or more different types of elements chemically combined together in a fixed ratio by mass. In the formation of a compound, the component elements must undergo chemical changes. The properties of the compound are different from the elements from which it is formed. For example water (H_2O) is a compound. It is made up of two elements hydrogen and oxygen. The ratio of hydrogen atoms to oxygen atom in water is

Fig: 1.1 Examples of compounds

always 2:1. Changing this ratio will give a different compound. For example, adding one more oxygen atom gives a ratio of 2:2 and the resulting compound will be hydrogen peroxide (H_2O_2).

Formula of Compound

A compound is always represented by a chemical formula. A formula shows the symbols of the elements of which compound is made and their combining ratio to each other.

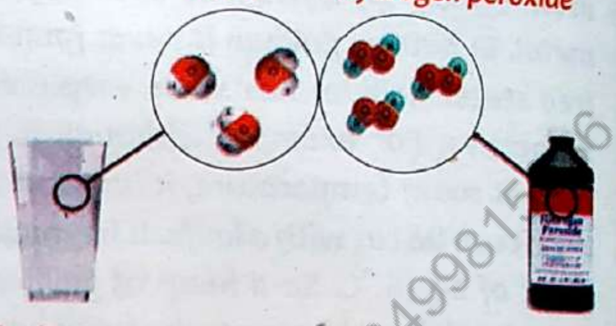


Fig: 1.2 Compounds of H_2O and H_2O_2

Table 1.3: Some common compounds their chemical names and formulae

S.No.	Common Name	Chemical Name	Chemical Formula
1.	Baking soda	Sodium bicarbonate	$NaHCO_3$
2.	Bleaching powder	Calcium hypochlorite	$CaOCl_2$
3.	Caustic soda	Sodium hydroxide	$NaOH$
4.	Chile salt petre	Sodium nitrate	$NaNO_3$
5.	Lime stone	Calcium carbonate	$CaCO_3$
6.	Lime water / milk of lime / slaked lime	Calcium hydroxide	$Ca(OH)_2$
7.	Lime/ Quick Lime	Calcium oxide	CaO
8.	Magnesia	Magnesium oxide	MgO
9.	Marsh gas	Methane	CH_4
10.	Table salt	Sodium chloride	$NaCl$

(iii) Mixtures

A mixture is made up of two or more substances that are not chemically combined. A mixture may consist of elements, compounds or both.

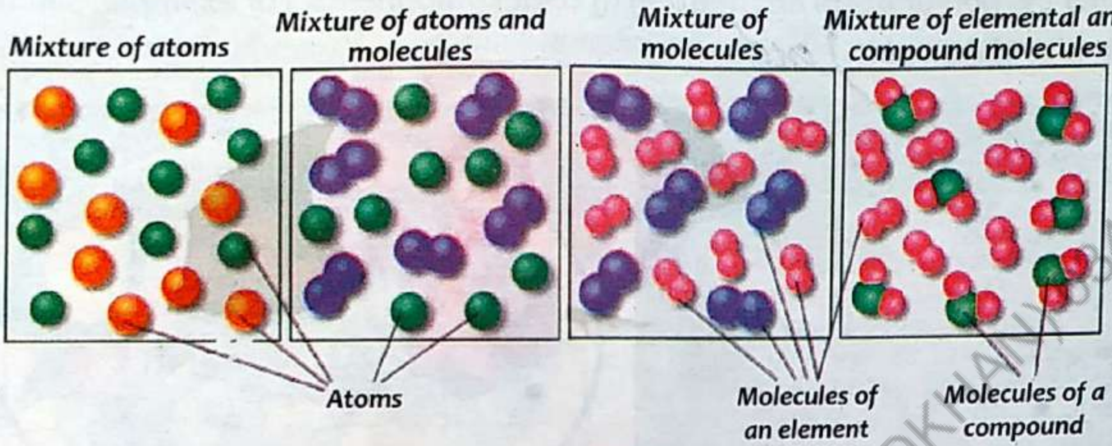


Fig: 1.3 Mixtures

The components of a mixture are not in a fixed ratio. Mixture can be separated by physical methods and constituents substances of the mixture retain their characteristic properties in the mixture. For example, when sugar is dissolved in water, it forms a mixture which is sweet in taste. The sweetness is due to the presence of sugar in water. A widely used example of mixture is alloys, which consist of a mixture of metals with other elements (metals or non- metals) e.g. brass and bronze etc.

Types of Mixtures

There are two types of mixtures.

(i) Homogenous Mixture

A homogenous mixture is one which has uniform composition throughout its mass. For example, Table salt ($NaCl$) dissolved in water (solution).

Fig: 1.4 Homogenous Mixture

(ii) Heterogeneous Mixture

A heterogeneous mixture is one, which does not have uniform composition throughout its mass. One part of the mixture having different composition than the other i.e. non-uniform distribution of their components. For example, Salad, dirt, sand and water etc.) end.

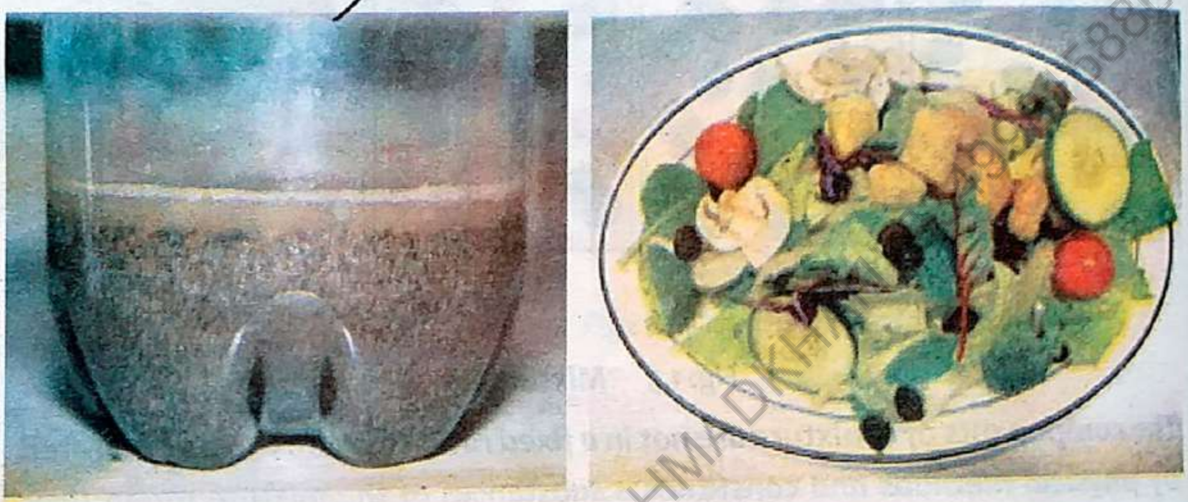


Fig: 1.5 Heterogenous mixture

EXAMPLE 1.1:

If we stir a teaspoon of sugar into a test tube of water and a teaspoon of sand into another test tube of water, the sugar will disappear into the water (dissolve), but the sand will not. Which mixture is a solution?

water and sugar

Solution

The sugar forms a homogeneous mixture with the water (solution). The sand and water form a heterogeneous mixture. Particles of sand are visible in the mixture, whereas, sugar particles are invisible (even with a microscope).

water and sand

Q14 Differentiate between compound and mixture?

Practice Problem: 1.1

Making mixtures

- *Make mixtures of sand, baking soda, ink and oil in water.*
- *Classify these as heterogeneous or homogeneous.*
- *Give reasons for your choice.*

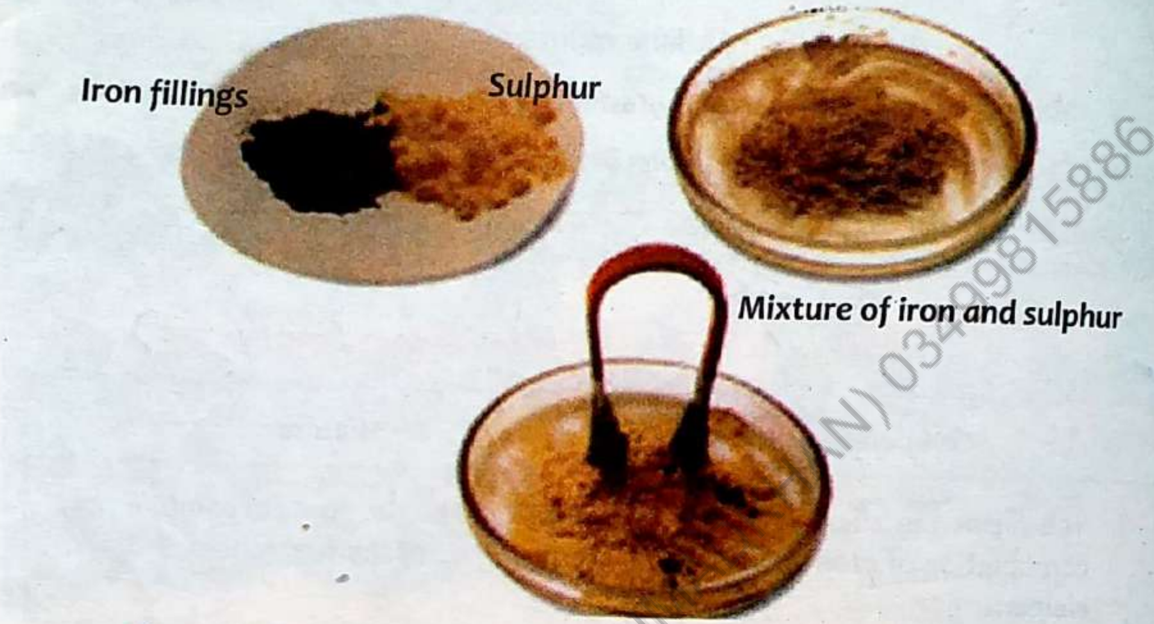
ans.

Table 1.4 (Differences between Compound and Mixture)

for Example

Compound	Mixture
i. It is formed by chemical combination of atoms of the elements.	*It is formed by the physical combination on mixing up of the substances.*
ii. The constituents lose their original properties.	*The constituents retain their properties.*
iii. Compound always has fixed composition by mass.	*Mixture does not have fixed composition by mass.*
iv. The components of the compound cannot be separated by physical methods.	*The components of the mixture can be separated by physical methods.*
v. Every compound is represented by its chemical formula.	*It consists of two or more components and does not have any chemical formula.*
vi. Compound has homogenous composition.	*Mixture may be homogenous or heterogeneous in composition.*
vii. Compound has sharp and fixed melting point.	*Mixture does not have sharp and fixed melting points.*

Iron (Fe) and sulphur (S) by mixing with each other form a mixture and can be separated with magnet. As Iron and sulphur retain its properties in this mixture.



Iron can be separated from sulphur using magnet

When Iron (Fe) and sulphur (S) is strongly heated, it react with each other and lose their properties and form a compound, iron sulphide (FeS). It cannot be separated with magnet.

Strongly heating Iron (Fe) and Sulphur (S), a chemical reaction occurs, a compound iron sulphide (FeS), formed.

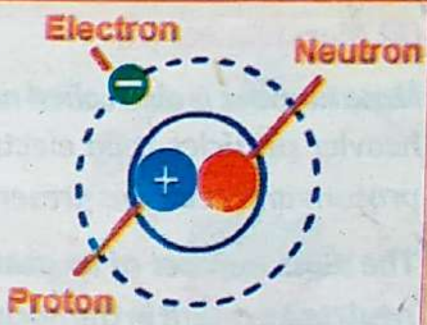
Iron sulphide, FeS dark black solid, not attracted by magnet

Fig: 1.6 Iron (Fe) and Sulphur (S) mixture and compound Iron sulphide (FeS)

Q15 Difine atomic and mass number with an example

Do you know?

Atoms consist of three fundamental particles i.e. protons, electrons and neutrons. Protons and neutrons are present inside the nucleus, whereas, electrons revolves around the nucleus.



1.2.2 Atomic Number and Mass Number

(i) Atomic Number (Z)

It has been found that the atoms of one element differ from those of other elements by the number of protons in their nuclei. No two elements have the same number of protons. As the atom is electrically neutral, so they have the same number of electrons, as the number of protons. The number of neutrons in the atom of an element is different and cannot be used to characterize the atom. Elements are therefore arranged in periodic table on the basis of number of protons in their nucleus, this number of protons is called the atomic number. It is represented by the symbol Z.

Atomic Number (Z) = Number of proton(s) or Number of electron(s) in an atom.

Ans **The atomic number of an element is the number of protons present in the nucleus of one atom of that element.**

Hydrogen, carbon and oxygen contain one, six and eight protons in their nucleus respectively. Thus, the atomic number of hydrogen, carbon and oxygen are one, six and eight respectively.

Table 1.5 Some elements with their atomic number

For Exampl.

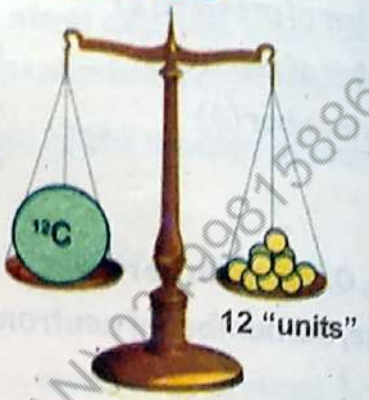
S.No.	Element	Atomic number (Z)
1.	Hydrogen (H)	1
2.	Carbon (C)	6
3.	Oxygen (O)	8
4.	Sodium (Na)	11
5.	Chlorine (Cl)	17

Q16 Define

1.2.3 Relative Atomic Mass and Atomic Mass Unit ?

If we measure the mass of an atom in grams or kilograms, it will be an extremely small value (10^{-24} to 10^{-22} g). Measuring such a small masses are not only impossible but are also impractical to work with. For that purpose, atomic masses are expressed by comparing with a mass of standard atom. The atom chosen nowadays, as the standard for comparison is that of carbon - 12, which has a mass of exactly 12 units.

masses of atoms are expressed relative to the mass of a Carbon-12 atom.



1 atom of C-12 = 12 units

1/12 the mass of C-12 = 1

Ans **The relative atomic mass of an element is the mass of an atom of an element relative to the mass of $\frac{1}{12}$ the mass of C – 12.**

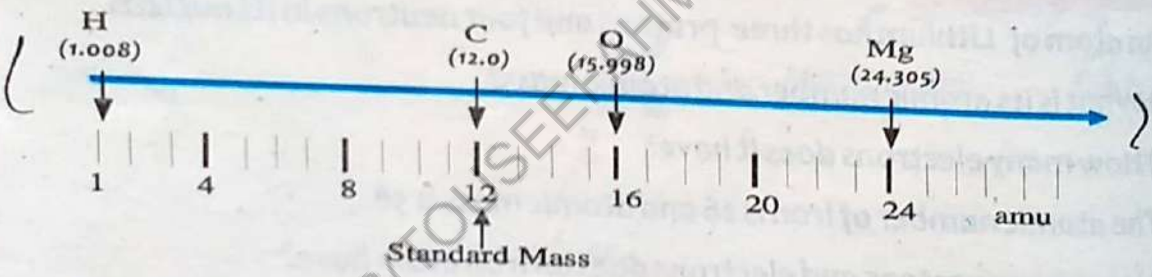


Fig: 1.7 Atomic scale showing relative atomic masses of some elements and that of the standard element C-12.

(I) Atomic Mass Unit (a.m.u.)

Atomic mass unit is the mass of the $\frac{1}{12}$ the mass of an atom of C – 12.

Atomic weight of C – 12 = 12 g = 1 mol

Similarly, 1 mol of C–12 contain atoms = 6.023×10^{23} atoms (Avogadro's number)

$$1 \text{ atom of C} - 12 = \frac{12}{6.022 \times 10^{23}} = 1.993 \times 10^{-23} \text{g}$$

So, we can calculate

$$1 \text{ a.m.u} = \frac{1}{12} \times 1.99 \times 10^{-23} = 1.67 \times 10^{-24} \text{g} = 1.67 \times 10^{-27} \text{kg}$$

for example

~~Table 1.6~~ Some Elements with their Relative Atomic Masses

S. No.	Element	Relative atomic mass in a.m.u.
1.	Hydrogen	1.0079 a.m.u.
2.	Oxygen	15.8994 a.m.u. End
3.	Sodium	22.9897 a.m.u.
4.	Magnesium	24.3050 a.m.u.
5.	Silver	107.8082 a.m.u.

(ii) Average Atomic Mass

The atomic masses, as can be seen in the table 1.6, are rarely found to be exactly whole numbers. This is because most elements are composed of two or more naturally occurring isotopes and the relative atomic mass takes into account the abundance of each isotope.

Average atomic mass is the weighted average of the atomic masses of the naturally occurring isotopes of an element. *Average atomic mass of an element can be calculated by using following formula.*

$$\text{Average atomic mass} = \frac{\text{Atomic mass of } 1^{st} \text{ Isotope} \times \text{its \%age abundance}}{100} + \frac{\text{Atomic mass of } 2^{nd} \text{ Isotope} \times \text{its \%age abundance}}{100}$$

For example, the average atomic mass of Chlorine is 35.5 a.m.u. because there are two isotopes of Chlorine. One is Chlorine – 35 and the other is Chlorine – 37. The percentage occurrence in nature of Chlorine –35 is 75% and the percentage of Chlorine –37 is 25%. Therefore, the average atomic mass of Chlorine is,

$$\text{Average atomic mass} = \frac{\text{Atomic mass of Isotope Cl}-35 \times \text{its \%age abundance}}{100} + \frac{\text{Atomic mass of Isotope Cl}-37 \times \text{its \%age abundance}}{100}$$

Average atomic mass of Cl $= (0.75)(35\ a.m.u.) + (0.25)(37\ a.m.u.)$

$= 26.25 + 9.25$

$= 35.50\ a.m.u.$

Practice Problem 1.2:

- Naturally occurring gallium consists of 60.108% Ga – 69, with a mass of 68.9256 amu, and 39.892% Ga – 71, with a mass of 70.9247 amu. Calculate the average atomic mass of gallium.
- Calculate the atomic mass of neon from the given data.

$^{20}_{10}Ne = 90.51\%$ $^{21}_{10}Ne = 0.27\%$ $^{22}_{10}Ne = 9.22\%$

1.2.4 Chemical Formula

As you have learnt earlier that compound are represented by a chemical formula. The chemical formula of a compound is written by putting together the chemical symbols of the elements that make up the compound and their relative ratio by atoms.

It is an abbreviation used for the full name of a compound with the help of symbols. In other words, we can say that the symbolic representation of a molecule of a compound is called chemical formula.

The chemical formula tells us:

- The type of atoms (elements) present in the compound.
- The ratio of the different atoms present in the compound.

For example, the chemical formula of Sodium Nitrate is $NaNO_3$ and that of glucose is, $C_6H_{12}O_6$.

Following are the two types of chemical formula.

(i) Empirical formula

(ii) Molecular formula

(i) Empirical Formula is called Empirical formula.

The empirical formula of a chemical compound is the simplest formula which shows the smallest whole number ratio of the atoms, of the different elements present in a compound. It may not be the actual formula of the compound. For example, the actual formula of benzene is C_6H_6. ~~Benzene~~ has the simplest ratio of 1:1. So, its empirical formula is CH and the actual formula of glucose is $C_6H_{12}O_6$. Glucose has the simplest ratio of 1:2:1. So, its empirical formula is CH_2O.

(ii) Molecular Formula

Molecular formula shows the actual numbers of atoms of different elements present in one molecule of that compound. For example,

Molecular formula of Glucose = $C_6H_{12}O_6$

Molecular formula of Benzene = C_6H_6 End

Molecular formula is derived from empirical formula by the following relationship.

Molecular formula = n (Empirical Formula)

Where n = 1, 2, 3, 4…… and so on

For example the molecular formula of benzene is,

Molecular formula = n (Empirical Formula)

Where the value of 'n' in this case is 6,

Molecular formula = 6 (CH)

Molecular formula = C_6H_6.

A molecular formula may be the same or multiple of the empirical formula. A few compounds showing their empirical and molecular formula are shown in table 1.7.

Table 1.7 Some compounds with empirical and molecular formulae

S. No.	Name of Compound	Empirical Formula	Molecular Formula
1	Acetic acid	CH_2O	CH_3COOH ($C_2H_4O_2$)
2	Acetylene	CH	C_2H_2
3	Benzene	CH	C_6H_6
4	Glucose	CH_2O	$C_6H_{12}O_6$
5	Hydrogen Peroxide	HO	H_2O_2
6	Sulphuric Acid	H_2SO_4	H_2SO_4
7	Iron Oxide	Fe_2O_3	Fe_2O_3
8	Ammonia	NH_3	NH_3
9	Methane	CH_4	CH_4
10	Water	H_2O	H_2O

Q 18 Define formula unit with examples.

It is clear from the table 1.7 that some compounds have identical empirical and molecular formula.

(iii) Formula unit

All the ionic compounds are represented by their formula unit. **Formula unit is the smallest repeating unit of an ionic compound showing the simple ratio between the ions.**

For example, in the whole crystal lattice of sodium chloride, the ratio of Na^+ and Cl^- ion is 1:1 and its formula unit is NaCl, similarly KCl is the formula unit of potassium chloride.

End

Practice Problem: 1.3

How many atoms of each element are present in one formula unit of each of the following compounds?

(a) Hg_2Cl_2 (b) $NH_4H_2PO_4$ (c) $MgCl_2$ (d) NH_4OH

Q19 Define molecular and formula mass?

1.2.5 Molecular Mass and Formula Mass

(19) ml

(i) Molecular mass

Molecular mass is the sum of the relative atomic masses of all the atoms present in the molecule. Molecular mass of a substance can be calculated by adding relative atomic masses of all the atoms in one molecule of the substance. For example,

Molecular formula of Ethane = C_2H_6

Atomic mass of carbon = 12

Atomic mass of Hydrogen = 1

Molecular mass of Ethane = 2(Atomic Mass of C) + 6(Atomic Mass of H)

= 2 (12) + 6(1)

= 24 + 6

Thus, molecular mass of Ethane = 30 a.m.u.

(ii) Formula Mass

Formula mass is the sum of the atomic masses of ions (atom) present in the formula unit of an ionic compound. The formula mass is calculated in the same manner as we did for molecular mass. Thus for NaCl,

Formula mass of NaCl = 1(atomic mass of Na) + 1(Atomic mass of Cl)

= 1(23 amu) + 1(35.5 amu)

= 23 amu + 35.5 amu

Formula mass of NaCl = 58.5 amu. End.

EXAMPLE: 1.2

Calculate the formula mass of (a) $(NH_4)_2SO_3$ and (b) $Fe(NO_3)_2$

Solution

(a) $(NH_4)_2SO_3$

2 N 2 x 14 amu = 28 amu

8 H 8 x 1 amu = 8 amu

1 S 1 x 32 amu = 32amu

3 O 3 x 16 amu = 48 amu

Formula mass = 116 amu

(b) $Fe(NO_3)_2$

1 Fe 1 x 56 amu = 56 amu

2 N 2 x 14 amu = 28 amu

6 O 6 x 16 amu = 96 amu

Formula mass = 180 amu

Q2: Define Ion and its types with examples.

Practice Problem: 1.4

Calculate the formula mass of (a) KCl and (b) Na_2CO_3.

1.3 Chemical Species

Chemical species that is atoms, molecules, ions, etc. take part in a chemical reaction.

Ans. ### (i) Ion

An atom is a neutral particle, the number of electrons and protons are equal in an atom. An atom can lose or gain electron and the balance between number of protons and electrons is disturbed. So there is net charge on the atom. **The particle that carries an electrical charge, positive or negative due to the loss or gain of one or more electrons is called an ion.**

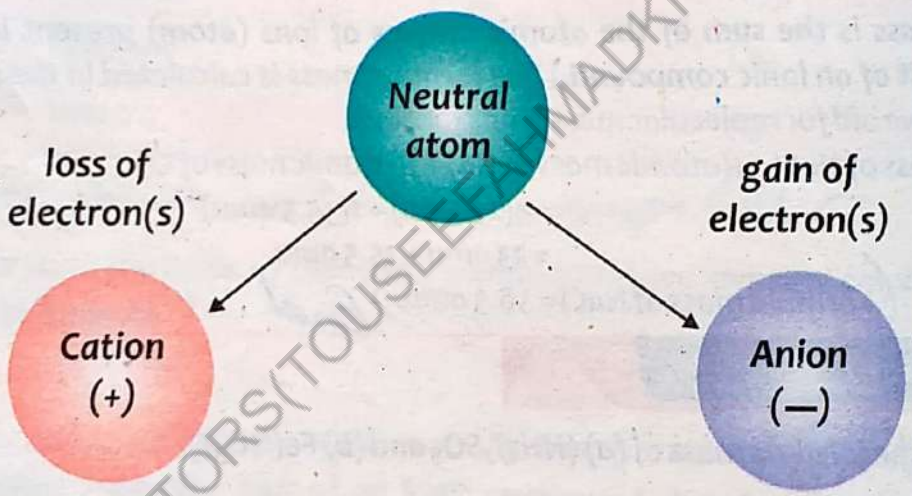


Fig: 1.9 Types of Ions

There are two types of ions, a) Cation (Positive ions) and b) Anion (Negative ions).

a) Cation

The atom which loses electron(s) from its outer most shell, a net positive charge is appeared on that atom. This positively charged species is called cation. Positive ions always have less number of electrons than the number of proton. The following equations show the formation of cations from different atoms.

for, examples

Atoms		Cations
H	$\longrightarrow$	H^+ + $1e^-$
Na	$\longrightarrow$	Na^+ + $1e^-$
Mg	$\longrightarrow$	Mg^{2+} + $2e^-$

b) Anion

The atom which gains electron(s), a net negative charge is appeared on the atom. This negatively charged species is called an anion. Negative ions always have more number of electrons than the number of protons. The following equations show the formation of anion from different atoms.

For Examples

Atoms		Anions
F + $1e^-$	$\longrightarrow$	F^{-1}
Cl + $1e^-$	$\longrightarrow$	Cl^{-1}
O + $2e^-$	$\longrightarrow$	O^{-2}

Q#: Differentiate between Atom and ions?

~~Table 1.8:~~ Difference between an atom and ion

Atom	Ion
i. It is neutral. It has same number of protons and electrons.	It has a net charge (either negative or positive) on it. The number of protons is different than electrons.
ii. It is the smallest particle of an element.	It is the smallest unit of an ionic compound.
iii. It can or cannot exist independently, for example Na, K, Fe.	It cannot exist independently, for example Na^+, K^+, Fe^{+2}.

Q[22] Define molecular ion with example?

(ii) Molecular ion

Just like atoms molecules also are electrically neutral particles. **When a molecule loses or gains one or more electrons, it forms a molecular ion. Molecular ions are also called radical.** Molecular ion can be cationic molecular ion (when they carry positive charge) or anionic molecular ion (when they carry negative charge). The examples of cationic molecular ions are CH_4^{+1}, CO^{+1} etc. while the examples of anionic molecular ions are $C_2H_5OH^{-1}$ etc. Cationic molecular ions are more common than anionic molecular ions.

Q[23] Define

(iii) Free Radical

with example?

Free radicals are atoms or group of atoms that has a single (unpaired) electron (odd) in an outer shell with no charge. It is represented by putting a dot over the species e.g. $\dot{H}$, $\dot{Cl}$, $H_3\dot{C}$ etc. Free radicals are formed by the homolytic fission (Equal breakage) of the bond between two atoms. For example, the chlorine molecule first forms chlorine free radical.

$$Cl_2 \rightarrow 2Cl^{\cdot} \quad \text{Chlorine free radical}$$

The chlorine free radical ($Cl^{\cdot}$) react with CH_4 to form methyl radical $\dot{CH}_3$.

$$CH_4 + Cl^{\cdot} \rightarrow \underset{\text{Methyl free radical}}{\dot{CH}_3} + HCl$$

A free radical is reactive specie which does not exist independently. The $\dot{CH}_3$ is very reactive and react with another Cl_2 molecule to form chloromethane and $Cl^{\cdot}$ again.

$$\dot{CH}_3 + Cl_2 \rightarrow \underset{\text{chloromethane}}{CH_3-Cl} + Cl^{\cdot}$$

Q[24] Define molecule and Explain its types with example?

1.3.2 Molecule

Ans **A molecule is formed by the chemical combination of atoms. It is the smallest unit of a substance. Molecule is the smallest particle of matter which can exist free in nature and may be composed of like or unlike atoms. It shows all the properties of that particular substance. For example, H_2, N_2, O_2, H_2O, etc.** There are different types of molecules depending on the number and types of atoms combining. Some types of molecule are following.

End

Molecules

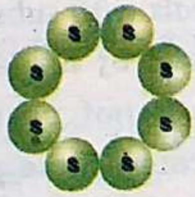
Mono-Atomic Molecule (mono = one)
Made up of only one atom
For example, the inert gases He, Ne, Ar etc.

Di-Atomic Molecule (di = two)
Consists of two atoms
For example, Hydrogen (H_2), Oxygen (O_2), Nitrogen (N_2), Hydrochloric acid (HCl) etc

Oxygen (O_2)

Tri-Atomic Molecule (tri = three)
Consists of three atoms
For example, water (H_2O), carbon dioxide (CO_2), Ozone O_3 etc.

Poly-Atomic Molecule (Poly= Many)
Made up of many atoms
For example, sulphuric acid (H_2SO_4), sugar ($C_{12}H_{22}O_{11}$), sulphur molecule S_8 etc



Sulphur S_8

~~Fig: 1.10~~ Types of molecules

It is noted that some molecules are very big and large and they are called macromolecules (macro means very large). For example, haemoglobin is a macromolecule found in blood.

Q25 Define molecule? Explain classification of molecules on the basis of nature of atoms.

On the basis of the nature of atoms present in the molecules, they may be classified in two groups i.e. homoatomic and heteroatomic molecules.

~~Table 1.9~~: Comparison of homo-atomic molecule and Hetero-atomic molecule

Homo-Atomic Molecule (Homo= Same)	Hetero-Atomic Molecule (Hetero= Different)
When a molecule consists of atoms of same element, it is called homoatomic molecule. They are also called homo-nuclear molecule. For example, oxygen (O_2), hydrogen (H_2), chlorine (Cl_2), sulphur (S_8) etc.	When a molecule consists of atoms of the different elements, it is called heteroatomic molecule. They are also called heteronuclear molecule. For example, carbon dioxide (CO_2), water (H_2O), Hydrochloric acid (HCl), etc.

1.4 Avogadro's number and Mole

and mole.

Q26 what do you means by Avogadro numbers

1.4.1 Avogadro's Number

A chemist deals with substances which are composed of large number of particles. It is not possible for the chemist to count these particles. The concept of Avogadro's number helped in the counting of these particles present in the given mass of a substance. Avogadro's number is collection of 6.023×10^{23} particles. **Avogadro's number is defined as the number of particles (atoms, molecules, ions) in one mole of any substance and is numerically equal to 6.023×10^{23}.**

Avogadro

Scientific Information

The molecule made up of more than one atom are termed as polyatomic molecules e.g.
diatomic = O_2
tri atomic = CO_2
tetra atomic = NH_3
Penta atomic = CH_4

This number was determined by an Italian scientist Amedo Avogadro and is called Avogadro's number represented by symbol "N_A".

Scientists are agreed that Avogadro's number of particles is present in one molar mass of a substance. In order to understand the relationship between the Avogadro's number and the mole of a substance let consider a few examples.

For Examples
i. 1.0g of hydrogen = 1 mole = 6.023×10^{23} atoms of hydrogen
ii. 12 g of carbon = 1 mole = 6.023×10^{23} atoms of carbon
iii. 18g of water = 1 mole = 6.023×10^{23} molecules of water
iv. 58.5 g of NaCl = 1 mole = 6.023×10^{23} formula unit of NaCl

Society, Technology and Science

The mole concepts, students often have trouble comprehending the huge size of Avogadro's number, 6.023×10^{23}. The following analogies may be helpful for better understanding of students.

- If there were a mole of rice grains, all the land area in the whole world would be covered with rice to a depth of about 75 meters.
- One mole of rice grains is more than all the grain that has been grown since the beginning of time.
- A computer counting with a speed of 10 million atoms a second, would take two billion years to count one mole of atoms.
- If one mole of marbles were spread over the surface of the Earth, our planet would be covered by a three miles thick layer of marbles.

1.4.2 Mole Ans

We use different units in our daily life to represent different quantities of items so that we can count them easily. How do we count shoes? As shoes come in pairs, so we would most likely count them by pairs rather than individual shoes. Similarly eggs, banana etc are counted in dozens, but paper by ream. Thus, the counting unit depends on what we are counting. Chemists also use a unit for counting atoms, molecules and ions. They use a counting unit called mole to measure the amount of a substance.

A mole is defined as the amount (mass) of a substance which contains Avogadro's number (6.023×10^{23}). Mole is represented as 'mol'. It forms a link between mass of substance and number of particles.

The mass of a substance is may be atomic mass, molecular mass or formula mass. These masses are expressed in atomic mass units (amu). But when these masses are expressed in grams they are called molar masses.

So, the mole can be quantitatively defined as the atomic mass or molecular mass or formula mass of a substance expressed in grams is called mole.

For example,

Atomic mass of carbon = 12	Expressed in grams = 12g	= 1 mole of carbon
Molecular mass of H_2O = 18	Expressed in grams = 18g	= 1 mole of water
Molecular mass of H_2SO_4 = 98	Expressed in grams = 98g	= 1 mole of sulphuric acid
Formaula mass of NaCl = 58.5	Expressed in grams = 58.5g	= 1 mole of sodium chloride

Bud

Q 9.7 Definition

1.4.3 Gram Atomic Mass, Gram Molecular Mass and Gram Formula Mass

As we know that all substances are made up of atoms, molecules or formula units (ions). Their masses are expressed as atomic mass, molecular mass, or formula mass respectively. Beside this, other units can also be used to express these masses. One of them is to express in grams.

(i) Gram Atomic Mass

When the relative atomic mass is expressed in grams, we call it the gram atomic mass. One gram atom of any element is the relative atomic mass of that element expressed in grams. This is also called mole.

Table 1.10 : Elements and their gram atomic mass

1gram atom of hydrogen	1.008g	1 mole of hydrogen
1 gram atom of oxygen	15.998 g	1 mole of oxygen
1 gram atom of carbon	12.00 g	1 mole of carbon

It is clear from the above examples 1 gram atom of different elements have different masses.

(ii) Gram Molecular Mass

The molecular mass of a substance expressed in grams is called gram molecular mass or gram molecule of that substance. It is also called mole.

Table 1.11: Compounds and Their Gram Molecular Mass

Name of compound	Chemical formula	Gram MolecularMasses
Water	H_2O	18g (1 mole)
Carbon dioxide	CO_2	44g (1 mole)
Sulphuric acid	H_2SO_4	98g (1 mole)

(iii) Gram Formula Mass

When the formula mass of an ionic compound is expressed in grams, it is called gram formula mass or gram formula. It is also called mole.

Table 1.12 Ionic compounds and their gram formula mass

Name of ionic compound	Chemical formula	Gram formula mass
Sodium Chloride	NaCl	58.5 g (1 mole)
Calcium chloride	$CaCl_2$	111g (1mole)

1.5 Chemical Calculations

In chemical calculations, we can calculate the number of moles from given mass of a substance and number of particles from a given number of moles of a sample. These calculations are based on the mole into mass and mole into particles conversion concept. The examples of these calculations are:

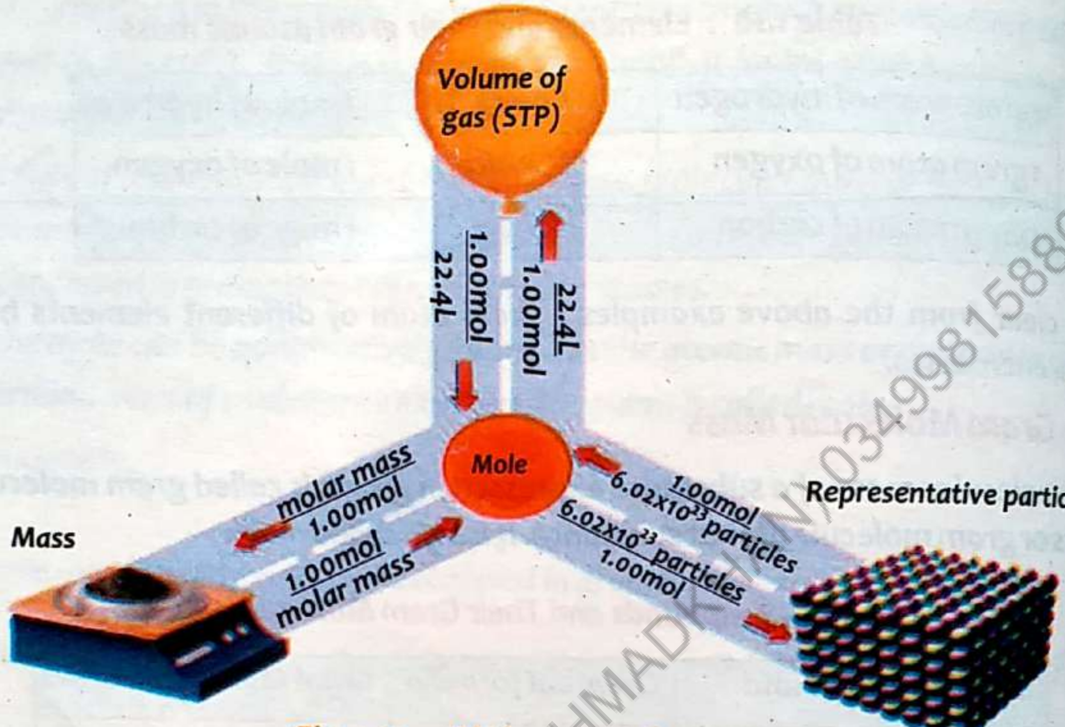


Fig: 1.10 Mole conversion map

Q (28) How will you calculate mole from mass and mass from mole?

1.5.1 Mole – mass calculations

In mole – mass calculations, we can calculate the number of moles of a substance from the known mass of the substance. When a chemical reaction occurs, changes take place. These changes can be expressed in terms of mole and masses.

When we say molar mass of water, we mean 18 grams of water. This relationship is often treated as conversion factor. Thus, 1 mole of water is 18 grams, 1 mole of Al is 27 grams.

We can convert any mass into moles and any number of moles into mass by using the following equation.

$$\text{No. of Mole}(n) = \frac{\text{Mass (in grams) of a Substance}}{\text{Molar mass of the Substance}}$$

By rearranging the above equation, we can calculate the mass of a substance from the number of moles of a given substance.

Mass (in grams) of a substance = No. of mole (n) x molar mass of the substance

EXAMPLE: 1.3

Calculate the moles in 60g of CO_2.

Solution

Mass of CO_2 = 60g

The molar mass of CO_2 = 12.00 g + 2 (16.00 g)

= 12 + 32

= 44 g/mol

$$\text{No. of moles of } CO_2 = \left(\frac{\text{Mass (in grams)}}{\text{Molar mass of } CO_2}\right)$$

$$= \frac{60}{44} = 1.364 \text{ mole}$$

Practice Problem: 1.5

How many moles are there in 90g of water (H_2O)?

How will you calculate mole from particles and particles from mole.

1.5.2 Mole – particle calculations

There is also a relationship between mole of a substance and Avogadro's number (N_A).

1 mole of any substance contain = 6.023×10^{23} particles (atom, molecules, and ions)

Ans

Based on the above relationship, we can convert any number of moles into particles, (atom, molecules and ions) and any number of particles into mole using the following formula.

$$\text{No. of Moles (n)} = \frac{\text{No. of particles}}{\text{Avogadro's number } (N_A)}$$

$$\text{No. of Moles (n)} = \frac{\text{No. of particles}}{6.023 \times 10^{23}}$$

No. of particles = No. of moles x N_A

No. of particles = No. of moles x 6.023×10^{23}

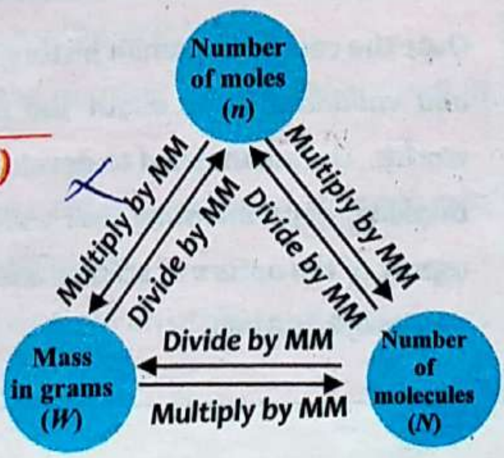


Fig: 1.11 Mole mass conversion

EXAMPLE : 1.4

How many moles of hydrogen are there in 8.9×10^{23} hydrogen atoms?

Solution:

No. of atoms of hydrogen = 8.9×10^{23}

No. of H atoms per mole = $N_A = 6.023 \times 10^{23}$

Therefore the number of moles present, can be calculated using equation.

$$\text{No. of moles of Hydrogen} = \frac{\text{No. of atoms or molecule}}{\text{Avogadro's No } (N_A)}$$

$$\frac{8.9 \times 10^{23}}{6.023 \times 10^{23}} = 1.48 \text{ moles}$$

Practice Problem: 1.6

a. How many moles of CO_2 are there in 3.01×10^{23} molecules of CO_2.

b. Calculate the number of molecules present in 5 moles of glucose, $C_6H_{12}O_6$.

Society, Technology and Science

Over the course of human history, people have developed many interconnected and validated ideas about the physical, biological, psychological and social worlds. The means used to develop these ideas are particular ways of observing, thinking, experimenting and validating. These ways represent a fundamental aspect of the nature of science and reflect how science tends to differ from other modes of knowing.

Key Points

- *Chemistry is the branch of science, which deals with the composition, structure and properties of matter, the changes occurring in matter and the laws and principles under which these changes occur.*
- *The main branches of chemistry are organic, inorganic, physical, analytical, industrial, nuclear, environmental and biochemistry.*
- *Matter is anything that occupies space and has mass.*
- *A piece of matter in pure form is called substance.*
- *Matter can be classified as an element, a compound or a mixture.*
- *An Element is a substance which cannot be split up into other simpler substances by any ordinary physical or chemical processes.*
- *An atom is the smallest particle of an element that can take part in a chemical reaction.*
- *Compound is a pure substance made up of two or more elements chemically combined together in a fixed ratio by mass.*
- *Mixtures can be classified as homogeneous or heterogeneous.*
- *The atomic number (Z) of an element is the number of protons present in the nucleus of an atom of that element.*
- *The mass number (A) of an element is the total number of protons and neutrons in the nucleus of an atom of that element.*
- *Relative atomic mass of an element is the mass of an atom relative to mass of 1/12 the mass of Carbon-12.*

- Atomic mass unit is the mass of the 1/12 of the mass of any atom of Carbon-12.
- Average atomic mass is the weighted average of the atomic masses of the naturally occurring isotopes of an element.
- Chemical Formula is a shorthand method of representing a compound. Elemental symbols are used in writing a formula.
- Empirical formula shows the smallest whole number ratio of the atoms/ions present, while molecular formula shows the actual number of atoms present in a molecule of the compound.
- Molecular mass and formula mass of a compound is sum of the relative atomic masses of the elements (atoms) present in the compound.
- Simple ions are atoms that have lost or gained electrons.
- Molecules can also form ions called molecular ions.
- Free radical is a specie, with an unpaired electron and is the result of homolytic fission of the bond.
- A molecule is the smallest particle of a compound or element that can exist independently. It may be mono-atomic or polyatomic.
- Molecules made up of similar atoms are homo-atomic and those composed of different atoms are hetero-atomic.
- A mole represents a definite quantity (mass) of an element or a compound. The gram atomic mass, gram formula mass or gram molecular mass of an element or compound is called a mole.
- Avogadro's number is a constant value equal to 6.023×10^{23} and this number of atoms or molecules is present in one mole of a substance.

Exercise

Choose the correct option.

i. Which one of the following group comprise of elements:

(a) Mercury, water, Ammonia (b) Iodine, Tin, Iron
(c) Copper, Aluminum, Methane (d) Coal, Smoke, Fog

ii. Which one of the following can be broken down in to simpler substances:

(a) Ammonia (b) Oxygen
(c) Sulphur (d) Iron filling

iii. The gram molecular mass of HNO_3 is:

(a) 60 (b) 100
(c) 63 (d) 98

iv. Which of these molecules is not a compound:

(a) N_2O (b) N_2
(c) NO (d) No_2

v. Which one of the following is equal to two moles of water(H_2O):

(a) 1.084×10^{25} molecules (b) 6.022×10^{23} molecules
(c) 1.204×10^{24} molecules (d) 1.806×10^{24} molecules

vi. A compound contains:

(a) different kinds of atoms mixed together
(b) the same kind of atoms mixed together
(c) different kinds of atoms chemically combined together
(d) the same kind of atoms chemically combined together

vii. Which one of the following is an example of triatomic molecule:

(a) CO_2 (b) O_2
(c) CH_4 (d) NH_3

viii. Hydrogen gas is:

(a) a monoatomic gas
(b) a mixture of hydrogen atoms
(c) a diatomic gas with each molecule made up of two atoms
(d) a diatomic atom made up of two molecules

ix. Which one of the following compounds has both empirical and molecular formula identical?

(a) benzene (C_6H_6) (b) hydrogen peroxide(H_2O_2)
(c) water (H_2O) (d) glucose ($C_6H_{12}O_6$)

x. Which one of the following is a homogeneous mixture

(a) smoke (b) air (c) fog (d) smog

SHORT QUESTIONS

Answer briefly the following questions.

i. How many elements are present in each of the following?

(a) HF and Hf (b) Co and CO
(c) Si and SiO_2 (d) $PoCl_2$ and $POCl_3$

ii. Cm is the chemical symbol for curium, named after the famous scientist Madam Curie. Why wasn't the symbol C, Cu or Cr used instead?

iii. What is the atomic number of an element? How does it differ from the mass number?

iv. Students often mix up the following elements. Give the name for each element.

(a) Mg and Mn (b) K and P (c) Na and S (d) Cu and Co

v. a) Classify the following molecules as mono-atomic, diatomic, tri-atomic and polyatomic molecules.

H_2O, N_2, S_8, He, HCl, CO_2, Ar, H_2SO_4, $C_6H_{12}O_6$

b) Classify the following as cation, anion, molecular ion, free radical and molecule.

CH_4^+, O^{-2}, $CH_3^{\bullet}$, CO^+, CO_2, Cl^-, Mg^{+2}, CO_3^{-2}, O_2, Na^+, $C_2H_5O^{-1}$, H_2O, Cl_2

vi. Calculate the number of moles of butane, C_4H_{10} in 151 g of butane (At. masses C=12 amu and H =1 amu). (Ans: 2.63 moles)

vii. What is the mass of 5 moles of ice? (Atomic masses: H= 1amu, O=16 amu) (Ans: 90g)

viii. Calculate the number of molecules in 6.50 mol of CH_4. (Ans: 3.914×10^{24} moles)

ix. Calculate the average atomic mass of lithium from the following data:

Isotope	Natural Abundance (%)	Relative Atomic Mass (amu)
^6Li	7.5	6.0151
^7Li	92.5	7.0160

(Ans: 6.940)

x. Calculate the mass of 6.68×10^{23} molecules of PCl_3. (Ans: 152.524g)

LONG QUESTIONS

(i) State and explain with examples:

(a) The empirical formula of a compound.

(b) The molecular formula of a compound.

(ii) What do you understand by the terms mole and Avogadro's number. Explain with suitable examples.

(iii) (a) Compare and contrast a mixture and a compound. Give examples of each of them.

(b) How will you classify molecules? Support your answer with at least two examples of each.

(iv) (a) What is the molecular mass of a compound? How will you differentiate it from formula mass?

(b) Calculate the molecular mass or formulae mass, as the case may be of the following compounds in amu.

(i) Benzene, C_6H_6 (ii) Ethane gas, C_2H_6

(iii) Aluminum chloride, $AlCl_3$ (iv) Iron oxide, Fe_2O_3

(v) (a) Find out the number of protons, electrons and neutrons in the following elements.

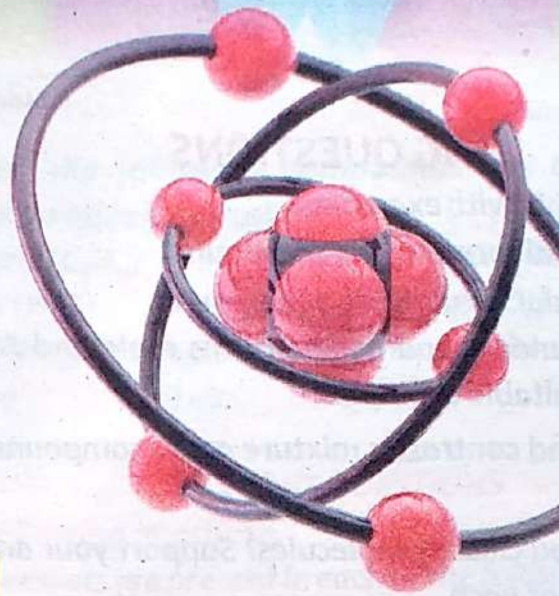
$$^{23}_{11}Na,\ ^{107}_{47}Ag,\ ^{56}_{26}Fe,\ ^{207}_{82}Pb,\ ^{40}_{18}Ar,\ ^{238}_{92}U$$

(b) Complete the following table:

	Symbol	Atomic No.	Number Protons	Number of electrons
a.	K			
b.		8		
c.			15	15
d.		20		
e.	Cl			

Project Work

Prepare the model of sodium, carbon and oxygen, showing number of protons, electrons and neutrons and present in the class.



Unit 2 Structure of Atom

After studying this unit, the students will be able to;

- ***Describe the contributions that Rutherford made to the development of the atomic theory.***
- ***Explain how Bohr's atomic theory differed from it.***
- ***Describe the structure of an atom including the location of the proton, electron and neutron.***
- ***Define isotopes.***
- ***Compare isotopes of an atom.***
- ***Discuss properties of the isotopes of H, C, Cl, U.***
- ***Draw the structure of different isotopes from mass number and atomic number.***
- ***State the importance and uses of isotopes in various fields of life.***
- ***Describe the presence of sub shells in a shell.***
- ***Distinguish between shells and sub shells.***
- ***Write the electronic configurations of the first 18 elements in the Periodic Table***

Quest. write down main postulates of Dalton atomic theory?

Introduction

The composition of matter is always a mystery to scientists. Many philosophers and scientists worked for centuries to solve this mystery.

A Greek philosopher Democritus (460 - 370 B.C.) for the first time suggested that all matter can be ultimately broken down into tiny particles that cannot be further divided. These particles were named atoms (from the Greek word "atomos" means indivisible).

In this unit, you will learn about the development of the atomic model and concepts like isotopes and electronic configuration in detail.

2.1 Theories and experiments related to atomic structure

Although the word atom was used in 400B.C, however no further work was done until 19th century. It was John Dalton, an English School teacher who after a series of experiments concluded that all matter must be composed of tiny particles, which are like solid balls and cannot be further divided. He called them atoms. He presented his theory under the title "A New System of Chemical Philosophy." The Main postulates of Dalton atomic theory (1808) are as follows:

John Dalton

1. **Matter is composed of very small particles called "Atoms".**
2. **Atom is an indivisible particle.**
3. **Atom can neither be created nor destroyed.**
4. **Atoms of particular element are identical in size, shape, mass and also in other properties.**
5. **Atoms of different elements are different in their properties.**
6. **Atoms combine with each other in small whole numbers ratio.**
7. **All chemical reactions are due to combination or separation of atoms.**

End

H.

Ans) Dalton atomic theory

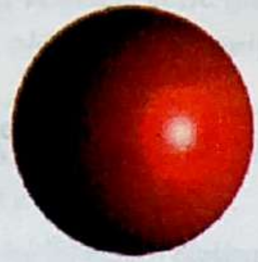


Fig 2.1 Dalton's Model of Atom as solid ball

2.1.1 Rutherford's atomic model

Q(2) In 1911, Rutherford performed an experiment in order to know the arrangement of electrons and protons in an atom. or: How nucleus is discovered?

Explain (i) Rutherford's Experiment?/ Discovery of Nucleus

Rutherford

Rutherford performed an experiment to determine the internal structure of the atom in 1911. For this purpose, he used a thin (0.00004cm) gold foil. The gold foil was surrounded by photographic plate or Zinc sulphide (ZnS) fluorescent screen to detect the particles emitting from the radiation.

He bombarded the gold foil with α – particles (20,000) from a radioactive source (polonium metal). He observed that most of the α – particles (19990) passed through the gold foil without any change in direction. It showed that most of the volume occupied by the atom is empty.

Some α – particles (8) were deflected through smaller angle. Only few (2) were bounced back at their original way. Rutherford concluded that an atom contains a positive portion i.e. nucleus in the center of the atom and So, α – particles passing near this portion were repelled. Because, α – particles are also positively charged particles and similar charges repel each other.

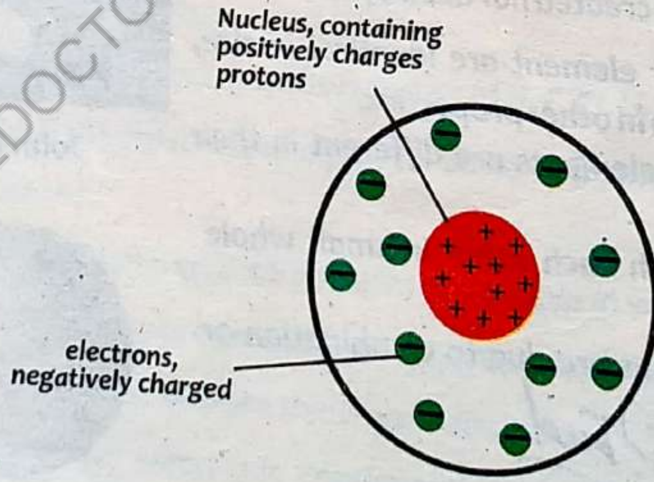
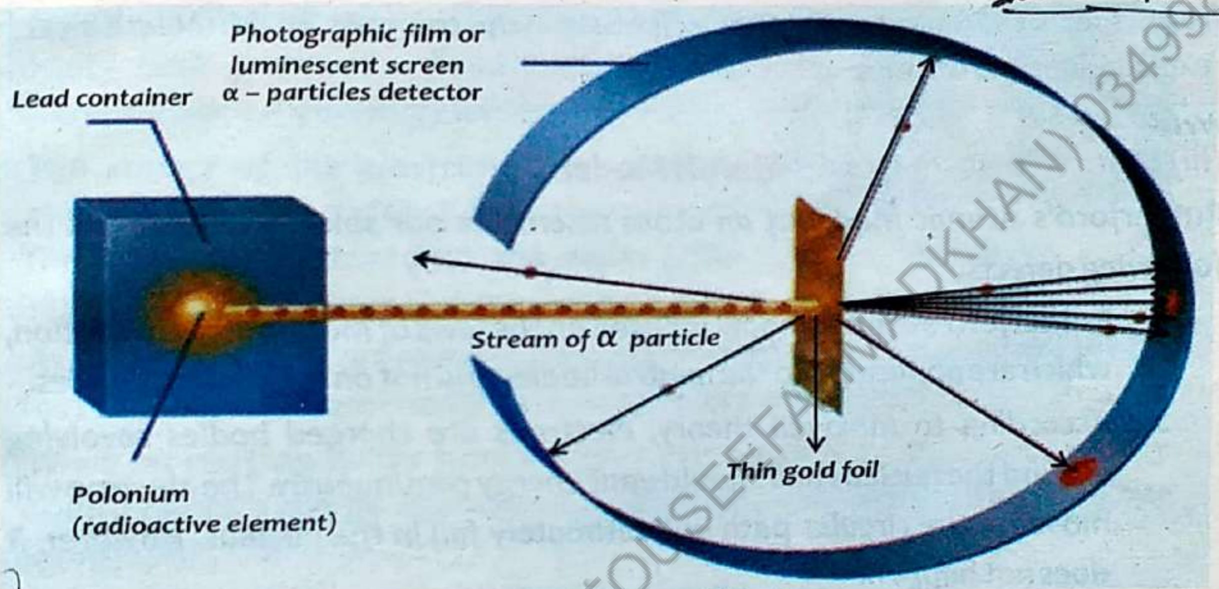


Fig: 2.3 Rutherford model of atom

If α – particles pass very closely to nucleus, they deflected through large angles. Similarly, if they do not pass close to nucleus, they either get deflected through very small angles or do not get deflected at all.

On the basis of his conclusion drawn from this experiment, he proposed a new model of atom called planetary model (similar to the solar system). Which put all the protons in the nucleus and the electrons orbited around the nucleus like planets around the sun.

Show the diagram



Q3
write down

Fig: 2.4 Discovery of Nucleus

(ii) Main points of Rutherford's atomic model ?

1. **Atom consists of positively charged central portion called nucleus. It contains protons and neutrons.**
2. **Electrons are revolving around the nucleus with very high speed just like the planets around the sun.**
3. **The electrons revolving around the nucleus would require centripetal force. The attractive force of the nucleus on electrons provides centripetal force to the electron.**
4. **The size of the nucleus is very small as compared to the size of the atom and most of the volume occupied by the atom is empty.**
5. **Atom is neutral. As the number of electrons is numerically equal to that of protons.**
6. **Nucleus is responsible for mass and energy of the atom.**) End.

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Rutherford in 1911, performed an experiment and proposed a model of an atom. These suggestions remain unchallenged for long time. But this model could not explain the stability of an atom and line spectrum for an atom. Bohr leaped over difficulty by using quantum theory of radiation that was proposed by Max Plank. Bohr atomic theory explains nicely the stability of an atom and also explains why an atom gives line spectrum. Development of Bohr's atomic model explains how interpretations of results of experiments of other scientists help chemists to formulate new explanations and theories.

Q(4) write down ?

(iii) Defects in Rutherford's Atomic Model

Rutherford's Atomic Model of an atom resembles our solar system. It has the following defects.

1. **Rutherford's atomic model is based on the laws of motion and gravitation, which are applicable to the neutral bodies and not on the charged bodies.**
2. **According to Maxwell theory, electrons are charged bodies revolving around the nucleus and should emit energy continuously. The electron will move on the circular path and ultimately fall in the nucleus. However, it does not happen.**
3. ***If electrons radiate energy continuously then continuous spectrum will be obtained but actually line spectrum is obtained.***
4. **It does not provide any explanation about the chemical properties of the elements.** End

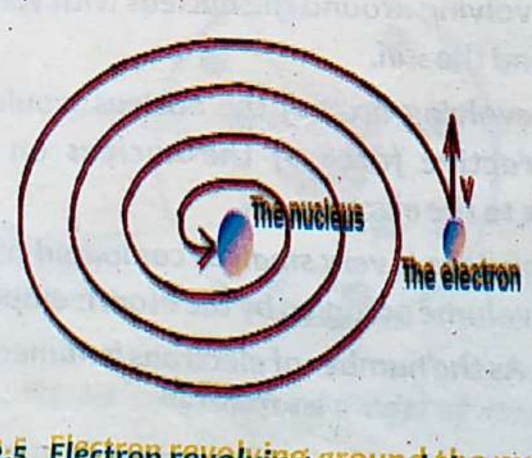


Fig: 2.5 Electron revolving around the nucleus write down main points of

Q8) ~~Explain~~ model?

2.1.2 Neil Bohr's atomic ~~theory~~

Neil Bohr

In 1913, Neil Bohr presented a new atomic theory to overcome the defects of Rutherford's atomic model. Bohr considered hydrogen atom as a model. This atomic model was based on the following assumptions.

1. *Electrons are revolving around the nucleus in one of the fixed circular paths called shells or orbits. Each orbit has a fixed energy. These orbits are also called energy levels.*
2. *The energy of the electron in an orbit is proportional to its distance from the nucleus. The farther the electron from the nucleus, the higher will be the energy and vice versa.*
3. *As long as electrons are revolving around the nucleus in fixed circular orbits, they do not gain or lose energy .i.e. energy of an orbit is fixed.*
4. *When an electron jumps from a higher energy orbit to the lower energy orbit, it radiates energy and when it jumps from lower energy orbit to higher energy orbit, it absorbs energy.*

The energy difference between two levels is given by,

$$\Delta E = E_2 - E_1 = h\upsilon$$

Where,

- h = *Plank's constant*
- υ = *frequency*
- E_1 = *lower energy orbit*
- E_2 = *higher energy orbit*
- ΔE = *energy difference*

5. *Electron can reside in any one of the orbits and cannot stay in between them.*
6. *Angular momentum (mvr) of an electron is an integral multiple of* $\frac{h}{2\pi}$

$$mvr = \frac{nh}{2\pi}$$

Where,

n = number of shell = 1, 2, 3, 4....

m= mass of electron

v = velocity of an electron

r = radius of the orbit in which electron is revolving, and

h= Plank's constant its value is 6.6262×10^{-34} J.s

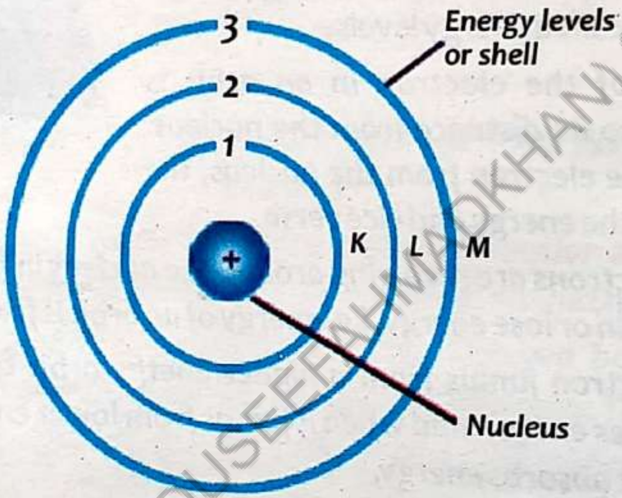


Fig: 2.6 Bohr' s Atomic Model

The orbit defined by Bohr, is also called a shell or an energy level. These are represented by alphabets K, L, M, N and so on. The electrons present in K–shell (n=1) have the least energy and it is nearest to the nucleus. L–shell (n= 2) is the next shell which is higher in energy than the K – shell and so on.

Activity:2.1

Make a model/diagram based on Bohr's atomic theory. Display this model in your classroom/ laboratory.

(6) What do you means by fundamental particel ?
Compaire properties of Electron, protons, and neutrons.

2.1.3 Fundamental particles of an atom

Modern research showed that an atom consists of many subatomic particles. Three subatomic particles Proton, Neutron and Electron are ~~very important to the chemists. These particles are~~ the fundamental particles.

1. Electron

Electron is a negatively charged particle. Its mass is equal to 0.000548597 a.m.u. or 9.11×10^{-31} kg. Charge of an electron is 1.6022×10^{-19} coulomb with negative sign. Electrons are very light small particles which revolve around the nucleus in orbits.

2. Proton

Proton is a positively charged particle. Its mass is equal to 1.0072766 a.m.u. or 1.6726×10^{-27} kg. Charge of proton is 1.6022×10^{-19} coulomb with positive sign. Proton is 1837 times heavier than an electron. Protons are present in the nucleus of an atom.

3. Neutron

Neutron is a neutral particle because it has no charge. Its mass is equal to 1.0086654 a.m.u. or 1.6749×10^{-27} kg. Neutron is 1842 times heavier than an electron. Neutrons are present in the nucleus of an atom.

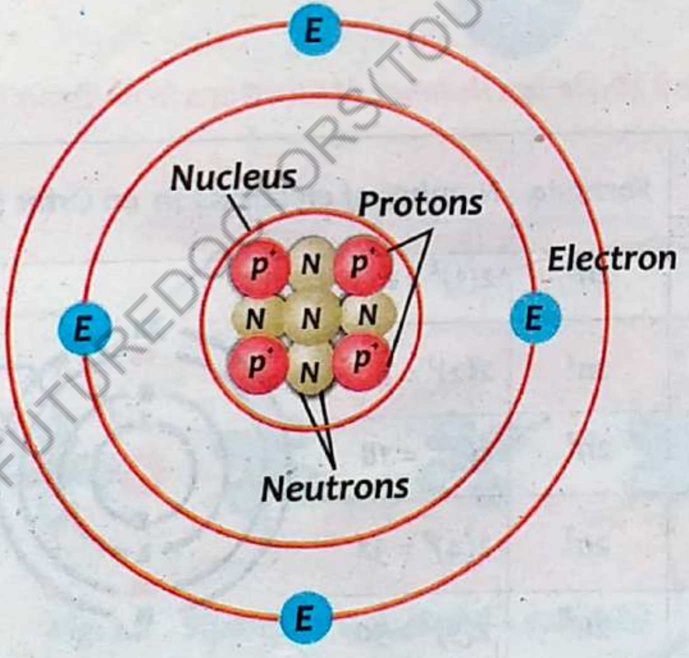


Fig: 2.7 Structure of Atom

Table 2.1 : Summary of the subatomic particles

Particle	Symbol	Unit charges	Charge (C)	Relative mass (amu)	Mass (kg)
Electron	e^-	–1	1.6022×10^{-19}	0.00054859	9.11×10^{-31}
Proton	P^+	+1	1.6022×10^{-19}	1.0072766	1.6726×10^{-27}
Neutron	n^0	0	0	1.0086654	1.6749×10^{-27}

Ans (7) Define electronic with example.

2.2 Electronic configuration

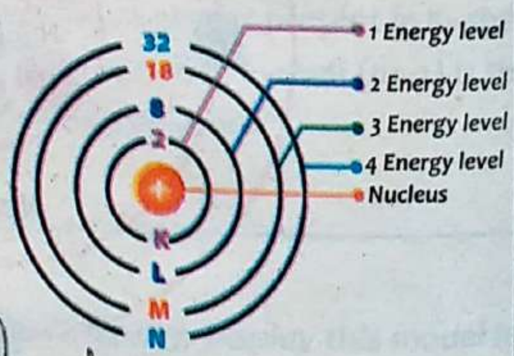
The distribution or arrangement of electrons around the nucleus in orbits or shells is called the electronic configuration.

The electrons are distributed around the nucleus in various orbits or energy levels or shells. These orbits are numbered as 1, 2, 3, 4, 5, 6, 7…… These orbits are represented by K, L, M, N etc. The maximum number of electrons in a particular orbit is given by the formula $2n^2$, Where $n = 1, 2, 3, 4...$

The maximum number of electrons in first, second and third orbit are 2, 8 and 18 respectively.

Table 2.2: Orbits and Maximum Number of Electrons in an Orbit (Energy Level)

Number of Orbit	Name of Orbit	Formula	Number of electrons in an Orbit (Energy Level)
1	K	$2n^2$	$2(1)^2 = 2$
2	L	$2n^2$	$2(2)^2 = 8$
3	M	$2n^2$	$2(3)^2 = 18$
4	N	$2n^2$	$2(4)^2 = 32$
5	O	$2n^2$	$2(5)^2 = 50$



End (7)

Q (8) Difine sell and sub-shell with examples

2.2.1 Concepts of s and p sub-shell

(i) Shell (Orbit)

According to Bohr's atomic theory, **electrons are revolving around the nucleus in one of the fixed circular paths called shells or orbits.** Each shell has a fixed energy. These orbits are also called energy levels. Each shell is described by 'n' value. 'n' can have values 1,2,3, 4...

As the value of 'n' increases, shell number increases. The distance of electron from the nucleus increases and energy of the electron also increases.

(ii) Sub-Shell (sub-orbit)

A shell or energy level is divided into sub-shells or sub – energy levels. 'n' value of a shell is placed before the symbol for a sub-shell.

There are four sub-shells. The sub – shell are represented by s, p, d and f. These are spectroscopic terms which stand for sharp, principal, diffused and fundamental.

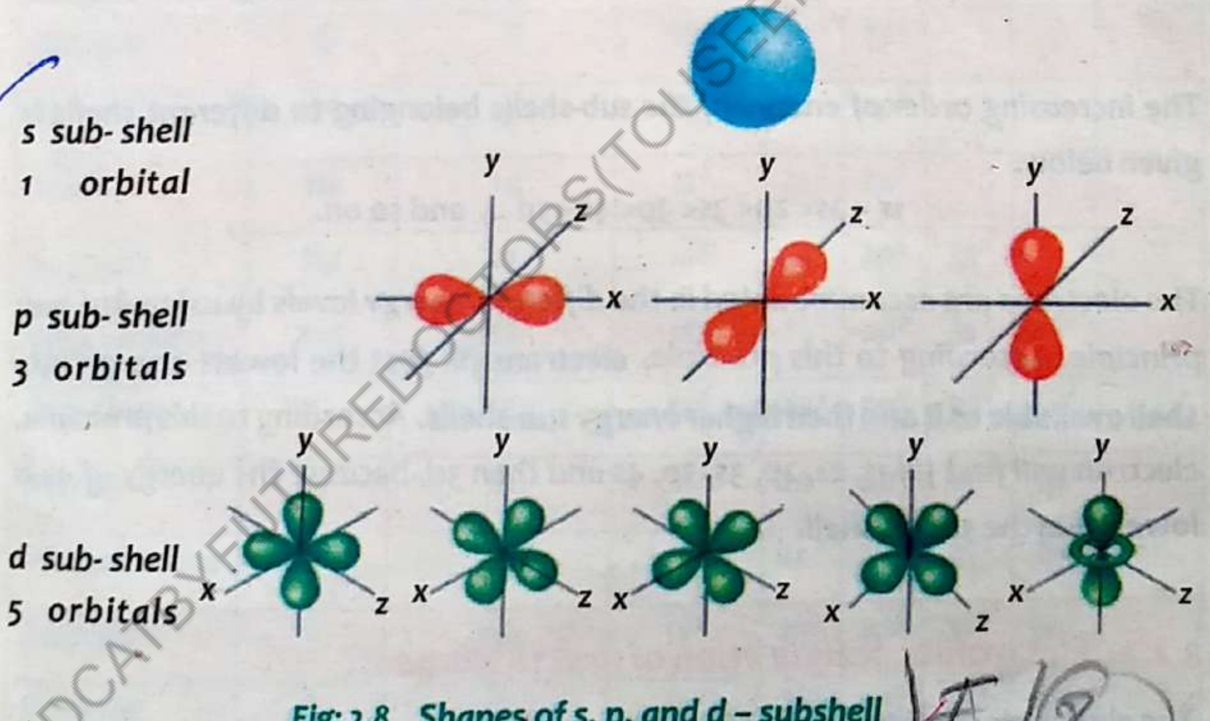


Fig: 2.8 Shapes of s, p, and d – subshell

End (8)

Note: The f sub-shell has seven orbitals and has complicated shape.

i. s sub-shell can accommodate a maximum of '2' electrons,

ii. p sub-shell can accommodate a maximum of '6' electrons,

iii. d sub-shell can accommodate a maximum of '10' electrons, and

Iv. f sub-shell can accommodate a maximum of '14' electrons. Excel clients.

Table 2.3: Shells, Maximum Number of Electrons in a Shell and sub-shells

Number of Shell (n)	Name of Shell	Formula	Number of electrons in Shell	Sub shell(s)	Electrons in the sub shell
1	K	$2n^2$	$2(1)^2 = 2$	s	$1s^2$
2	L	$2n^2$	$2(2)^2 = 8$	s and p	$2s^2, 2p^6$
3	M	$2n^2$	$2(3)^2 = 18$	s, p and d	$3s^2, 3p^6, 3d^{10}$
4	N	$2n^2$	$2(4)^2 = 32$	s, p,d and f	$4s^2, 4p^6, 4d^{10}, 4f^{14}$

The increasing order of energy of the sub-shells belonging to different shells is given below.

$$1s < 2s < 2p < 3s < 3p < 4s < 3d \ldots \text{ and so on.}$$

The electrons are accommodated in the different energy levels by using Auf Bau principle. According to this principle, **electrons fill first the lowest energy sub-shell available to it and then higher energy sub-shells.** According to this principle, electron will first fill 1s, 2s, 2p, 3s, 3p, 4s and then 3d, because the energy of 4s is lower than the 3d sub-shell.

2.2.3 Electronic configuration of first 18 elements

The electronic configuration of first 18 elements is given below.

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p.

write down electronic

Eighteen

Table 2.4: Electronic Configuration of First 18 Elements

of periodic table?

try q)

Element	Symbol	Atomic Number	Electronic Configuration					
			n= 1 = K	n= 2 = L		n= 3 = M		
			1s	2s	2p	3s	3p	3d
Hydrogen	H	1	$1s^1$					
Helium	He	2	$1s^2$					
Lithium	Li	3	$1s^2$	$2s^1$				
Beryllium	Be	4	$1s^2$	$2s^2$				
Boron	B	5	$1s^2$	$2s^2$	$2p^1$			
Carbon	C	6	$1s^2$	$2s^2$	$2p^2$			
Nitrogen	N	7	$1s^2$	$2s^2$	$2p^3$			
Oxygen	O	8	$1s^2$	$2s^2$	$2p^4$			
Fluorine	F	9	$1s^2$	$2s^2$	$2p^5$			
Neon	Ne	10	$1s^2$	$2s^2$	$2p^6$			
Sodium	Na	11	$1s^2$	$2s^2$	$2p^6$	$3s^1$		
Magnesium	Mg	12	$1s^2$	$2s^2$	$2p^6$	$3s^2$		
Aluminum	Al	13	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^1$	
Silicon	Si	14	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^2$	
Phosphorous	P	15	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^3$	
Sulphur	S	16	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^4$	
Chlorine	Cl	17	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^5$	
Argon	Ar	18	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	

End

Test Yourself.

1. Sodium has 11 protons and 12 neutrons. Write the electron configuration of the sodium atom.
2. Potassium has the electronic arrangement 2,8,8,1.
   a. What is its proton number?
   b. Draw its electronic structure.
3. Write the electronic configurations of N, P and Ar.

What is isotopes? write down isotopes hydrogen

2.3 Isotopes

These are the atoms of the same element which have same atomic number but different mass number. The difference in mass of an atom in an element is due to the difference in number of neutrons. In other words, we can say that **these are the atoms of the same element which have same number of protons and electrons but different number of Neutrons.** Isotopes have the same chemical properties but different physical properties.

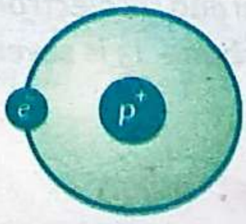
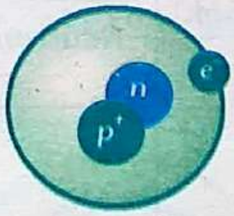
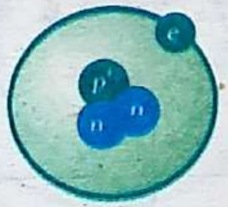
2.3.1 Examples of Isotopes

(i) Isotopes of Hydrogen

Hydrogen has three isotopes. These are Protium (H), Deuterium (D) and Tritium (T). Protium (H) is ordinary hydrogen. It is the most abundant isotope (99.985%) of hydrogen. Deuterium (D) is about 0.015 % and Tritium is rarely found in nature. All the isotopes have one proton in their nuclei and one electron in their orbit, the difference lies in the number of neutrons. The number of neutrons in Protium (H) is zero, Deuterium (D) is one and Tritium (T) is two in its nucleus.

The chemical properties of the Isotopes are same due to the same number of valence electrons but there is difference in physical properties like density, melting point and boiling point etc.

Table 2.5: Isotopes of Hydrogen

Isotope Name	Atomic structure	Symbol	Number of Neutrons	Isotopic abundance (percentage)
Protium (H)	 Protium	$^{1}_{1}H$	0	99.985
Deuterium (D)	 Deuterium	$^{2}_{1}H$ or D	1	0.015
Tritium (T)	 Tritium	$^{3}_{1}H$ or T	2	very rare isotope

10
End

Table 2.6: Physical properties of H_2O and D_2O

S. No	Property	H_2O	D_2O
1.	Molecular mass (a.m.u)	18.02	20.03
2.	Density at 0°C ($g.\ cm^{-3}$)	1.000	1.105
3.	M.P (°C)	0.00	3.82
4.	B.P (°C)	100.00	101.42

Q write down ?

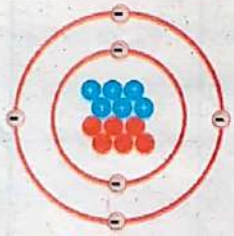
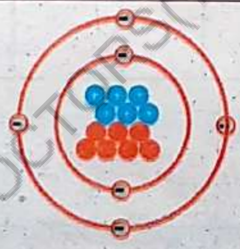
(ii) Isotopes of Carbon

Atomic number of carbon is six. It is the first member of the 4th group. There are three isotopes of carbon in nature, Carbon – 12, Carbon – 13 and Carbon – 14. These are represented by $^{12}_{6}C$, $^{13}_{6}C$ and $^{14}_{6}C$.

All the isotopes have six protons in their nuclei and six electrons in their orbit. The number of neutrons in Carbon – 12 is six, Carbon – 13 is seven and Carbon – 14 is eight in their nucleus.

Ans (11)

Table 2.7: Isotopes of Carbon

Isotope Name	Atomic structure	Symbol	Number of Neutrons	Isotopic abundance (percentage)
Carbon – 12	 Carbon - 12 6 Protons 6 Neutrons	$^{12}_{6}C$	6	98.98
Carbon – 13	 Carbon - 13 6 Protons 7 Neutrons	$^{13}_{6}C$	7	1.109
Carbon – 14	 Carbon - 14 6 Protons 8 Neutrons ⊕ Proton ○ Neutron ⊖ Electron	$^{14}_{6}C$	8	1 part per trillion (ppt)

Q12 write down Isotopes of Chlorine?

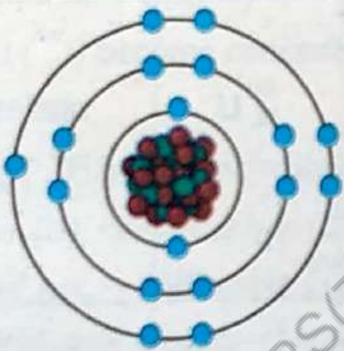
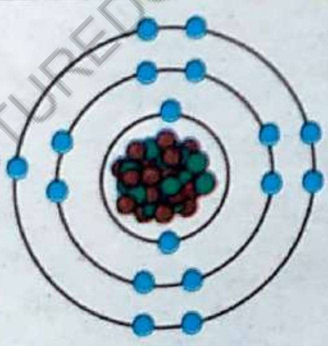
(iii) Isotopes of Chlorine

Atomic number of Chlorine is 17. It is the second member of the 7^{th} group. There are two isotopes of Chlorine in nature, Chlorine – 35 and Chlorine – 37. The abundance of Chlorine – 35 is 75.53% and that of Chlorine – 37 is 24.47% in nature. These are represented by $^{35}_{17}Cl$ and $^{37}_{17}Cl$.

Both isotopes have 17 protons in their nuclei and 17 electrons in their orbits. The number of neutrons in Chlorine – 35 is 18 and in Chlorine – 37 is 20.

Ans 12

Table 2.8: Isotopes of Chlorine

Isotope Name	Atomic structure	Symbol	Number of Neutrons	Isotopic abundance (percentage)
Chlorine – 35	 Chlorine - 35 17 Protons 18 Neutrons	$^{35}_{17}Cl$	18	75.53
Chlorine – 37	 Chlorine - 37 17 Protons 20 Neutrons	$^{37}_{17}Cl$	20	24.47

End

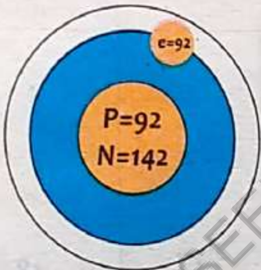
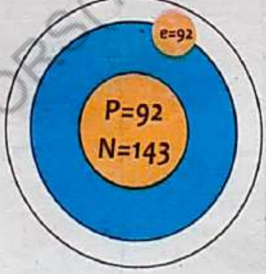
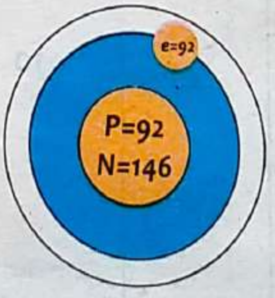
Q13 Write down

(iv) Isotopes of Uranium ?

Atomic number of Uranium is 92. There are three isotopes of Uranium in nature, Uranium – 234, Uranium - 235 and Uranium–238. The abundance of Uranium – 234 is 0.05%, Uranium – 235 is 0.75% and Uranium - 238 is the most abundant 99.245% in nature. These are represented by $^{234}_{92}U$, $^{235}_{92}U$ and $^{238}_{92}U$.

All the isotopes have 92 protons in their nuclei and 92 electrons in their orbits. The number of neutrons in Uranium – 234 is 142, Uranium – 235 is 143, and Uranium – 238 is 146.

Table 2.9 : Isotopes of Uranium

Isotope Name	Atomic structure	Symbol	Number of Neutrons	Isotopic abundance (percentage)
Uranium – 234		$^{234}_{92}U$	142	0.05
Uranium – 235		$^{235}_{92}U$	143	0.75
Uranium – 238		$^{238}_{92}U$	146	99.245

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Estimation of the age of the plant and animal remains by measuring the amount of radioactive decay products is called chemical or carbon dating. A more recent method involves the use of a radioactive isotope, C – 14. This method is used in the study of deep sea sedimentation, dates of volcanic and glacier activity. The age of a uranium containing material can be determined by measuring the percentage of lead formed as a result of disintegration of uranium.

Q(4) Write down uses of Isotopes?

2.3.2 Uses of Isotopes:

Isotopes are used in chemical, agricultural and medical research for diagnosing and treatment of diseases. Isotopes of certain elements show radioactivity while others do not. Some uses of isotopes are given below.

1. **Iodine-131 ($^{131}_{53}I$) become concentrated in the Thyroid gland and is used as cure for goiter.**
2. **Iodine-123 ($^{123}_{53}I$) is used for brain imaging.**
3. **Deuterium (D), heavy Carbon (C-13), heavy Nitrogen (N-15), heavy Oxygen (O-18) and heavy Iodine-131 ($^{131}_{53}I$) are being used as tracer elements in biochemical and physio–chemical research to trace the path of the element to the defective or obstructed part.**

Fig: 2.9 Chemotherapy (Treatment of cancer)

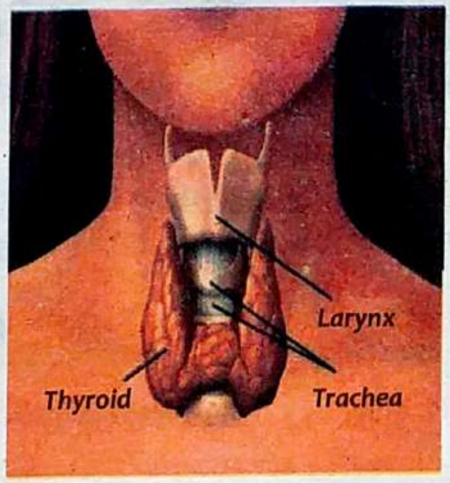


Fig: 2.10 Thyroid gland

4. *Radium irradiation and Cobalt-60 ($^{60}_{27}Co$) are used in the treatment of cancer and for diagnosis of tumors.*
5. Sodium (Na-24) is used for the identification of blood circulation problems in patients.
6. Carbon-14 is used to trace the path of carbon in photosynthesis.
7. Americium -241 is used in smoke detectors. It is also use to determine where oil wells should be drilled.) End.
8. *Californium – 252 is used to measure the moisture content of soil in road construction and in building construction. It is also used to inspect airplane luggage for hidden explosives.*
9. *Krypton -85 is used in electrical cloth washers to measure the dust and pollutants level.*

Carbon dating (carbon -14) | Smoke Detectors (americium-241) | Agricultural Research (carbon -14, Phosphorous-32)

Fig: 2.10 Different uses of isotopes

Key Points

- Greeks proposed that matter is composed of small particles called atoms.
- According to Dalton's theory, matter is composed of small indivisible particles called atoms.
- Electrons, protons and neutrons are the fundamental subatomic particles of an atom.
- Rutherford's model of atom consists of a positively charged nucleus. This model states that, electrons revolve around the nucleus like the planets around the sun. Most of the space occupied by an atom is empty.
- According to Bohr's atomic theory, the electrons revolve around the nucleus in certain fixed circular orbits called shells or energy levels.
- The arrangement of electrons around the nucleus in orbits or shells is called the electronic configuration.
- Electrons are revolving around the nucleus in one of the fixed circular paths called shell or orbits. Each shell has a fixed energy. These orbits are also called energy levels.
- Shell or energy levels are named alphabetically as K, L, M, N and so on.
- The maximum number of electrons in a particular shell is given by the formula $2n^2$.
- Energy levels or shells further split into sub-energy levels or sub-shells. The sub-shells are designated by s (Sharp), p (principal), d (diffused) and f (fundamental).
- Isotopes are atoms of the same element having different number of neutrons.
- Isotopes are used in chemical, medical and agricultural research.

Exercise

Choose the correct option.

i. The maximum number of electrons in third energy level is:

(a) 10 (b) 18

(c) 32 (d) 64

ii. Mass of an atom is mostly due to its

(a) nucleus (b) neutrons

(c) electrons (d) protons

iii. If Rutherford had used neutrons instead of alpha particles in his scattering experiment, the neutrons would

(a) not deflect because they have no charge

(b) have deflected more often

(c) have been attracted to the nucleus easily

(d) have given the same results

iv. Electron in its ground state does not

(a) spin (b) revolve

(c) radiate energy (d) reside in orbit

v. Which statement about ${}^{12}_{6}X$ and ${}^{14}_{6}Y$ is false

(a) they are isotopes

(b) they are the same elements

(c) they have the same number of electrons

(d) they have the same number of neutrons

vi. The neutron particle

(a) has a mass equal to that of a electron

(b) has a mass approximately equal to that of a proton

(c) has charge equal to but opposite to that of an electron

(d) has a positive charge

vii. Isotopes of the same element have

(a) the same number of protons (b) the same number of neutrons

(c) different number of electrons (d) the same mass number

viii. Which one is the lightest?

a) an alpha particle b) a hydrogen atom

c) an electron d) a proton

ix. The nucleus of an atom has all of the following characteristics except that it

(a) is positively charged

(b) is very dense

(c) contains nearly all of the atom's mass

(d) contains nearly all of the atom's volume

x. L – shell has sub-shell(s)

(a) s (b) s and p

(c) s, p and d (d) s, p, d and f

SHORT QUESTIONS

Answer briefly the following questions.

i. Aluminum is represented as $^{27}_{13}Al$. Draw the structures of Aluminum. Write its electronic configuration.

ii. The energy of an electron in K and L shells is the same or different. Explain.

iii. Draw the structures of hydrogen isotopes.

iv. How many electrons are present in each of the following atoms? Assuming that each is a neutral atom, identify the element.

a. $1s^2 2s^2 2p^6 3s^1$ b. $1s^2 2s^2 2p^6 3s^2 3p^5$ c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

v. Why an atom is considered as neutral particle? Give reason(s).

vi. The mass of an atom is present in its nucleus. Can you explain it?

vii. What is the reason that physical properties of the isotopes are different but their chemical properties are the same?

viii. Draw the structures of carbon isotopes. Then write down the number of proton, neutron and electron.

ix. How many electrons could be contained in the K, L, M and N energy levels.

x. Write detailed electronic configurations for $^{7}_{3}Li$, $^{12}_{6}C$, and $^{24}_{12}Mg$.

xi. Write the symbol for an isotope:

(a) Containing one proton and two neutrons

(b) For which the atomic number is one and there is one neutron

(c) For which the atomic number is one and the mass number is also one.

LONG QUESTIONS

(i) Why Dalton's atomic theory is considered as a base for modern atomic concepts.

(ii) Summarize Rutherford's model of an atom and explain how he developed this model based on the results of his famous gold-foil experiment.

(iii) State the postulates which Bohr suggested to overcome the shortcomings of the Rutherford's atomic model.

(iv) Complete the following table for neutral atoms of specific isotopes:

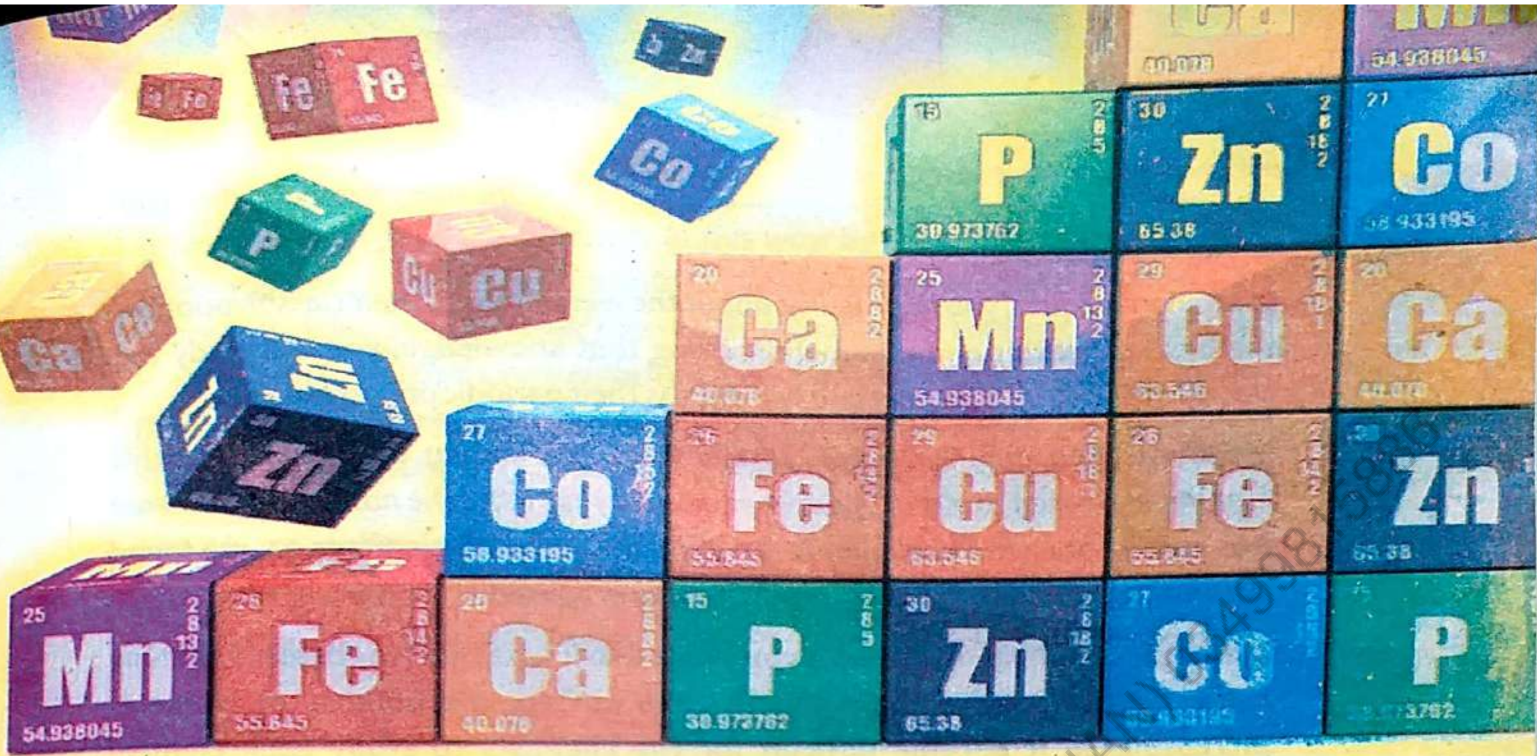
	Isotopic Symbol	Atomic No.	Mass No.	Number of Protons	Number of electrons	Number of neutrons
a	$^{131}_{54}Xe$					
b		27	59			
c			144			84
d		22				26
e					72	106
F			128	52		
g				18		22

(v) (a) Define energy level and sub-energy level.

(b) Explain the distribution of electrons in various energy levels and sub energy levels for first four elements of the periodic table.

Project Work

- The study of atomic structure and the nucleus produced a new field of medicine called nuclear medicine. Describe the use of radioactive tracers to detect and treat diseases.
- In groups of 5 to 8 students, make and present different models based on Rutherford and Bohr's atomic theories.



Unit 3 Periodic Table and Periodicity of Properties

After studying this unit, the students will be able to;

- *Distinguish between a period and a group in the periodic table.*
- *State the periodic law.*
- *Classify the elements (into two categories: groups and periods) according to the configuration of their outer most electrons.*
- *Determine the demarcation of the periodic table into an s block and p block.*
- *Explain the shape of the periodic table.*
- *Determine the location of families on the Periodic Table.*
- *Recognize the similarity in the chemical and physical properties of elements in the same family of elements.*
- *Identify the relationship between electronic configuration and the position of an element on the periodic table.*
- *Explain how shielding effect influences periodic trends.*
- *Describe how electronegativities change within a group and within a period in the periodic table.*

Introduction

In previous grades, you have learnt about the elements and their classification on the basis of metals and non-metals. Using that knowledge, you will study the classification of all elements and will discuss their periodic properties.

In the year 1800, only thirty four (34) elements were known. By the year 1870, this number increased to almost double and in the year 1974, the number of elements was 105. At present 118 elements are known. Out of 118 elements 92, are found naturally on earth and the rest are artificially prepared.

Initially the number of elements was limited so it was easy for chemists to study these elements individually. But as time passed, the number of elements increased, so it became difficult for the chemists to study each element individually. It was felt that the elements might be properly arranged so that they could be easily studied. By doing so, one could easily get more information with little effort. Many chemists, including Lavoisier (1787), Newlands (1864), and others attempted to classify the elements systematically. It was a Russian chemist Mendeleev (1869), who made the most successful classification based on the atomic masses of the elements.

Society, Technology and Science

The development of periodic table started in 1800. Many scientists in different ages attempted for classifying the elements in different ways. Initially elements were divided into metal and non-metals. In 1817, J. W. Dobereiner, a German chemist classified chemically similar elements in groups of three on the basis of their atomic masses. These groups are called triads. In triads, the atomic weight of the middle element was the arithmetic mean of the atomic weights of the other two elements. For example, Li^7, Na^{23}, K^{39}

$$\text{atomic weight of } Na = \frac{7+39}{2} = 23$$

It failed to classify all the elements in this way. By 1864, more elements had been discovered. An English chemist J. Newland classified the elements in order of their increasing atomic masses. He classified the elements into groups of seven and found that every eighth element had the properties similar to the first one of the series. He called these groups octaves. Lithium (Li) resembles Sodium (Na) and fluorine (F) resembles chlorine (Cl), as shown below.

Li 7	Be 9	B 11	C 12	N 14	O 16	F 19
Na 23	Mg 24	Al 27	Si 28	P 31	S 32	Cl 35.5

He could not able to classify all the elements in the same manner, so the law of octaves failed to classify all the elements.

In 1869, two chemists, Lother Meyer in Germany and Dmitri Mendeleev in Russia, at about the same time independently developed very similar arrangements of the elements. They found that while arranging the elements in order of their increasing atomic mass, certain physical and chemical properties are repeated at regular intervals.

Lother Meyer plotted atomic volumes (densities) of the elements against their atomic masses and found that similar elements occupied similar positions in the graph.

Dmitri Mendeleev arranged the elements in the form of table. He arranged 65 elements in periods and groups. In this table, he left vacant spaces for the elements which were not discovered at that time and to be placed in these vacant spaces. On the basis of this table, he predicted the properties of certain elements very accurately that were yet to be discovered.

Mendeleev classification was based on atomic mass. However, with the discovery of atomic structure, it became clear that elements varied regularly with atomic number not with the atomic mass. It had also become known that the properties of the elements depended on the number and arrangement of electrons in each of its atom.

3.1 Modern Periodic Table

Jeffreys Moseley

The modern periodic table was put forward by Henry Gwyn Jeffreys Moseley, an English physicist. This table is based on the atomic number. The modern periodic law states that the physical and chemical properties of the elements are the periodic function of their atomic number. Periodic table is formed on the basis of this law. In the simplified form it is divided into eight vertical columns known as groups / families and seven horizontal rows known as periods.

3.1.1 Periods

The horizontal rows of elements in the periodic table are called periods. They are numbered from 1 to 7. From the table, we can see that all the elements in the same period have the same number of electronic shells. In the first, second and third period, the number of shells are 1, 2 and 3 respectively as shown in fig. 3.1

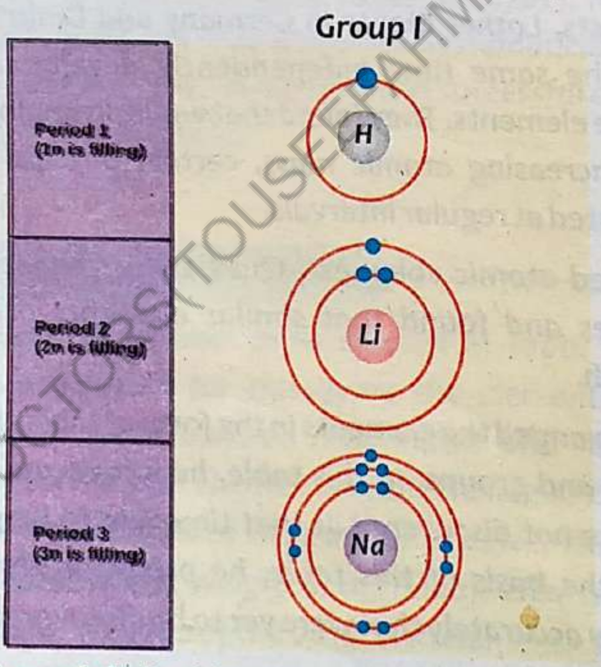


Fig. 3.1 Numbers of shell with correspondence to the period number

Scientific Information

- *The only letter that does not appear on the periodic table is J.*
- *The rarest naturally-occurring element in the earth's crust is Astatine. The entire crust appears to contain about 28 g of this element.*

Periodic Table of the Elements

Black	Alkali Metals
Blue	Alkaline Earth Metals
Peach	Transition Metals
Yellow	Other Metals
Light green	Nonmetals
Light grey-green	Lanthanoids
Dark green	Noble Gases
Orange	Actinoids
White	Metalloid
Purple	Halogens

key

element name
atomic number
symbol
atomic weight

hydrogen 1 **H** 1.00794																	helium 2 **He** 4.002602
lithium 3 **Li** 6.941	beryllium 4 **Be** 9.012182											boron 5 **B** 10.811	carbon 6 **C** 12.0107	nitrogen 7 **N** 14.00674	oxygen 8 **O** 15.9994	fluorine 9 **F** 18.9984	neon 10 **Ne** 20.1797
sodium 11 **Na** 22.98977	magnesium 12 **Mg** 24.3050											aluminium 13 **Al** 26.981538	silicon 14 **Si** 28.0855	phosphorus 15 **P** 30.97376	sulphur 16 **S** 32.065	chlorine 17 **Cl** 35.453	argon 18 **Ar** 39.984
potassium 19 **K** 39.0983	calcium 20 **Ca** 40.078	scandium 21 **Sc** 44.95591	titanium 22 **Ti** 47.867	vanadium 23 **V** 50.9415	chromium 24 **Cr** 51.9961	manganese 25 **Mn** 54.93805	iron 26 **Fe** 55.845	cobalt 27 **Co** 58.9332	nickel 28 **Ni** 58.6934	copper 29 **Cu** 63.546	zinc 30 **Zn** 65.409	gallium 31 **Ga** 69.723	germanium 32 **Ge** 72.64	arsenic 33 **As** 74.9216	selenium 34 **Se** 78.96	bromine 35 **Br** 79.904	krypton 36 **Kr** 83.798
rubidium 37 **Rb** 85.4678	strontium 38 **Sr** 87.62	yttrium 39 **Y** 88.90585	zirconium 40 **Zr** 91.225	niobium 41 **Nb** 92.90638	molybdenum 42 **Mo** 95.94	technetium 43 **Tc** [98]	ruthenium 44 **Ru** 101.07	rhodium 45 **Rh** 102.9055	palladium 46 **Pd** 106.42	silver 47 **Ag** 107.8682	cadmium 48 **Cd** 112.411	indium 49 **In** 114.818	tin 50 **Sn** 118.710	antimony 51 **Sb** 121.760	tellurium 52 **Te** 127.60	iodine 53 **I** 126.9045	xenon 54 **Xe** 131.293
caesium 55 **Cs** 132.90545	barium 56 **Ba** 137.327	lutetium 71 **Lu** 174.967	hafnium 72 **Hf** 178.49	tantalum 73 **Ta** 180.9479	tungsten 74 **W** 183.84	rhenium 75 **Re** 186.207	osmium 76 **Os** 190.23	iridium 77 **Ir** 192.217	platinum 78 **Pt** 195.078	gold 79 **Au** 196.96655	mercury 80 **Hg** 200.59	thallium 81 **Tl** 204.3833	lead 82 **Pb** 207.2	bismuth 83 **Bi** 208.980	polonium 84 **Po** [209]	astatine 85 **At** [210]	radon 86 **Rn** [222]
francium 87 **Fr** [223]	radium 88 **Ra** [226]	lawrencium 103 **Lr** [262]	rutherfordium 104 **Rf** [261]	dubnium 105 **Db** [262]	seaborgium 106 **Sg** [266]	bohrium 107 **Bh** [264]	hassium 108 **Hs** [269]	meitnerium 109 **Mt** [278]	darmstadtium 110 **Ds** [271]	roentgenium 111 **Rg** [272]	copernicium 112 **Cn** [285]	nihonium 113 **Nh** [286]	flerovium 114 **Fl** [287]	moscovium 115 **Mc** [289]	livermorium 116 **Lv** [293]	tennessine 117 **Ts** [294]	oganesson 118 **Og** [294]

lanthanum 57 **La** 138.9055	cerium 58 **Ce** 140.116	praseodymium 59 **Pr** 140.90765	neodymium 60 **Nd** 144.24	promethium 61 **Pm** [145]	samarium 62 **Sm** 150.36	europium 63 **Eu** 151.964	gadolinium 64 **Gd** 157.25	terbium 65 **Tb** 158.9253	dysprosium 66 **Dy** 162.50	holmium 67 **Ho** 164.930	erbium 68 **Er** 167.259	thulium 69 **Tm** 168.934	ytterbium 70 **Yb** 173.04	
actinium 89 **Ac** [227]	thorium 90 **Th** 232.038	protactinium 91 **Pa** 231.0359	uranium 92 **U** 238.0289	neptunium 93 **Np** [237]	plutonium 94 **Pu** [244]	americium 95 **Am** [243]	curium 96 **Cm** [247]	berkelium 97 **Bk** [247]	californium 98 **Cf** [251]	einsteinium 99 **Es** [252]	fermium 100 **Fm** [257]	mendelevium 101 **Md** [258]	nobelium 102 **No** [259]	

There are seven periods (horizontal rows) in the periodic table.

- 1^{st} period contain two elements i.e. hydrogen (H) and Helium (He). This is also called the shortest period of the periodic table.
- 2^{nd} and 3^{rd} periods each contains eight elements and are called short periods.
- 4^{th} and 5^{th} periods have 18 elements each. In these elements, eight elements are called representative elements and the rest of 10 elements are called Transition elements. In 4^{th} period it ranges from (scandium) to $^{45}_{21}Sc$ $^{65}_{30}Zn$ (zinc) and in 5^{th} periods it ranges from $^{89}_{39}Y$ (Yttrium) to $^{112}_{48}Cd$ (cadmium).
- 6^{th} period contains 32 elements of which 8 are representative elements, 10 are transition elements and the rest 14 are known as Lanthanides series.
- 7^{th} period has also 8 representative elements, 10 transition element and the rest are called Actinides series.

Periodic table is still incomplete and with the discoveries of new elements, it will keep on adding these elements over the time. The number of valence electrons (electrons in the outermost shells) of the elements in the same period increases gradually by one, across the period from left to right. This leads to a parallel development of properties for elements in each period. For example, the metallic properties tend to decrease across each period while the non-metallic properties tend to increase.

Scientific Information

Representative Elements: All s-block and p-block elements are called representative elements. **Transition elements:** Those elements which show variable valency are called transition elements. **Lanthanides Series:** Lanthanides series are those elements which come after lanthanum in 6^{th} periods. The lanthanide series contain 14 elements, which ranges from $^{8}_{10}\partial\,\partial$ (Cerium) to $^{1}_{LS}Tn$ (Lutetium). Placed below the periodic table. Lanthanide series elements are also called rare earth elements.

Actinide series are those elements which come after Actinium in 7^{th} periods. The Actinide series contain 14 elements, which ranges from $^{232}_{90}Th$ (Thorium) to $_{103}Lr$ (Lawrencium), placed separately at the bottom of the table beneath Lanthanide series.

Practice Problem: 3.1

- *Which is the first element of the second period of the periodic table?*
- *Which one is the last element of the second period of the periodic table?*
- *How many elements are in that period?*

3.1.2 Groups

The vertical columns of elements in the periodic table are called Groups. They are numbered from I to VIII.

Relationship between Group Number and Electronic Structure (Top to Bottom)

The periodic table is arranged in such a manner that the elements in the same group have the same number of electrons in the outermost shell of their atoms i.e. they have the same number of valence electrons. It is clear from the fig. 3.2, that atoms having the same number of valence electrons are present in the same group.

*Because of the similarity in the number of valence electrons, the elements show resemblance in their chemical properties in the same group. Hence the first group elements namely, Lithium (Li), Sodium (Na), and Potassium (K) possess one valence electron each and they show similar chemical properties. These elements are very reactive, electropositive and show metallic character. They are also called **alkali metals or alkali family.***

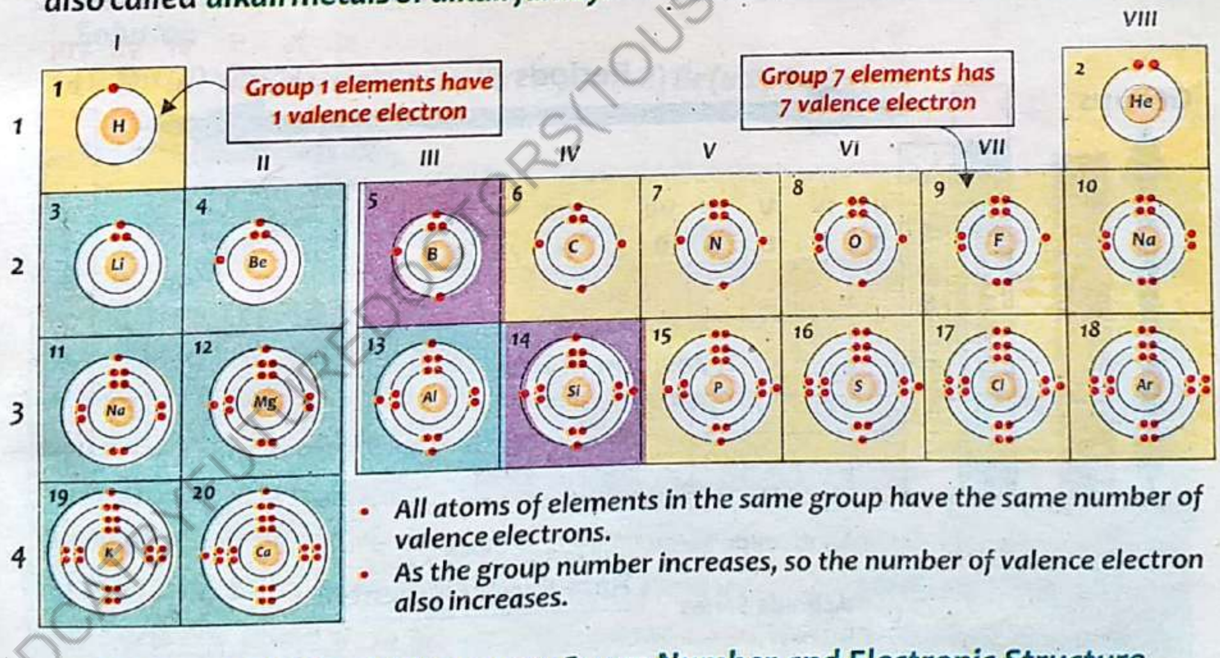


Fig. 3.2 Relationship between Group Number and Electronic Structure

Beryllium (Be), Magnesium (Mg) and Calcium (Ca) belong to second group. These elements are also called alkaline earth metals. They possess two valence electrons and show similar chemical behavior in the group. They are less electropositive than alkali metals.

Group III is called boron family. All the members of this family have three electrons in their outermost shell. Group IV is known as carbon family, having four outermost electrons in all family members. Group V is called nitrogen family and VI is called oxygen family, having five and six outermost electrons in all the members of the family respectively.

The seventh group is called halogens. They have seven electrons in their valence shell. They require only one electron to complete their valence shell. They are very electronegative. They show non-metallic character.

The eighth group is called noble gases or noble family. They have two or eight electrons in their valence shell. They possess a chemical inertness in their character due to the completion of their valence shell.

Based on the electronic sub – shell, the group – I and group – II elements are called s – block elements. As in these elements the 's' sub – shell are in the process of completion. Group – III to group – VIII elements are called p – block elements.

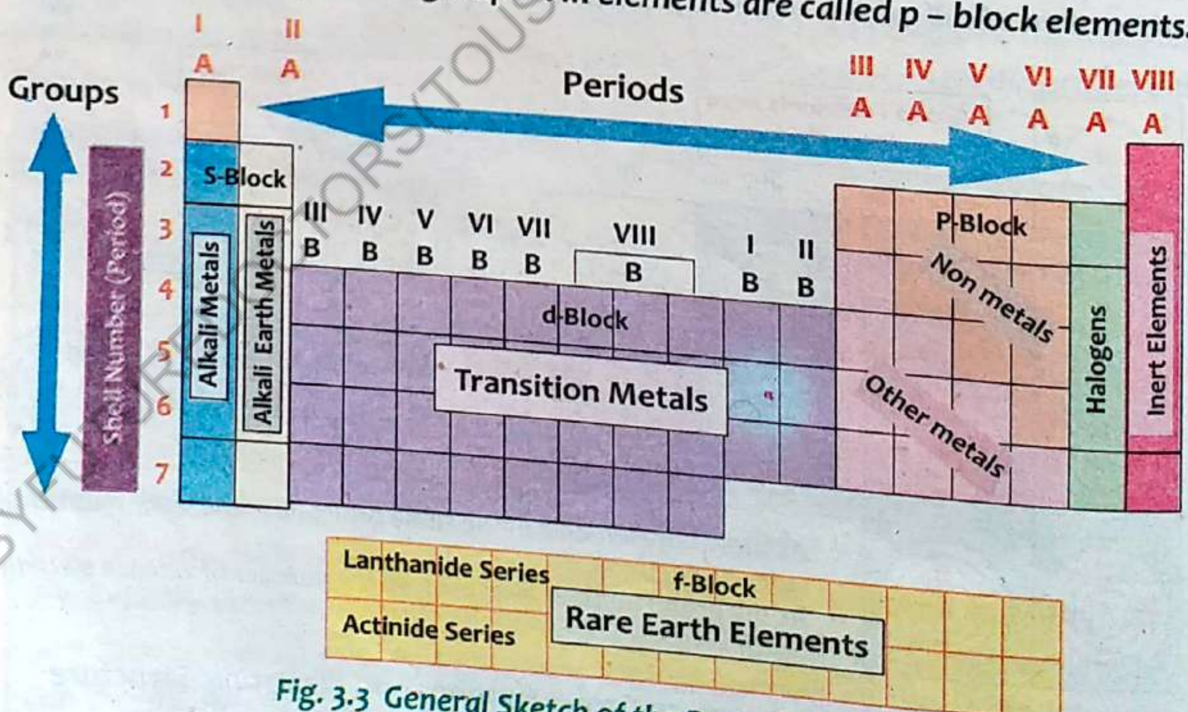


Fig. 3.3 General Sketch of the Periodic Table

As in these elements the 'p' sub – shell is in the process of completion. All s – block elements and p – block elements are called representative elements. The transition elements are called d – block elements, while the lanthanide and actinide series elements are called f – block elements. A general sketch of the periodic table is shown in figure 3.3.

EXAMPLE: 3.1

Use the periodic table to identify each of the following:
(i) The fifth element of the first transition series
(ii) The element of the fourth period that is also in group VB
(iii) The last lanthanide
(iv) The seventh transition element
(v) The second actinide metal
(vi) The first element of group VIII
(vii) The third halogen
(viii) The first alkaline earth metal
(ix) The first coinage metal

Solution

(i) Mn (ii) V (iii) Lu (iv) Co (v) Pa (vi) Fe (vii) Br (viii) Be (ix) Cu

Practice Problem: 3.2

Identify the second noble gas.

Scientific Information

Hydrogen is the lightest, most abundant and explosive gas on Earth. Fluorine is the most reactive and most electronegative of the elements, making elemental Fluorine a dangerously powerful Oxidant. This leads to direct reactions between Fluorine and most elements, including noble gases Krypton, Xenon and Radon. There are 17 gases in total, which can be found in the natural atmosphere on Earth. Only Oxygen and Nitrogen are found in large concentrations; 20.9476% and 78.084% respectively. Oxygen concentrations below 16% are considered unsafe for humans.

Scientific Information

- Group number indicates the number of electron in the outermost shell of an atom.
- Period number indicates the number of electron shells in an atom.

Test Yourself.

- To which group does the element Potassium belong?
- To which period does the element Arsenic belong?
- To which group does the element Titanium belong?
- State whether the following elements are metals or non-metals.
  Cesium, Boron, Selenium, Phosphorous, Astatine, and Osmium
- How many shells and valence electrons, the element Calcium will have?
- Element X is located in period III, group II of the periodic table, deduce its electronic configuration.
  - A. 2, 2
  - B. 2, 3
  - C. 2, 8, 2
  - D. 2, 8, 3

3.2 Periodicity of Properties

In a same group, all the elements will have similar general physical and chemical properties. It means after a certain interval the properties of the given elements are repeated. A pattern of repeating properties at regular interval is called periodicity. This is the repetition of the properties in the different groups and periods of the periodic table, when the elements are arranged in the increasing order of their atomic numbers.

The process /pattern by which there is repetition of properties in all the groups and the periods after a certain interval, are called the periodicity of properties.

Here we will discuss the periodicity of some physical properties of the elements in the periodic table.

3.2.1 Atomic Size

The size of an atom is not rigidly fixed but it varies when combined with different atoms. The same atom may have different sizes in different combination. It must be remembered that an atom in isolation has different size (atomic volume) than when it is in combination with other atoms. The atomic size of an atom is expressed in terms of Atomic radii, covalent radii and Ionic radii.

(i) Atomic Radius

The distance between the nucleus and the valence shell (outer shell) of the atom is termed as atomic radius. The atomic radius is represented by "r". The atomic radius is directly proportional with the number of shells. The atomic radius are expressed in the nanometer (1.0×10^{-9} m) or the Pico meter (pm = 1.0×10^{-12} m).

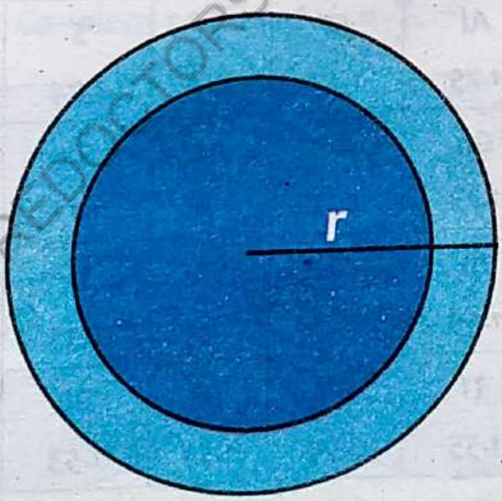


Fig. 3.4 Atomic Radius

Trends of Atomic Radius in the Periodic Table

(a) Trends of Atomic Radius in Groups

The atomic radius in periodic table increases in group from top to bottom, due to the addition of new shells in successive periods and shielding effect of electrons.

(b) Trends of Atomic Radius in Periods

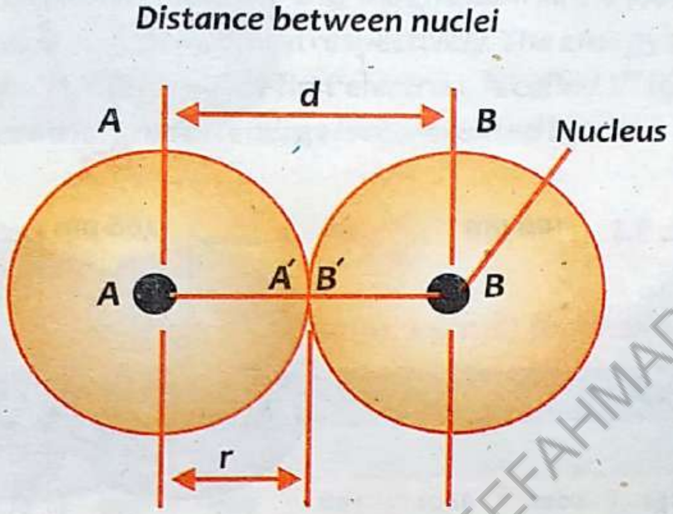
The atomic radius decreases in the period from left to right, due to the addition of electron in the same shell. As the number of electrons increases, the number of protons also increases, which increase the nuclear pull on the electrons and pulls them nearer to the nucleus. Thus, the atomic radius decreases.

Table 3.1 Atomic Radii in angstrom units (10^{-8} cm) of the representative elements

H 0.37							He 0.49
Li 1.23	Be 0.89	B 0.80	C 0.77	N 0.74	O 0.74	F 0.72	Ne 0.51
Na 1.35	Mg 1.36	Al 1.25	Si 1.17	P 1.10	S 1.04	Cl 0.99	Ar 0.88
K 2.03	Ca 1.74	Ga 1.25	Ge 1.22	As 1.21	Se 1.17	Br 1.14	Kr 1.03
Rb 2.16	Sr 1.91	In 1.59	Sn 1.41	Sb 1.41	Te 1.37	I 1.53	Xe 1.24
Cs 2.35	Ba 1.98	Tl 1.55	Pb 1.54	Bi 1.52	Po 1.53	At 1.63	Rn 1.34

(ii) Covalent Radii

The one half of the distance between the nuclei of two similar atoms of the same molecule containing a single covalent bond. Therefore, the bond distance between the two atoms 'A' and 'B' is the average of the lengths A – A' and B – B'.



Covalent radius (r) = $\frac{d}{2}$

Covalent radius of $H_2 = \frac{d}{2}$

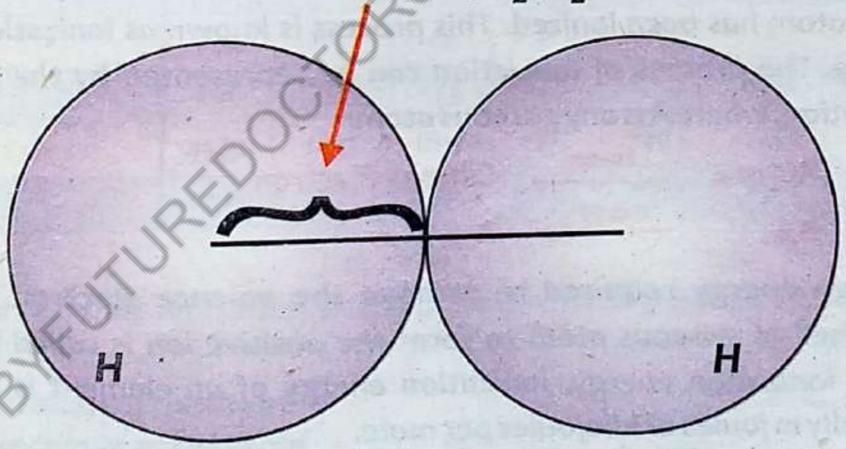


Fig. 3.5 Covalent Radius of Hydrogen molecules

Examples of halogens

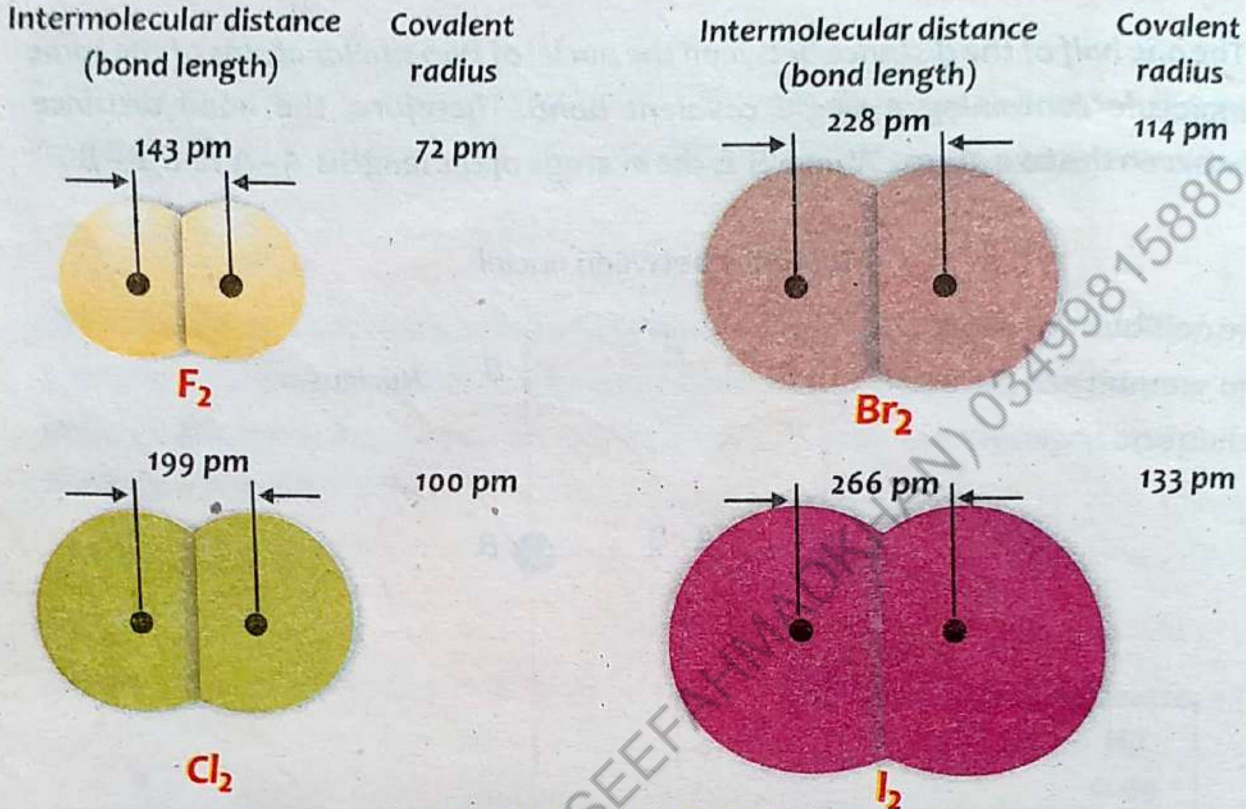


Fig. 3.6 Covalent Radii

3.2.2 Ionization Energy or Ionization Potential

When an electron is completely removed from the valence shell of an atom, we say that the atom has been ionized. This process is known as ionization, which needs energy. The process of ionization can be represented by the following general equation, where A is any gaseous atom:

$$\text{Atom}_{(g)} \xrightarrow{\text{Energy}} \text{Cation} + \text{Electron}$$

$$A_{(g)} \xrightarrow{\text{Energy}} A^{+}_{(g)} + e^{-}$$

The minimum energy required to remove the valence electron from the outermost shell of gaseous atom to form the positive ion is called ionization potential or ionization energy. Ionization energy of an element is measured experimentally in joules or kilojoules per mole.

For example, the first ionization energy of sodium and magnesium is represented by the equation:

$$Na_{(g)} \xrightarrow{\text{Energy}} Na^{+}_{(g)} + e^{-} \qquad I.E = +496kJ/mol$$

Similarly,

$$Mg_{(g)} \xrightarrow{\text{Energy}} Mg^{+}_{(g)} + e^{-} \qquad I.E = +738kJ/mol$$

The above two values for sodium and magnesium are called the 1st ionization potential for sodium and magnesium respectively. The energy required to remove a 2nd electron, after the removal of first electron, is called 2nd ionization potential. Magnesium's second ionization energy is represented by

$$Mg^{+}_{(g)} \xrightarrow{\text{Energy}} Mg^{+2}_{(g)} + e^{-} \qquad I.E = +1450kj/mol$$

Table 3.2 ionization energies (in. KJ/mol) for elements of Period 1 -3

	Period 1		Period 2							
	H	He	Li	Be	B	C	N	O	F	Ne
IE_1	1312	3272	520	900	801	1086	1402	1314	1681	2081
IE_2		5250	7298	1757	2427	2353	2856	3388	3374	3952
IE_3			11815	14849	3660	4621	4578	5300	6050	6122
			Period 3							
			Na	Mg	Al	Si	P	S	Cl	Ar
IE_1			496	738	578	787	1010	1000	1251	1521
IE_2			4562	1451	1817	1577	1903	2251	2297	2666
IE_3			6912	7733	2745	3232	2912	3361	3822	3931

It is observed that ionization energies of atoms depend upon several factors. These are,

(i) Atomic radius of the atom
(ii) Nuclear charge of the atom
(iii) Shielding effect of low lying electrons
(iv) Electronic configuration of the atom

Trends of Ionization Potential in the Periodic Table

(a) Trends of Ionization Potential in Groups

The ionization potential in a group decreases as we move down the group because of increase of atomic size. The elements at the top of the Group have maximum ionization potential and the elements at the bottom of the group have minimum ionization potential.

(b) Trends of Ionization Potential in Period

The ionization potential increases in a period in the periodic table from left to right. The elements, which are on the extreme right of the period, has maximum ionization potential and the elements, which are on the extreme left of the period, has minimum ionization potential.

Table 3.3 First ionization energies (kJ/mol) of the representative elements

H 1312							He 2372
Li 520	Be 899	B 801	C 1086	N 1402	O 1314	F 1681	Ne 2081
Na 496	Mg 738	Al 578	Si 786	P 1012	S 1000	Cl 1251	Ar 1521
K 419	Ca 590	Ga 579	Ge 762	As 947	Se 941	Br 1140	Kr 1351
Rb 403	Sr 549	In 558	Sn 709	Sb 834	Te 869	I 1008	Xe 1170
Cs 376	Ba 502	Tl 589	Pb 716	Bi 703	Po 812	At 1002	Rn 1038

EXAMPLE: 3.2

Which atom in each of the following sets has the largest ionization energy?

(a) K, Ga, Se (b) O, S, Se (c) In, As, Cl

Solution

(a) Se (The ionization energy gets larger as we go to the right in the periodic table.)

(b) O (The ionization energy gets smaller as we go down the group.)

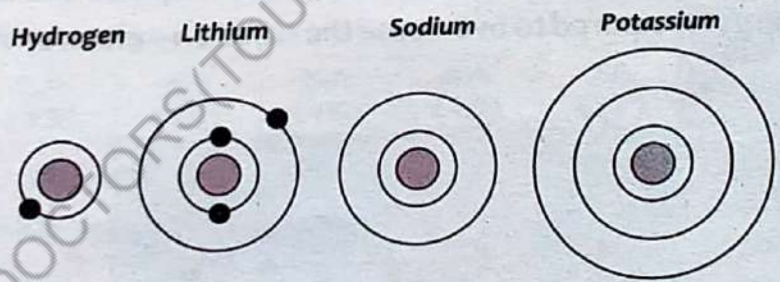
(c) Cl (The ionization energy gets larger as we go to the right and as we go up.)

Practice Problem: 3.3

- Which atom in each of the following sets has the largest ionization energy?

(a) Be, C, N (b) Mg, Ca, Sr (c) Sn, Se, F

Test Yourself.



Electron diagrams for same of the Group-I elements

- Use a periodic table, to complete the last two diagram for sodium (Na) and potassium (K).
- What do you notice about the number of electrons in the valence energy level in each case?
- Explain why elements from group-I are more reactive than elements from group-II of the periodic table (Hint: Think back to 'ionization energy').

3.2.3 Electron Affinity (E.A)

Electron affinity means love or attraction for accepting electron. All these elements in the periodic table possess varying tendencies towards the acceptance of electrons in their outer shell. The new incoming electron when absorbed by the atom is slightly bound by the nucleus of the atom through attractive force. This causes an evolution of energy. Thus, the electron affinity of the atom is measured in terms of energy.

The minimum amount of energy released when an electron is added to gaseous atom of an element in its outermost shell to form an anion is called electron affinity. It is represented by E.A. It is assigned negative values and expressed in $KJ.mol^{-1}$. For example,

$$\underset{(2,8,7)}{Cl} + e^- \longrightarrow \underset{(2,8,8)}{Cl^{-1}} \quad E.A = -348.8 KJ/mol$$

The energy released by the addition of the first electron is called the first electron affinity; the energy released by the addition of the second electron is called the second electron affinity. By addition of the first electron, the energy is released, but the addition of the second electron requires the absorption of energy. This energy is required to overcome the electron – electron repulsion. For example,

$$\underset{(2,6)}{O} + e^- \longrightarrow \underset{(2,7)}{O^{-1}} \quad E.A = -141.0 KJ/mol$$

$$\underset{(2,7)}{O^-} + e^- \longrightarrow \underset{(2,8)}{O^{-2}} \quad E.A = +844.0 KJ/mol$$

Factors Affecting the Electron Affinity

Following factors affect electron affinity.

(i) Nuclear Charge

(ii) Atomic Radius

(iii) Shielding Effect

(iv) Electronic configuration

Atoms with smaller atomic radii, greater nuclear charge and poor shielding effect have usually high electron affinity values.

Trends of Electrons Affinity values in the Periodic Table

(a) The Electron Affinity in Groups

When we move in a group from top to bottom in the periodic table, the electron affinity decreases. The decrease in the electron affinity is due to the addition of new shells. These shells make a shield between the nucleus and outer most shell and force of attraction between the nucleus and valence electrons is reduced. Therefore, the tendency of taking electron is decreased.

(b) The Electron Affinity in Periods

When we move in a period from left to right, the electron affinity increases. This increase in the electron affinity is due to increase in the nuclear charge and decrease in the atomic radius from left to right. Therefore, the tendency of taking electron is increased.

Table 3.4 Trends of electron affinity

1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.8							He (+21)
Li −59.6	Be (+241)	B −26.7	C −122	N 0	O −141	F −328	Ne (+29)
Na −52.9	Mg (+230)	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+34)
K −48.4	Ca (+156)	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr (+167)	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+40)
Cs −45.5	Ba (+52)	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

3.2.4 Shielding Effect

Electrons in atom are distributed in different shells or orbits. They are revolving around the nucleus. The electrons in the valence shell spend most of the time away from the nucleus than the inner orbital electrons because these electrons do not feel the pulling effect of the positive charge of the nucleus. As the number of shells increases between the nucleus and the valence shell, shielding effect will be increased and there will be more protection of the electrons from the attraction of the nucleus. So, the removal of the electrons from the outer most shell is easy. Therefore, shielding effect has a direct impact on the atomic radii, ionization potential and electron affinities of the elements.

The inner electrons shield the outer electrons from the nuclear charge and reduce the hold of the nucleus on these valance electrons. This effect is called shielding effect or screening effect."

Trends on shielding effect in the periodic table

(a) Shielding effect in groups

When we move in a group from top to bottom, the shielding effect are more effective than the nuclear charge. This is because a new shell is added as we move down in the group. This shell screens the valance electrons from the hold of the nucleus. This causes a decrease in the ionization energies and electron affinities of the elements down the group.

(b) Shielding effect in periods

When we move in a period from left to right, the atomic number increases. The

Fig. 3.7 Shielding effect trend in first group

positive charge on the nucleus also increases. There is no change in inner orbits, shielding electrons. Consequently, the increasing nuclear charge wins over shielding effect, which remains constant.

3.2.5 Electronegativity

Electronegativity is a property associated with the atoms when they are chemically bonded to each other. In a covalent bond, the two atoms involve in bond formation, mutually share their electrons. This shared pair of electrons is then attracted by the nuclei of both the atoms. But different atoms have different abilities to attract the shared pair of electrons of the covalent bond. The ability of an atom to attract the shared pair of electrons towards itself in a covalent bond is called electronegativity. In other words, we can say that the power of attraction for the shared pair of electron is called electronegativity. Electronegativity is represented by E.N. If the two atoms have the same ability to attract the shared pair of electron; we say that they have the same electronegativity. In such a case the covalent bond between them is said to be a non-polar covalent bond e.g. H_2. On the other hand, if the bond is formed between atoms of different electronegativity, it is said to be a polar covalent bond e.g. HCl.

Factor Affecting the Electronegativity

Following factors affect electronegativity of elements.

(i) Nuclear Charge: Greater the nuclear charge, greater will be the electro-negativity. When we move in period from left to right, the nuclear charge increases, so the electronegativity increases.

(ii) Atomic Radius: Greater the atomic radius, lower will be the electronegativity. When we move in a group from top to bottom, the atomic radius increase, so electronegativity decreases.

(iii) Shielding Effect: As the number of shells increases between the nucleus and the valence shell, so the removal of the electrons from the outer most shell is easy. Therefore, the electronegativity will be lower.

(iv) Electronic Configuration: The completely filled orbitals and half-filled orbitals are stable. The addition of electrons to these orbitals is difficult.

Therefore, their electronegativity is very low.

Beside this the electronegativity values also depends to some extent on

- the nature of the combining atoms
- the atomic volume of the combining atoms
- the value of electron affinity and
- the value of Ionization potential.

Pauling calculated the electronegativities values of the elements. For this purpose, he developed a scale from bond energies of diatomic molecules. On this scale the electronegativity of Fluorine is 4.0 which is the highest of all the periodic table elements, Oxygen is 3.5 and the electronegativity of Cesium is 0.7, which is the lowest.

Test Yourself.

- The least electronegative element is Cs (0.7), can you guess where is it located in the table?
- Can you tell, elements of which group have the highest electronegativity values?

Trends in Electronegativity values in the Periodic Table:

(a) Electronegativity in the Groups

In groups, the electronegativity of the elements at the top of the group have maximum, while the elements at the bottom of the group have minimum values. This is due to the addition of shells and increase in atomic radius of the electrons. The new inner shell increases the shielding effect and decrease the electronegativity.

(b) Electronegativity in the Periods

In the periods, the electronegativity increases as we move in the periods from left to right. The elements at the left of the period have minimum electronegativity, while the elements at the extreme right of period, have maximum electronegativity. This may be attributed to the decreasing atomic size and increases nuclear charge due to the addition of electrons in the same shell.

It can be concluded from the above discussion that the elements located in the upper right corner of the table will be the most electronegative elements.

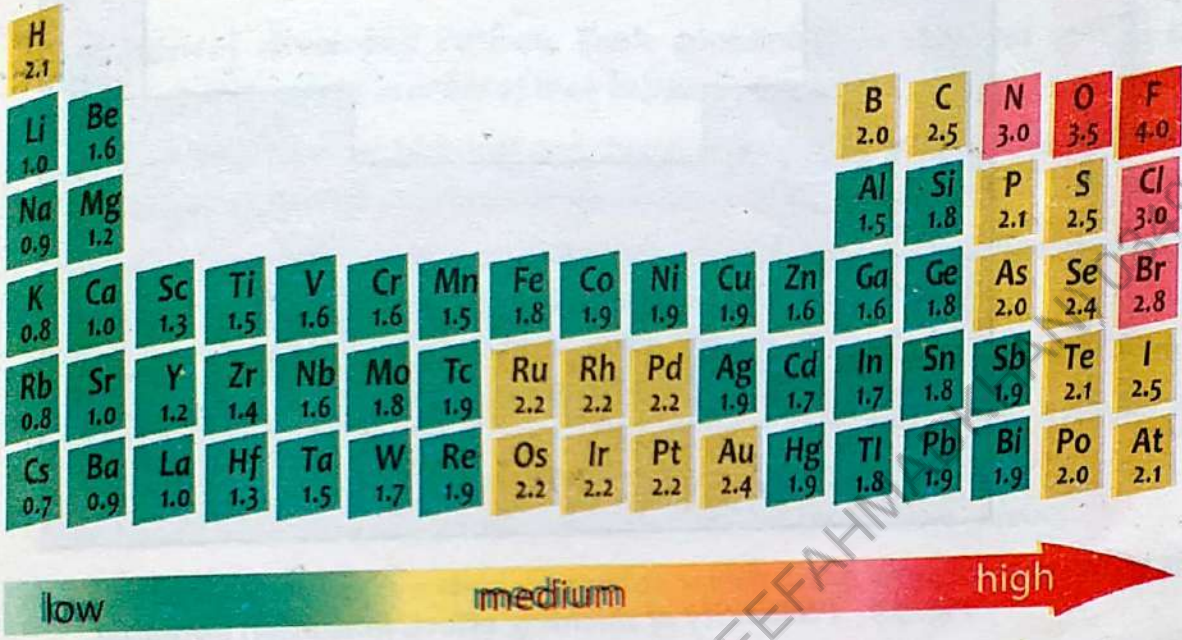


Fig. 3.8. Electronegativity trends of elements in the periodic table

EXAMPLE: 3.3

Which set has the more similar chemical properties?
(a) N, P, As (b) Li, C, F

Solution
Set (b) is in the same period (2); set (a) is in the same group (VA). Set (a) has the more similar chemical properties because these elements are in the same group.

Practice Problem: 3.4

Is sulphur more likely to be similar to selenium (Se) or to chlorine in its chemical properties?

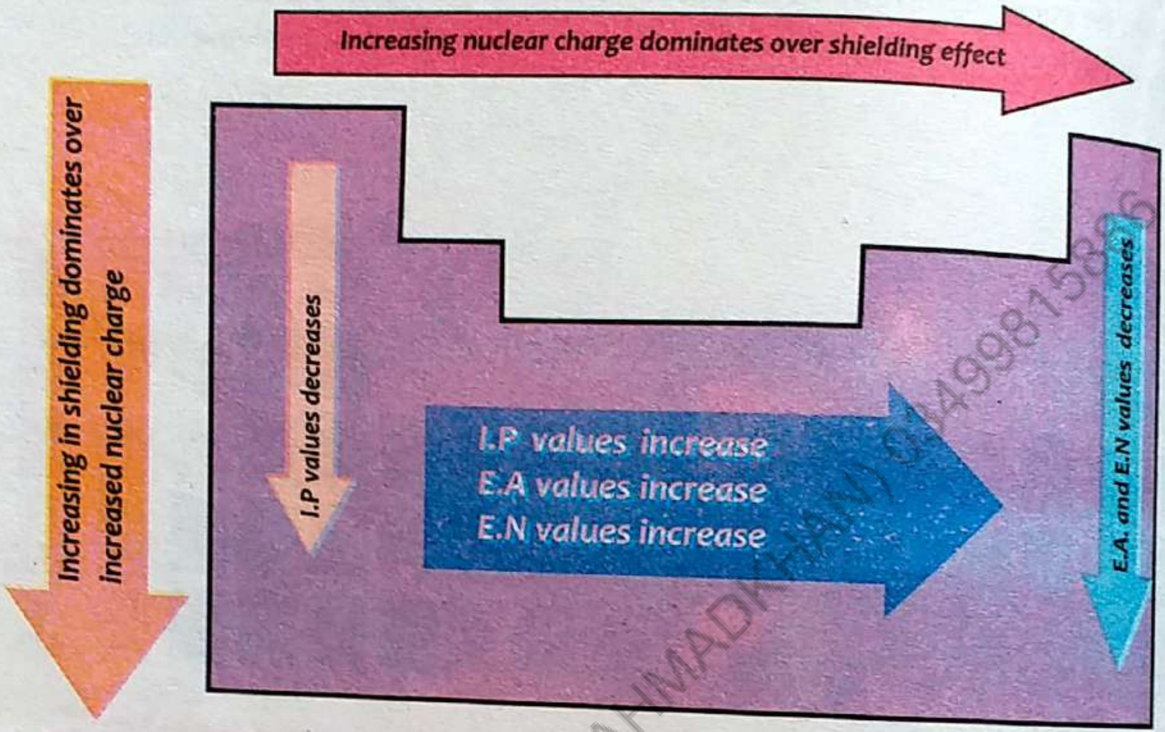


Fig. 3.9 Periodicity of shielding effect, ionization potential, electron affinity and electronegativity

Key Points

- Mendeleev developed Periodic Table containing periods and groups by arranging elements in order of their increasing atomic weights.
- Old periodic law, the physical and chemical properties of the elements is the periodic function of their atomic weights.
- Modern periodic law, the physical and chemical properties of the elements is the periodic function of their atomic number.
- There are total eight groups and seven periods in the modern Periodic Table.
- Depending on outermost electrons and electronic configuration, element in periodic table is grouped in s, p, d and f blocks elements.
- Transition elements are those elements which show variable valency.
- Lanthanides series are those elements which come after lanthanum in 6^{th} period.
- Actinide series are those elements which come after Actinium in 7^{th} period.
- The properties, which are repeated in all the groups and the periods after a certain interval, are called the periodic properties and this repetition of properties is called periodicity of properties.
- The distance between the nucleus and the valence shell of the atom is termed as atomic radius
- The minimum energy required to remove the valence electron from the outermost shell of gaseous atom to form the positive ion is called ionization potential or ionization energy.
- The minimum amount of energy released when an electron is added to gaseous atom of an element in its outermost shell to form an anion is called electron affinity.
- Shielding effect is greater in atoms with greater number of electrons.
- Electronegativity increases along a period and decreases down the group.

Exercise

Choose the correct option.

(i) Which of the following elements is in the same family as fluorine:
(a) silicon (b) antimony
(c) iodine (d) arsenic

(ii) Which of the following would have the smallest ionization energy:
(a) K (b) P
(c) S (d) Ca

(iii) An element has configuration 2, 8, 1. It belongs to, ________:
(a) group I and III period (b) group III and I period
(c) group I and VII period (d) group VII and III period

(iv) Which of the following elements would be most similar to carbon:
(a) nitrogen (b) boron
(c) oxygen (d) silicon

(v) s – block elements are:
(a) metals (b) non-metals
(c) metalloids (d) transition

(vi) Which of the following would have the largest ionization energy:
(a) Na (b) Al
(c) H (d) He

(vii) Elements in a ________ have similar chemical properties:
(a) period (b) group
(c) both a and b (d) neither a nor b

(viii) An element has 8 electrons in its valence shell. It is a member of:
(a) alkali family (b) halogen family
(c) noble family (d) carbon family

(ix) The modern periodic table is based on:
(a) atomic number (b) mass number
(c) neutron number (d) isotope number

(x) Shielding effect is due to:
(a) neutron (b) Proton
(c) proton and neutron (d) electron

SHORT QUESTIONS

Answer briefly the following questions.

(i) Which element of group IA is not an alkali metal and why?

(ii) Place the following elements in order of increasing ionization energy: Na, S, Mg and Ar.

(iii) Name the group and state the group number of each of the following elements:

(a) K (b) Ne (c) Be (d) Cl (e) C

(iv) Which element is the most electronegative among C, N, O, Br and S? Which group does it belongs to?

(v) How do the first ionization energies of representative elements vary across a period and down a group?

(vi) Which element is found in,

(a) Period 2, group VII (b) Period 4, group III

(c) Period 5, group VI (d) Period 1, group VIII

(vii) How will you differentiate between representative and transition elements?

(viii) Make a general sketch of the periodic table showing s, p, d and f-block elements (without showing the symbols of elements)

(ix) Why the s-block elements have two groups only?

(x) What type of element is Sulphur (S), a representative element, a transition element or lanthinide element?

LONG QUESTIONS

(i) How modern periodic table is different from the Mendeleev's periodic table.

(ii) Differentiate between atomic radii and covalent radii. Explain the trends of atomic radius in the periodic table.

(iii) What is electronegativity? Identify the most and least electronegative groups of elements in the periodic table. Why is fluorine special in terms of electronegativity?

(iv) Define shielding effect and draw it affects the ionization energy, electron affinity and electronegativity.

(v) Explain the following terms.

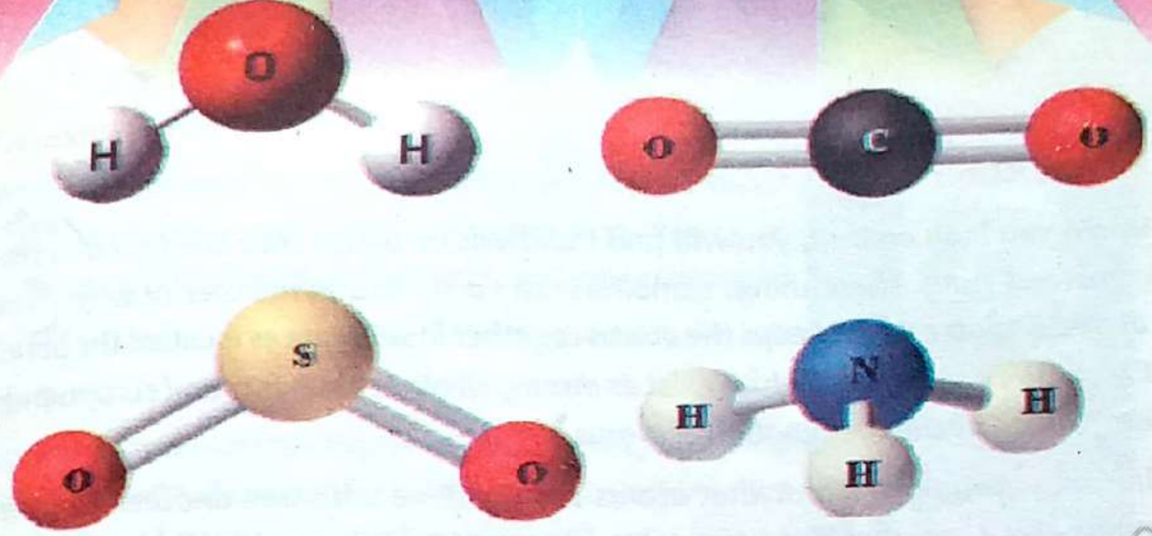
(a) Periodicity of properties

(b) Electron affinity

(c) Modern periodic law

Project Work

Prepare a report tracing the evolution of the current periodic table since 1900. Cite the chemists involved and their major contributions.



Unit 4 Structure of Molecules

After studying this unit, the students will be able to;

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

Introduction

Ans.

When you look around, you will find hundreds of things that are found in the combined state. These things combine with each other in number of ways. (① The attractive force which keeps the atoms together in substances is called the bond. Except the noble gases, which exist as atoms, all other elements and compounds are formed by the combination of atoms.) End ①

In unit 1, you have learnt that atoms can combine with one another to form molecules. Molecules of the same elements consist of the same kind of atoms e.g. O_2, O_3, while molecules of compounds are made up of different kinds of atoms e.g. H_2O , CO_2 etc. Atoms combine with one another by the process called chemical bonding. Why do atoms form bonds? In how many ways do they form bonds? What are the differences between the different kinds of bond? You will get answers of these questions in this unit.

4.1 Why do atoms form Chemical Bonds?

Every system in universe tends to lower its energy, in order to attain stability. Water flows from higher level to lower level. Similarly, electricity flows from higher potential to lower potential and heat flows from a hot body to a cold body. Why is it so? This happens because both water and electricity are trying to decrease their energy. Ans. ② (Atoms, ~~in the same way~~ have tendency to decrease their energy. The energy of the isolated hydrogen atoms is higher than the two bonded hydrogen atoms. That is, the combination of atoms gives stable molecule through emission of energy. They can decrease their energy by combining with other atoms and form a chemical bond.

There are two concepts, which explain the chemical bonding,

i. The valence concept and
ii. The Orbital concept) End ②

Ans ③ 4.1.1 (The Valence Concept (Electronic Theory of Valence)

In 1916, G. N. Lewis and W. Kossel gave the electronic theory of valence.

It states that in a chemical bond formation, atoms take part by losing, gaining or sharing of electrons, so as to attain the inert or noble gas electronic configuration.

G. N. Lewis

When atoms have two or eight electrons in their outer most shell, they are stable. The electron theory of valence can be named as the Octet or Duplet theory of valence.

(a) Octet Theory of Valence or Rule of Eight

Octet theory of valence explains the tendency of atoms to attain eight electrons in the outer most shell in order to attain stability. For example oxygen (O) atom has six electrons in their valence shell. It shares or gains two electrons in its outer most shell and attain the stability by completing its outermost shell with eight electron. This is also known as the Octet rule or rule of eight.

(b) Duplet Rule or Rule of two

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The tendency of atoms to attain two electrons in the valence shell in order to attain stability. For Example, Helium (He) has two electrons in its valence shell and is stable.

The elements in groups VIII of the periodic table, such as helium, neon and argon are known as noble gases. They are also called inert gases because they are very

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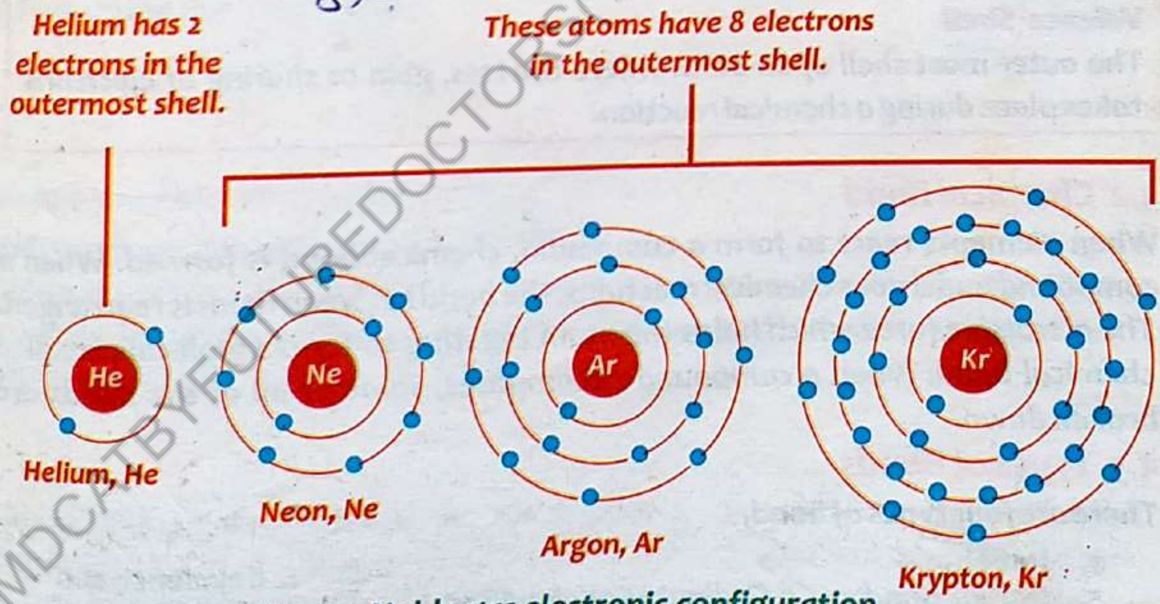


Fig. 4.1 Noble gas electronic configuration

stable and rarely take part in chemical reactions to form compounds. Their stability comes from their completely filled outermost shells.

Except for helium that has two electrons in the outermost shell, all other noble gases have their outermost shells filled with eight electrons. A shell with eight electrons is called an octet and is very stable. Thus, when atoms take part in chemical reactions, they tend to combine in ways to complete eight electrons in their outermost shell, to attain the electronic configurations of the noble gases (except helium).

(3) S, 6

4.1.2 The Orbital Concept:

This concept is based on the combination of atomic orbital to produce molecular orbital. The atomic orbitals have one electron. These orbitals when come close to one another, they overlap each other. This overlapping is either endwise or sidewise. Endwise overlapping produce sigma bond and sidewise overlapping produce pi-bond. End (3) End

Scientific Information

Valence Electrons

The electrons present in the valence shell of an atom are called the valence electrons. The electrons which take part in bond formation is called bonding electrons and the electrons which do not take part in the bonding are called the non bonding electrons.

Valence Shell

The outer most shell of an atom where the loss, gain or sharing of electrons takes place during a chemical reaction.

4.2 Chemical Bond

When elements react to form a compound, chemical bond is formed. When a compound undergoes chemical reactions, the bond between them is rearranged. The attractive force which holds the atom together to form a molecule is called chemical bond. When a compound decomposes, some or all of the bonds are broken down.

4.3 Types of Bonds

There are four types of bond,

1. Ionic bond
2. Covalent bond
3. Dative Bond or Co-Ordinate Covalent Bond
4. Metallic bond

(4)

Ans 4.3.1 Ionic Bond

The chemical bond which is formed due to the complete transfer of electron(s) from one atom to the other atom is called ionic bond or electrovalent bond. The transfer of electrons between atoms complete the octets and duplets. This type of bonding is always between a metal and non-metal. Metals always lose electrons to form cation and non-metals always gain electrons to form anions.

In ionic bond formation, one atom loses electron(s) and the other gains it. The atom that loses the electron acquires positive charge and the other atom, which gains the electron, becomes a negatively charged particle. Due to opposite charges, an electrostatic force of attraction is setup between them. This force holds these ions together. This force of attraction is referred as an ionic bond. S.Q

An atom having low ionization energy will lose one or more electrons to form cation. An atom having high electronegativity will gain one or more electrons to form anion. For Example -

(i) Formation of Sodium Chloride (NaCl)

Sodium (Na) and Chlorine (Cl) combine to give Sodium Chloride (NaCl).

The atomic number of Sodium (Na) is '11'.

Its electronic configuration is K = 2, L = 8, M = 1

It is clear that Sodium has one electron in its outer most shell. It has the tendency to lose one electron to attain the inert gas electronic configuration of Neon. So the Sodium ion (Na^+) is formed.

$$Na_{2,8,1} \longrightarrow Na^{+}_{2,8} + e^{-}$$

The atomic number of Chlorine (Cl) is '17'

Its electronic configuration is K = 2, L = 8, M = 7

It is clear that Chlorine (Cl) needs one electron to complete its outer most shell. It has the tendency to gain one electron to attain inert gas electronic configuration of Argon. So it forms Chloride ion (Cl^-).

$$Cl_{2,8,7} + e^{-} \longrightarrow Cl^{-}_{2,8,8}$$

We know that opposite charges attract each other, therefore, an electrostatic force of attraction is set up between Na^+ and Cl^- ions. This force unites these ions in a crystalline lattice. In this way, an ionic bond is formed between Na^+ and Cl^- ions, which results in the formation of sodium chloride. End (4)

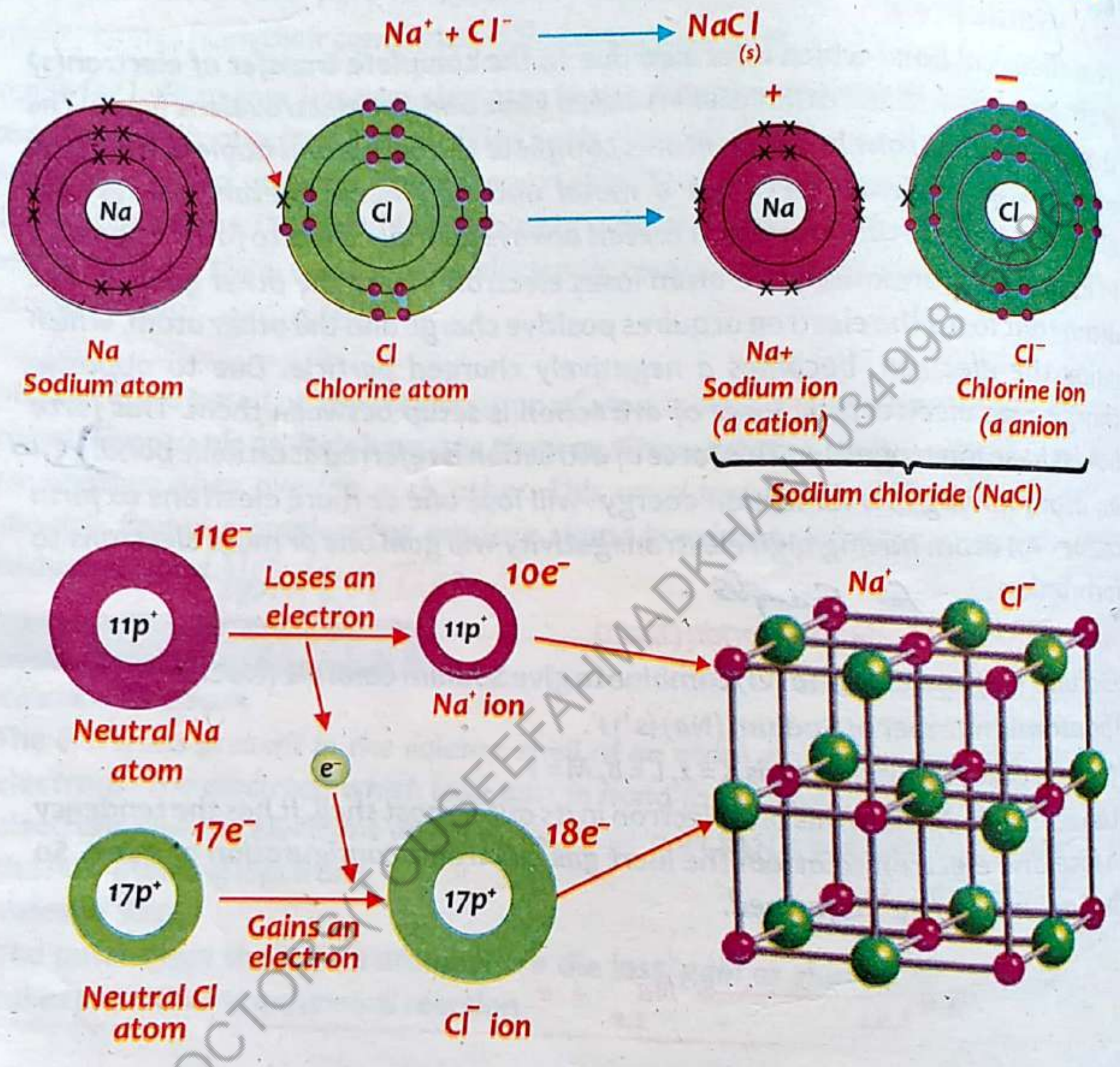


Fig. 4.2 Formation of NaCl compound

(ii) Formation of Calcium Chloride ($CaCl_2$)

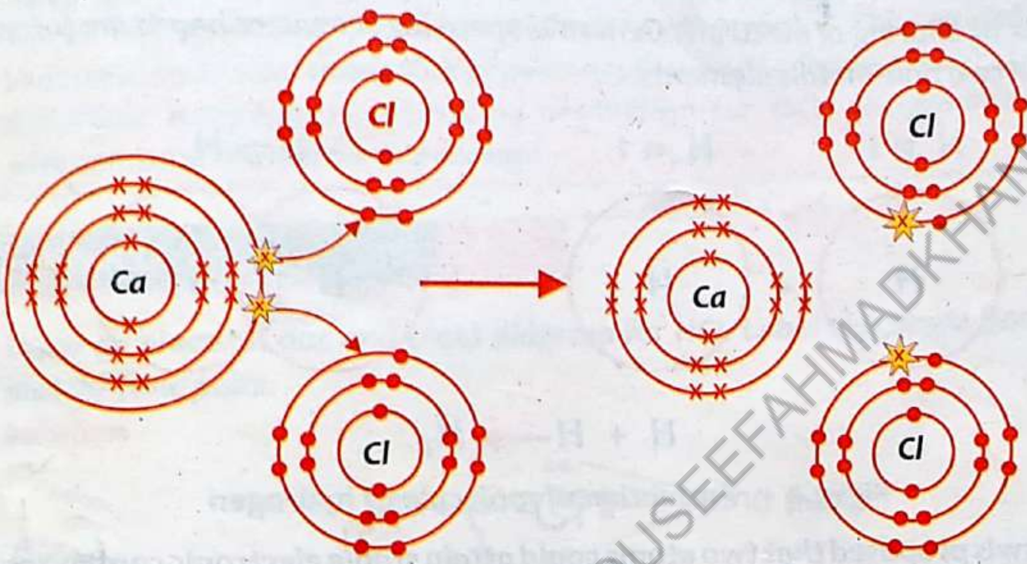
Calcium (Ca) and Chlorine (Cl) combine to give Calcium Chloride ($CaCl_2$). Calcium has two electrons in its outer most shell. It has the tendency to lose two electrons to attain the inert gas electronic configuration of Neon. So the Calcium ion (Ca^+) is formed.

$$Ca_{2,8,8,2} \longrightarrow Ca^{+2}_{2,8,8} + 2e^-$$

Chlorine (Cl) needs one electron to complete its outer most shell. It has the tendency to gain one electron to attain inert gas electronic configuration of Argon. So it form Chloride ion (Cl^-).

$$Cl_{2,8,7} + e^- \longrightarrow Cl^-_{2,8,8}$$

One Ca^{+2} and two chloride ion (Cl^-), which results in the formation of Calcium chloride ($CaCl_2$).



Ca + 2 Cl ⟶ $CaCl_2$

Fig. 4.3 Formation of $CaCl_2$

This explains the divalent nature of calcium (Ca), similarly, we can show the trivalent nature of Aluminum (Al) atom in the formation of $AlCl_3$ molecule.

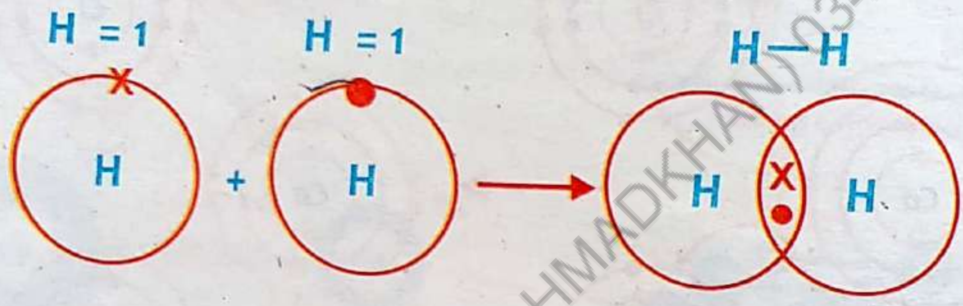
Test Yourself.

1. a. In what ways do atoms form ions?
   b. Why do atoms form ions?
2. Ionic bonds are formed between which two types of elements.
3. State the formula of the ions formed by the following atoms
   a. Calcium (2,8,8,2) b. Fluorine (2,7) c. Lithium (2,1) and d. Oxygen (2,6)
4. Draw a diagram to show the ionic bonds in magnesium oxide (MgO) and Magnesium chloride ($MgCl_2$).

Ans 5

4.3.2 Covalent Bond

We have learnt that molecules of elements are made up of the same kind of atoms. An example is a molecule of hydrogen, made up of two hydrogen atoms and its formula is H_2. How are the two hydrogen atoms joined or bonded? Since the two atoms are identical, we do not expect one of them to transfer its electron to the other. Instead of transferring electrons, the two atoms form bonds by sharing of electrons. Ans. 5 A covalent bond is formed when two atoms are joined together by sharing of electrons. Generally speaking, covalent bonds are formed between two non-metals elements.



$H + H \rightarrow H_2$

Fig. 4.4 Formation of molecule of hydrogen.

G. N. Lewis proposed that two atoms could attain stable electronic configuration by sharing electrons between them. The shared pair of electrons is represented by a dash (–) between the two bonded atoms. The shared pair of electrons remains between the two bonded atoms and are called the localized electrons. The bond which is formed by the mutual sharing of electrons is called covalent bond. For example, Hydrogen molecule (H_2) Oxygen molecule (O_2), hydrochloric acid (HCl) etc.

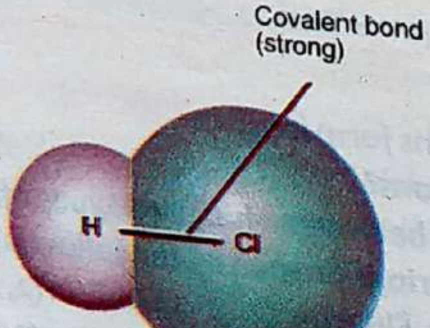


Fig. 4.5 Formation of molecule of HBr

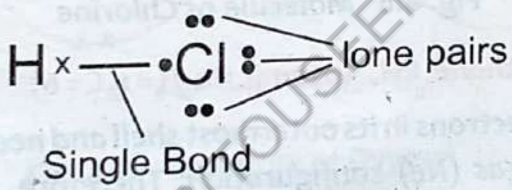
Society, Technology and Science

Glue is a substance that allows two surfaces to be bonded together. The term is commonly used interchangeably with "adhesive". An adhesive holds materials together. There are different types of adhesive e.g. cement, glue etc. There are many natural and synthetic adhesives in used. Adhesives can also be used as coatings on the surfaces that are subjected to corrosion and rust. Adhesives are also used in many repair applications. These are commonly used to fix broken dishes and to make other repairs that would be impossible or difficult by other means. This can also provide improved mechanical strength. A popular application of adhesives is to encapsulate electronic components, providing protection for that component against environmental and mechanical damage.

EXAMPLE: 4.1

Draw an electron dot and cross diagram for HCl. Label the single bond pair and the lone pairs.

Solution



Practice Problem: 4.1

Draw an electron dot and cross diagram for H_2O. Label the single bond and the lone pairs.

4.3.2.1 Types of Covalent Bond

The covalent bond is further divided into three types.

i. Single Covalent Bond:

The bonds in which two atoms share one electron each to form a pair of electrons is called single covalent bond. A single straight line (–) shows the single covalent bond. For example, H_2, Cl_2, F_2 etc.

(a) Chlorine (Cl_2)

Chlorine molecule is formed from two Chlorine atoms. The Chlorine atom electronic configuration is (2, 8, 7). A Chlorine atom has seven electrons in its valence shell. The two Chlorine atoms mutually shares one electron with each other to form chlorine molecule (Cl_2) . Therefore, both Chlorine atoms attain inert gas (Ar) electronic configuration and complete their octet.

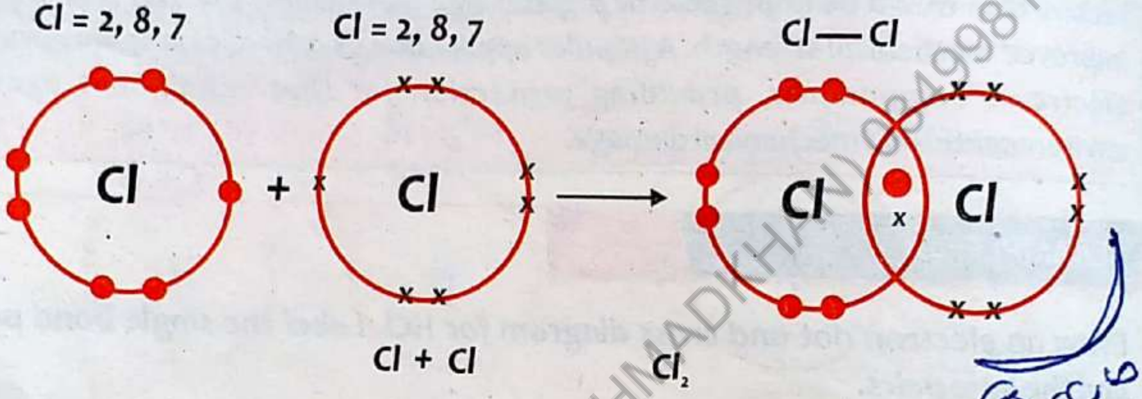
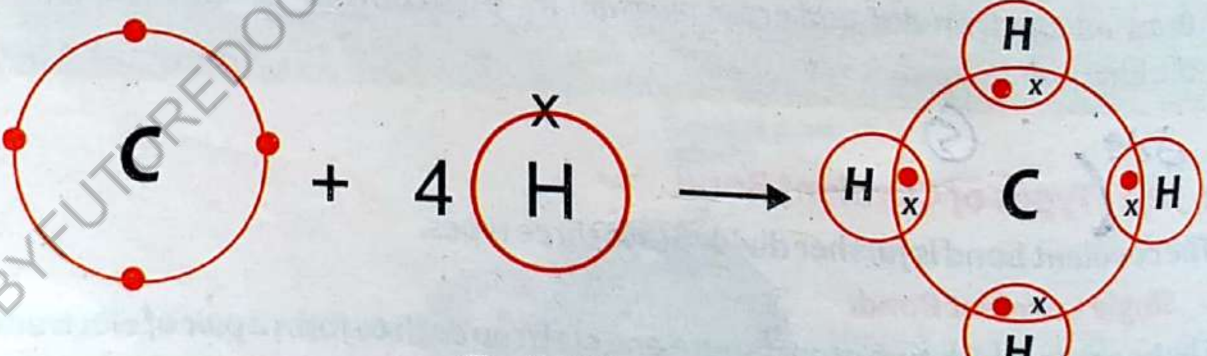


Fig. 4.6 Molecule of Chlorine

(b) Molecule of Methane, CH_4 (atomic No. H = 1, C = 6)

Carbon has four electrons in its outermost shell and needs four more electrons to attain the noble gas (Ne) configuration. Therefore, four atoms of hydrogen mutually share one electron each with a carbon atom to form a molecule of methane.



$$C_{(s)} + 2H_{2(g)} \longrightarrow CH_{4(g)}$$

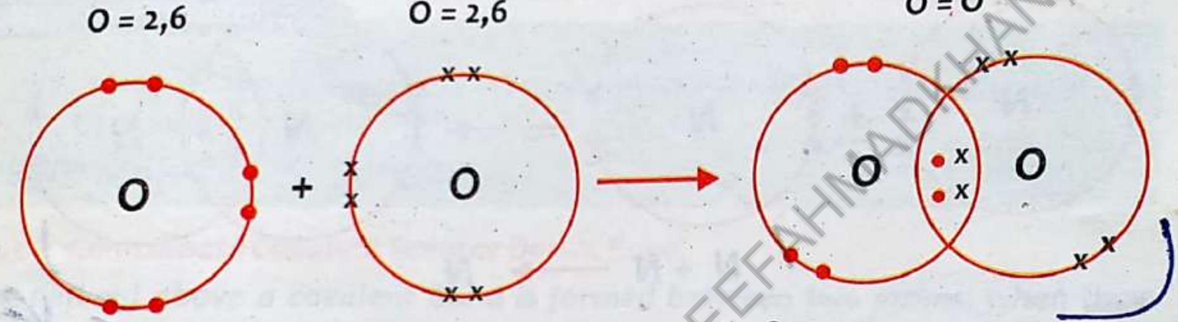
Fig. 4.7 Molecule of Methane

ii. Double covalent bond

The bond in which two atoms share two electrons each to form two pairs of electrons is called double covalent bond. A double straight line (=) shows such a covalent bond. For example, O_2, C_2H_4, CO_2, etc.

(a) Oxygen (O_2)

Oxygen molecule is formed from two Oxygen atoms. The Oxygen atom electronic configuration is (2, 6). An Oxygen atom has six electrons in its valence shell and it shares two electrons with another Oxygen atom to form oxygen molecule (O_2). In this way, both Oxygen atoms attain inert gas (Ne) electronic configuration and complete their octet.

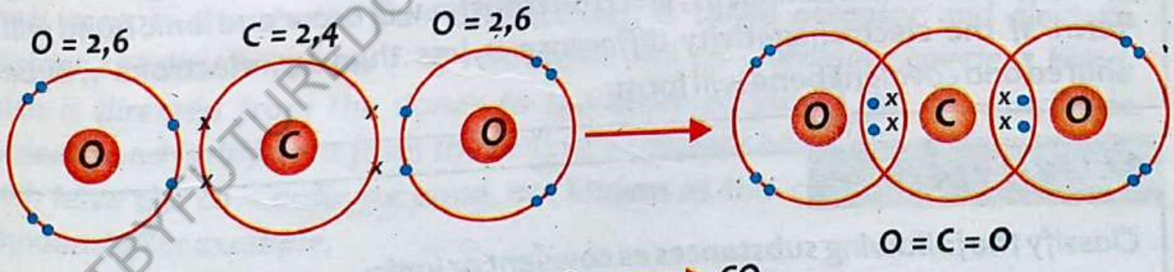


$O + O \longrightarrow O_2$

Fig. 4.8 Molecule of Oxygen

(b) Carbondioxide (CO_2)

Similarly, in carbon dioxide, carbon atom shares four electrons with two oxygen atoms and form two double covalent bonds.



$O + C + O \longrightarrow CO_2$

Fig. 4.9 Molecule of Carbondioxide

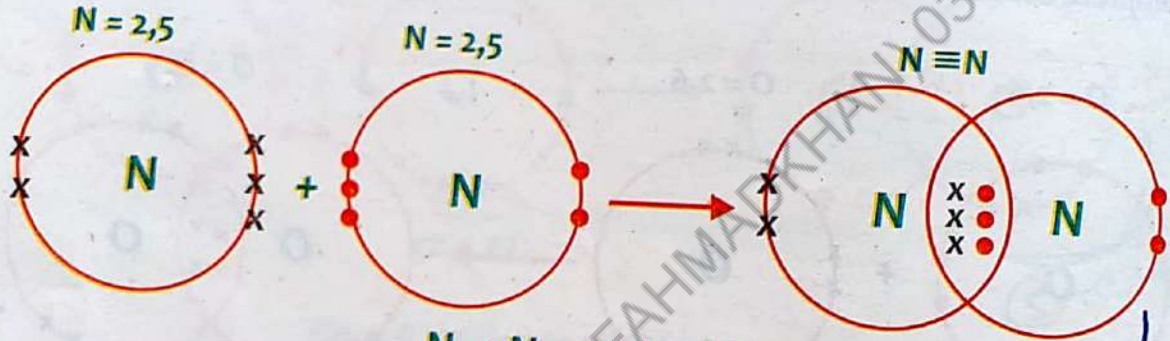
(5)

(iii) Triple Covalent Bond

The bond in which two atoms share three electrons each, to form three pairs of electrons is called triple covalent bond. A triple straight line (≡) shows a covalent bond in which total six electrons are shared. For example, N_2, C_2H_2 etc.

(a) Nitrogen (N_2)

Nitrogen molecule is formed from two Nitrogen atoms. The Nitrogen atom electronic configuration is (2, 5). A Nitrogen atom has five electrons in its valence shell and it shares three electrons with another Nitrogen atom to form nitrogen molecule. In this way, both Nitrogen atoms attain inert gas (Ne) electronic configuration and complete their octet.



$N + N \longrightarrow N_2$

Fig. 4.10 Molecule of Nitrogen

End (5)

Scientific Information

Sharing of electrons in the formation of covalent bonds rather than complete transfer from one atom to the other as in the formation of ionic bonds depends upon the differences in electronegativity of the bonding atoms. If this difference is more than 1.7, electron transfer will occur and ionic bond will form. If the electronegativity difference is less than 1.7, electrons will be shared and covalent bond will form.

Activity:4.1

Classify the following substances as covalent or ionic:

Co_2, NaCl, H_2O, C_6H_6, $MgCl_2$ and HCl.

EXAMPLE: 4.2

Which one of the following is (1) uncombined atoms (2) molecules (3) and consist of ions?

(a) O_2 (b) Ne (c) NO_2 (d) KCl (e) Na_2O

Solution

(a) O_2 consists of molecules.

(b) Ne consists of uncombined atoms.

(c) NO_2 consists of molecules.

(d) KCl is ionic.

(e) Na_2O consists of ions.

Practice Problem: 4.2

Which one of the following is (1) uncombined atoms (2) molecules and (3) consist of ions?

(a) H_2 (b) He (c) HCl (d) HgO

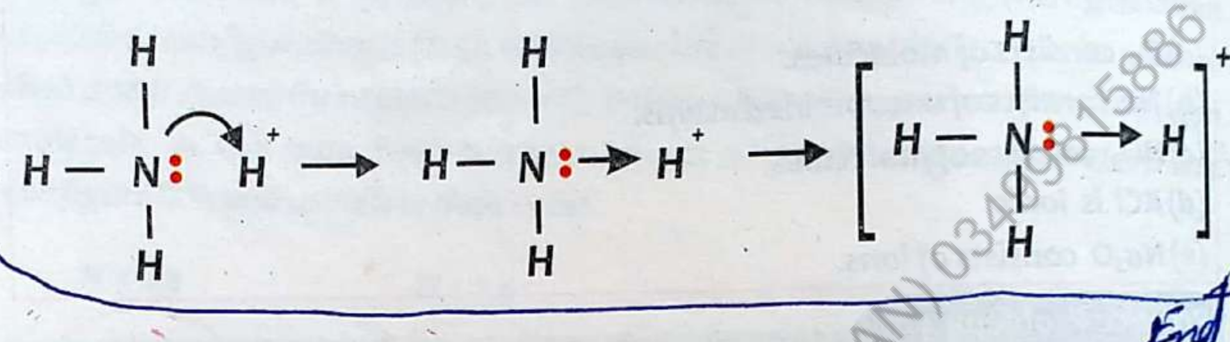
4.3.3 Co-ordinate Covalent Bond or Dative Bond

As defined above a covalent bond is formed between two atoms, when they share valence electrons with each other. In some compounds, it is observed that electrons shared for the formation of covalent bond between atoms do not come from both bonding atoms, the pair of electron is donated by one of the bonding atoms only. This type of covalent bond is called co-ordinate covalent bond or dative bond. The covalent bond in which only one atom donates the shared pair of electron is called co –ordinate covalent bond or Dative Bond.

The atom, which donates the shared pair of electrons, is called donor. The atom, which accepts the shared pair of electrons, is called acceptor and electron deficient species. An arrow (→) represents the co – ordinate covalent bond, which is directed from the donor to the acceptor atom. The co – ordinate covalent bond is different from the ordinary covalent bond. Those compounds, which have the co – ordinate bond, are known as the co – ordinate covalent compounds. For example,

(i) Ammonium ion (NH_4^+)

In ammonia molecule, the Nitrogen atom is bonded to three Hydrogen atoms. There is still one unshared pair of electrons with the Nitrogen, an electron rich species. The Hydrogen ion (H^+) is electron deficient species. Therefore, Nitrogen donates this lone pair of electron and the Hydrogen ion accepts this electron pair, forming the Ammonium ion (NH_4^+).



End (6)

Fig. 4.11 Structure of Ammonium

(ii) Addition compound of NH_3 and BCl_3

The Nitrogen atom of NH_3 has lone pair of electron and is electron rich species. The Boron atom of the BCl_3 is short of two electrons to complete its octet. An addition compound is formed when the Nitrogen of NH_3 donates this lone pair of electron to the Boron of BCl_3 which accepts it and forms the addition compound.

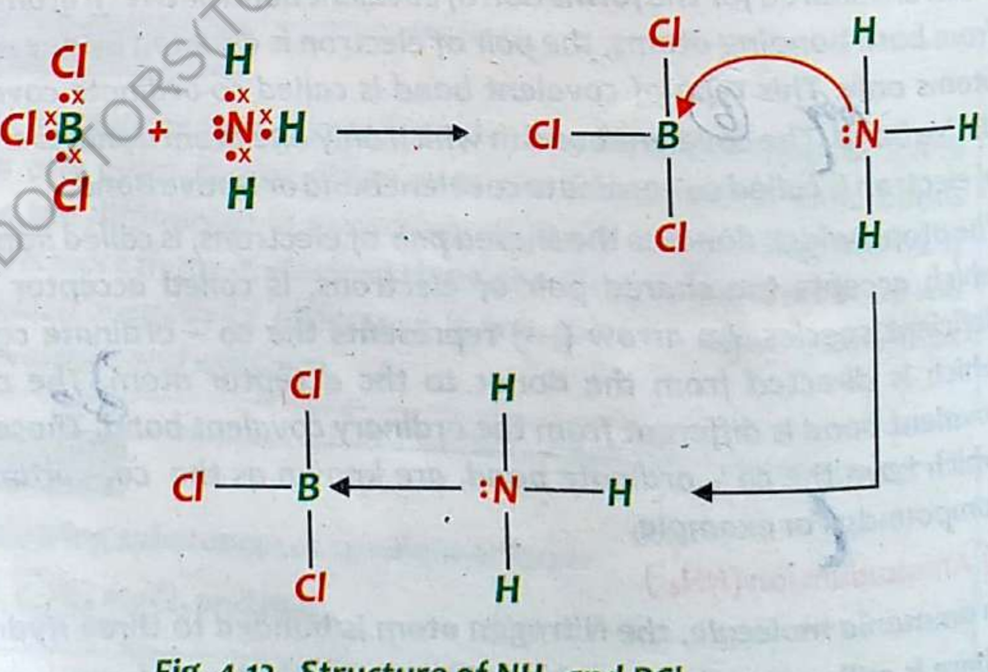


Fig. 4.12 Structure of NH_3 and BCl_3

Lewis structure

In Lewis structures, unshared (non-bonded) electrons are shown as dots or cross and shared pair (bonded) of electrons is shown as lines between atoms. An example of Lewis structure of water is,

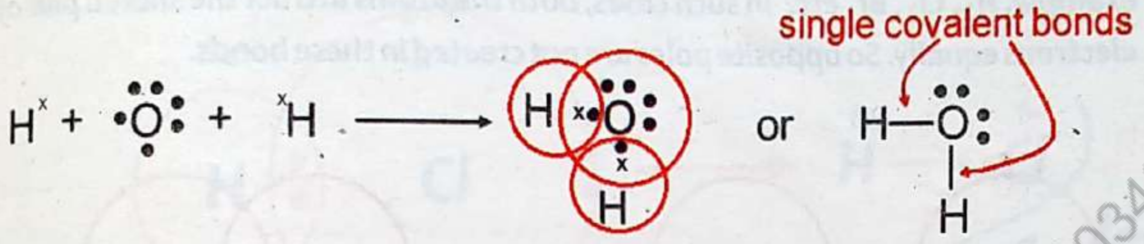


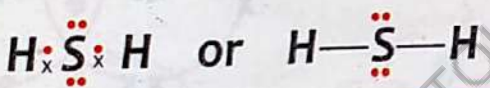
Fig. 4.13 Lewis structure of water

EXAMPLE: 4.3

Draw the Lewis structure for hydrogen sulphide, H_2S.

Solution

Answer:



Practice Problem: 4.3

Draw the Lewis structure of ammonia, NH_3.

Test Yourself.

1. a. which group of the elements form covalent bonds?
   b. How is a covalent bond formed between two atoms?
2. How many electrons are shared in a,
   a. Single bond b. double bond
3. Draw a 'dot and cross' diagram to show the bonding in each of the following molecules:
   a. Chlorine, Cl_2 b. Hydrogen chloride, HCl c. Ammonia, NH_3

Ans 7

4.3.4 Polar and Non-Polar Bonds

(i) Non – polar covalent bond

The covalent bond which is formed by the mutual sharing of electrons between atoms, having similar electronegativities is called non – polar covalent bond. For example, H_2, Cl_2, Br_2 etc. In such cases, both the atoms attract the shared pair of electrons equally. So opposite poles are not created in these bonds.

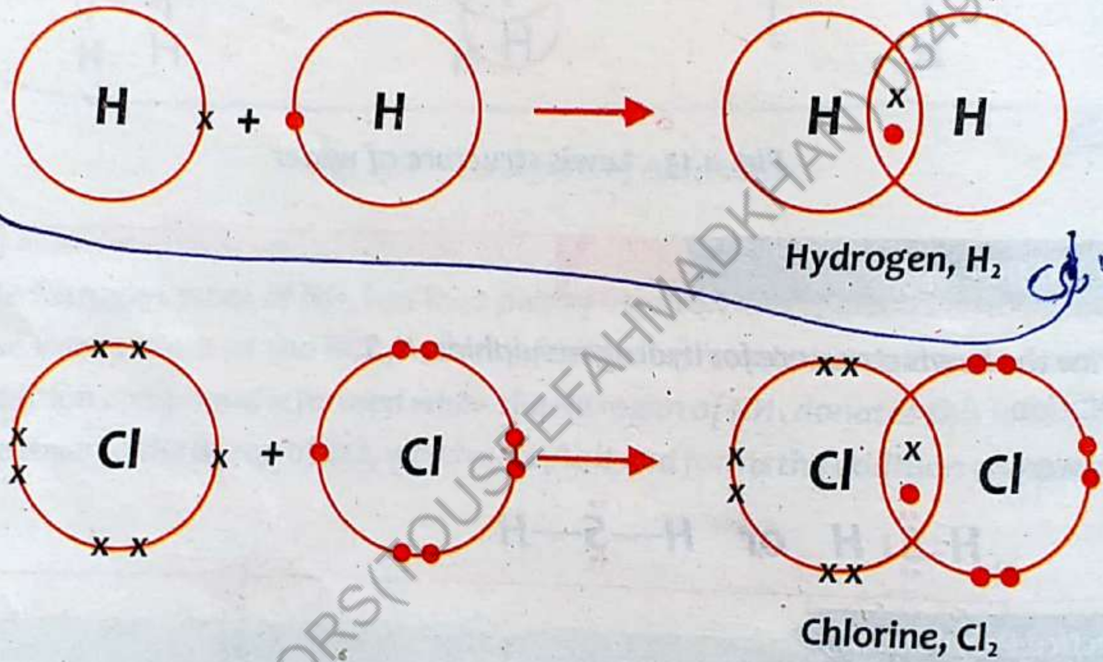


Fig. 4.14 Non- Polar bond in H_2 and Cl_2 molecules

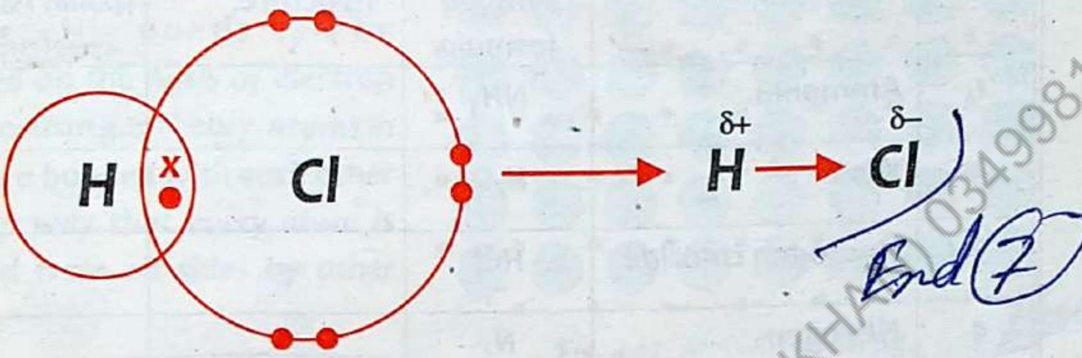
S.Q. 7

(ii) Polar covalent bond

The covalent bond which is formed by the mutual sharing of electrons between atom, having different electronegativities is called polar covalent bond.

Consider the 'HCl' molecule. There is single covalent bond between 'H' and 'Cl' atoms. As the electronegativity of Chlorine atom is greater than the electronegativity of Hydrogen atom. Therefore, the shared pair of electrons between 'H' and 'Cl' atoms is attracted more towards 'Cl' than Hydrogen atom.

Hence, a partial positive charge is produced on Hydrogen atom and a partial negative charge is produced on Chlorine atom. So opposite poles are formed in H– Cl molecules and the bond is called a polar bond.



End 7

Fig. 4.15 Polar bond in HCl molecule

Other examples are H_2O, NH_3, etc.

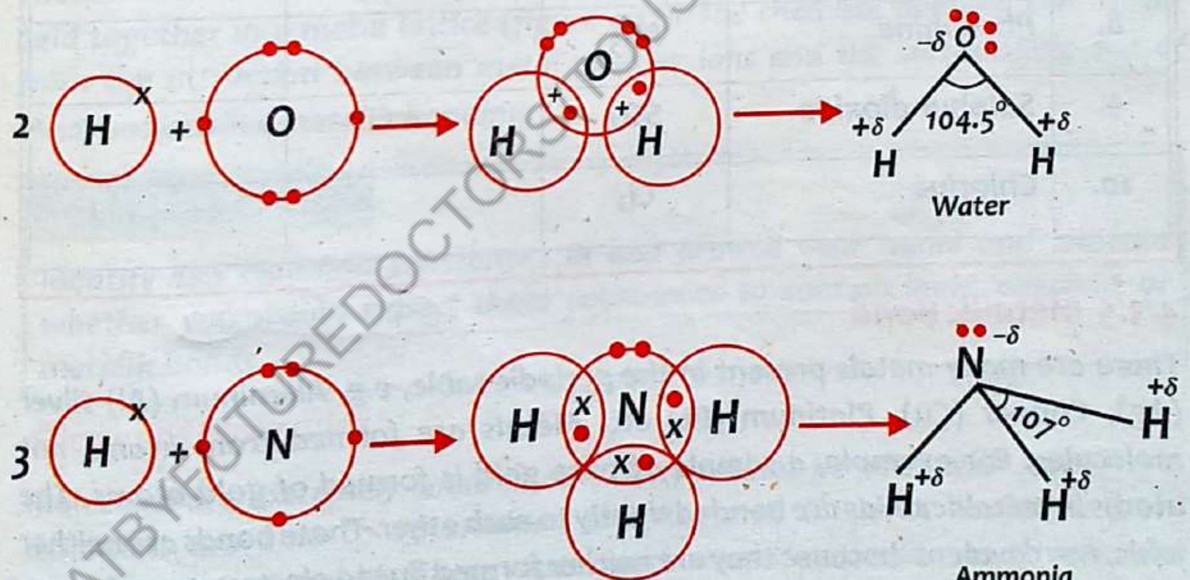


Fig. 4.16 Polar bond in H_2O and NH_3 molecules

Test Yourself.

S. No.	Compound	Chemical symbol/ formula	Lewis structure	Type of bond (polar/ non polar covalent bond)
1.	Ammonia	NH_3		
2.	Water	H_2O		
3.	Hydrogen chloride	HCl		
4.	Nitrogen	N_2		
5.	Oxygen	O_2		
6.	Methane	CH_4		
7.	Hydrogen	H_2		
8.	Phosphine	PH_3		
9.	Sulphur dioxide	SO_2		
10.	Chlorine	Cl_2		

4.3.5 Metallic Bond

There are many metals present in the periodic table, e.g. Aluminum (Al) silver (Ag), Copper (Cu), Platinum (Pt) etc. Metals are formed from atoms not molecules. For example, a sample of pure gold is formed of gold atoms. The atoms in metallic solids are bonded tightly to each other. These bonds are neither ionic, nor covalent because they are neither formed due to electron transfer, nor due to sharing of electrons between the atoms. In metals, we deal with different kind of bond called metallic bond.

(8) An (8)
Metallic bond

The metallic bonds cannot be explained on the basis of ionic, covalent bond or Van der waal's forces. There are various theories, which explain the metallic bonds. The metallic bonds can be explained on the basis of electron sea or electron gas theory. Atoms in metals are bonded with each other in such a way that every atom is attracted from all sides by other atom.

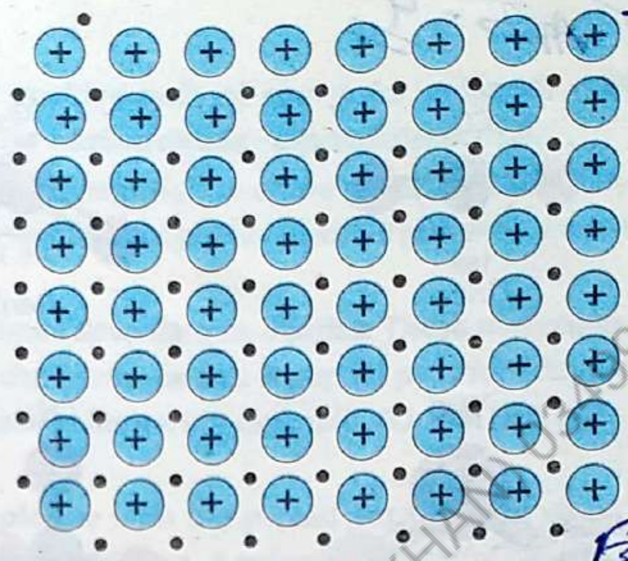


Fig. 4.17 Metallic bonding

End (8)

In the metal lattice, metal atoms lose their outer electrons and become positively charged ions. The outer electrons of the atom no longer belong to any metal atom and are said to be delocalized, which means that they do not belong to any one atom but move freely between the metal ions like a cloud of negative charge. These mobile electrons form a sea of electrons around the metal atoms, which are held together in a metal lattice (figure 4.17). The chemical bonding that result from the attraction between metal positive ions and the surrounding sea of electrons is called metallic bonding.

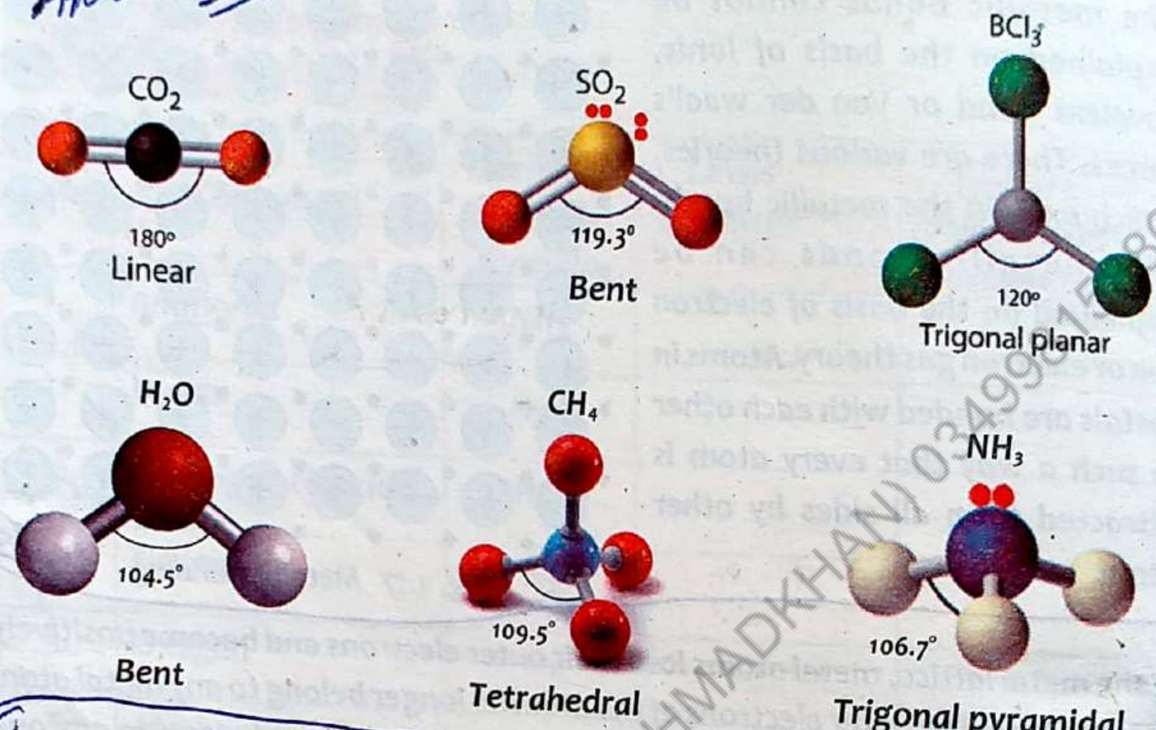
Activity:4.3

Identify ten common substances in and around your home and indicate whether you would expect these substances to contain ionic, covalent or metallic bonds.

Shapes of Molecules

Molecules are extremely small in size. They cannot be seen with naked eye. Scientists are able to determined the size and shape of molecule from experiments and analysis. These experiments show that the shapes of molecules are linear, triangular, tetrahedral, pyramidal or any other shape. For example, the molecular shape of CO_2 is linear, the water molecule is angular, BF_3 has trigonal planner, CH_4 is tetrahedral and NH_3 is trigonal pyramidal.

Ans - 9



End 9

Fig. 4.18 Shapes of CO_2, SO_2, BCl_3, H_2O, CH_4 and NH_3

Ans (10)

4.4 Intermolecular Forces

The forces of attraction between the molecules of a compound are called intermolecular forces.

Intermolecular forces are the weaker forces of attraction. It is 25 times weaker than the covalent bond. For example, in H_2O molecules, it requires 464 $kJ.mol^{-1}$ to break the H – O bonds within a water molecule and needs only 19 $kJ.\ mol^{-1}$ to break the intermolecular forces between water molecules. End (10)

The intermolecular forces are much weaker among the molecules of gases, whereas, they are stronger in the molecules of liquid and much stronger in solid substances. The melting point of solids and boiling point of liquids depends on the strength of these forces.

The intermolecular forces are of three types and collectively called Van Der Waals forces. We will discuss here two such types of forces, whereas the third type will be discussed in higher grades.

Ans: (11)

4.4.1 Dipole - Dipole Interaction

The attractive forces between the positive pole of one polar molecule and negative pole of other polar molecule are called dipole – dipole interaction.

Due to the difference between the electro-negativities of the atoms in molecules, the electrons is not shared equally; one atom has partial positive and other atom partial negative charges.

Polar covalent molecules are sometimes described as dipoles. These molecules have two poles. One end (pole) of the molecule has a partial positive charge while the other end has a partial negative charge. For example, CO, NO, H_2S, H_2O, NH_3 and HCl.

Hydrochloric acid (HCl) is a polar molecule with the partial positive charge on Hydrogen and partial negative charge on Chlorine. A network of partial positive δ(+) and partial negative δ(–) charges attract molecules to each other. Because of the force of attraction between oppositely charged ends (poles), there is a small dipole-dipole force of attraction between adjacent HCl molecules. The dipole-dipole interaction in HCl is relatively weak; only 3.3 kJ/mol energy is required to break that interaction. The force of attraction between HCl molecules is so small that Hydrogen Chloride (HCl) boils at –85.0°C.

End (11)

show diagram (11)

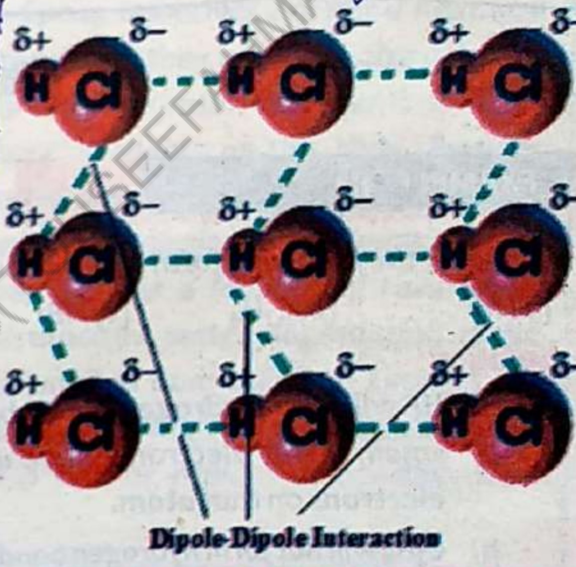


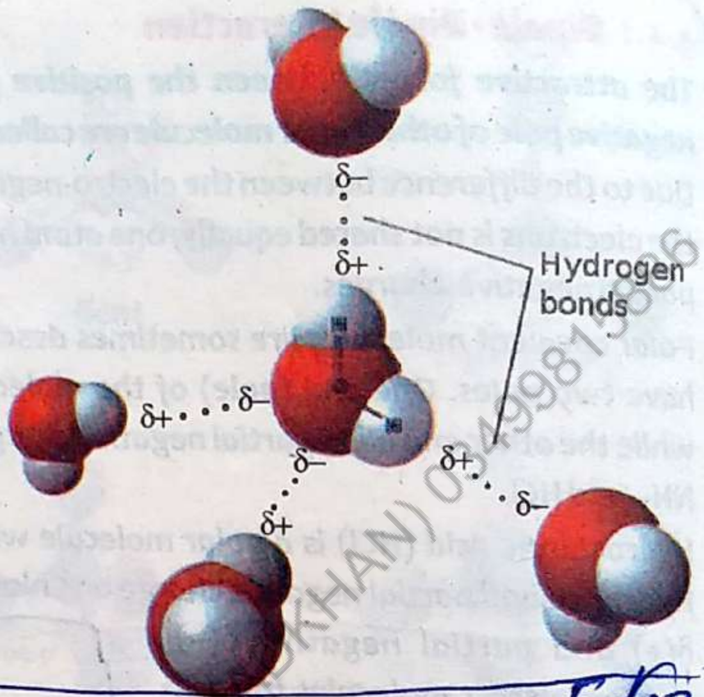
Fig. 4.19 Dipole-Dipole Interaction

Ans (12)

4.4.2 Hydrogen Bonding

Molecule in which hydrogen is covalently bonded to very electronegative atom such as Fluorine (F), Oxygen (O) or Nitrogen (N). An appreciably strong positive charge is thus developed on the hydrogen atom due the large difference in their electronegativity values with other atoms. As a result ,strong force of attraction is created when this positively charged hydrogen atom of one molecule is attracted by Fluorine (F), Oxygen (O) or Nitrogen (N) belonging to a neighbouring molecule. This force of attraction is called hydrogen bonding.

A hydrogen bond is the attractive force between the highly electron deficient hydrogen atom and nearby highly electronegative atom with lone pair of electrons such as F, O or N.



End 12

Fig. 4.20 Hydrogen Bonding in water

EXAMPLE: 4.4

Which of the following form hydrogen bonds?

(a) HF (b) C_2H_4 (c) HBr (d) NH_4^+ (e) H_2

Solution

a) HF will form hydrogen bonds because it has (1) hydrogen atom; (2) a small, highly electronegative atom; and (3) at least one free pair of electrons on that atom.

b) C_2H_4 will not form hydrogen bonds because, the carbon atoms have no lone pairs of electrons to bond with hydrogen atoms from other molecules.

c) HBr will not form hydrogen bonds, because the bromine atom is too large and not electronegative enough.

d) NH_4^+ by itself cannot form hydrogen bonds because it has no lone pairs.

e) The covalent bond in H_2 is a chemical bond, not the intermolecular force called hydrogen bonding.

Practice Problem: 4.4

Which of the following substances are expected to form hydrogen bonds?

(a) H_2O (b) CH_3OH (c) NH_3

Society, Technology and Science

Epoxy is polymer that is formed from two different chemicals. These are referred to as resin and the hardener. Epoxy adhesives are called structural adhesives. These high performance adhesives are used in the construction of aircraft, automobiles, bicycles, boats and golf clubs, where high strength bonds are required. Epoxy adhesives can be developed to suit almost any application. They can be made flexible or rigid, transparent or opaque even colored as well as fast or slow setting. Epoxy adhesives are good heat and chemical resistant. Because of these properties, they are given the name of engineering adhesives. The large family of epoxy resins represents some of the highest performance resins of those available at this time. Epoxies generally out-perform most other resin types in terms of mechanical properties and resistance to environmental degradation, which leads to their almost exclusive use in aircraft components. As a laminating resin their increased adhesive properties and resistance to water degradation make these resins ideal for use in applications such as boat building. Epoxies are widely used as a primary construction material for high-performance boats or as a secondary application to sheath a hull or replace water-degraded polyester resins and gel coats.

Ans (13)

(i) Application of Hydrogen Bonding

These intermolecular forces are extremely important in determining properties of water, biological molecules such as proteins, DNA etc. and synthetic materials such as glue, paints, resins etc. The adhesive action of paints and dyes is developed due to hydrogen bonding. Synthetic resins also bind two surfaces together by hydrogen bonding or dipole-dipole interactions. End (13)

Ans (14)

(ii) Properties of Hydrogen Bond

i. Hydrogen bond is stronger than Dipole-Dipole forces but weaker than covalent bond. It is about twenty times weaker than covalent bond and ten times stronger than the dipole-dipole interaction.

ii. Hydrogen bond is directional

iii. Hydrogen bond forms long chains and helps in the formation of network of molecules. End (14)

4.5 Nature of Bonding and Properties

The nature of bonding and properties of different compounds, are briefly given below.

4.5.1 Properties of Ionic Compounds

Ionic compounds are formed by the strong electrostatic force of attraction between the two oppositely charged ions. (15) Ans Some properties of ionic compounds are following.

i. Solid at room temperature.
ii. Having sharp melting and boiling points.
iii. Soluble in polar solvents like water.
iv. Good electrolytes in molten or solution form.
v. Non – directional bonds in ionic compound.
vi. Having reactions in molten state or in solution form.
vii. Compounds are composed of cations and anions in crystalline form. (15) End

Scientific Information

The greater the charges on the ions of an ionic compound, the higher will be its melting point. For this reason, the melting point of magnesium oxide is 2800°C, so it can be used as refractory material for lining of furnaces.

4.5.2 Properties of Covalent Compounds

Covalent compounds are also called molecular compounds, the properties of covalent compounds depends on,

i. Geometrical shape of molecules.
ii. Polarity and intermolecular forces among molecules.
iii. Bond type, whether single, double or triple.

Ans (16)

Some of the important properties of covalent compounds are following.

i. Covalent compounds have low melting points and boiling points.

ii. Covalent compounds are non electrolytes in their solution form.

iii. The bonds in covalent compounds are directional.

iv. The crystals of covalent compounds are composed of molecules.

v. Reactions of covalent compounds are slower than the ionic compounds.

vi. Polar covalent compounds are soluble in polar solvent like water(H_2O), Alcohol, etc. while non-polar covalent compounds are soluble in non-polar solvent like benzene (C_6H_6), Acetone (C_3H_6O), carbon tetrachloride (CCl_4) etc. End (16)

(17) Properties of Metals

Atoms of metals are held together by special type of bonds called metallic bond. Properties of metals result from this type of bonding are listed below.

i. All metals are solid at room temperature and pressure except Mercury (Hg)

ii. Metals are malleable; they can be beaten into sheets and foils.

iii. Metals are ductile; they can be drawn into wires as shown in fig. 4.21.

iv. Metals are good conductors of heat and electricity.

v. Metals are lustrous; they have shiny surfaces.

vi. Metals are sonorous; they produce specific ringing sounds when struck. End (17)

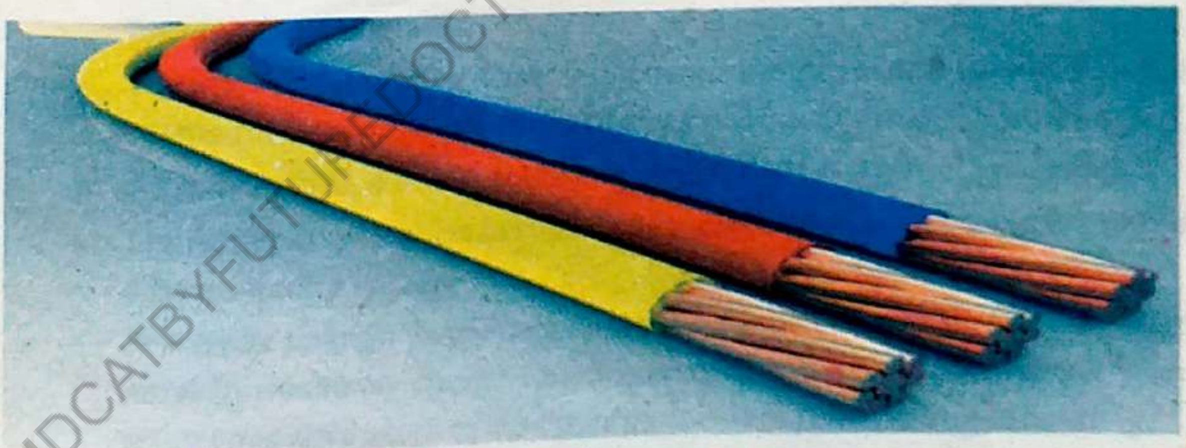


Fig. 4.21 Copper wire

Activity:4.4

The first column in the table below lists some general properties of metals. Complete the second and third columns of the table.

General property of metals	Correct name for this property	One use that depends on this property
can be drawn into wires		
can be bent into shape		
reflect light		
make a ringing sound when struck		
allow electricity to pass through		
heavy for their volume		
transfer heat well		

Key Points

- The attractive force which holds the atom together in the form of a compound is called a chemical bond.
- The electrons present in the valence shell of an atom are called the valence electrons.
- The chemical bond which is formed due to the complete transfer of electron(s) from one atom to the other atom is called ionic bond or electrovalent bond.
- The bond which is formed by the mutual sharing of electrons is called covalent bond.
- Covalent bond may be single, double or triple.
- The covalent bond in which only one atom donates the shared pair of electron is called Dative Bond.
- A polar covalent bond is formed by the mutual sharing of electrons between different atoms.
- A non – polar covalent bond is formed by the mutual sharing of electrons between similar atoms.
- The chemical bonding that result from the attraction between metal atoms and the surrounding sea of electrons is called metallic bonding.
- Structure of CO_2 is linear, H_2O is angular, and NH_3 is trigonal pyramidal.
- The forces of attraction between the molecules of a compound are called intermolecular Forces.
- The weak linkage formed between two molecules in such a way that partially positively charged Hydrogen of one molecule is attracted to an electronegative atom of the other molecule to form hydrogen bonding.

Exercise

Choose the correct option.

i. An atom with a charge is called:

a. an electron b. a molecule

c. a metal d. an ion

ii. An element X is in group VI of the periodic table. The ion will be represented by:

a. X^+ b. X^-

c. X^{-2} d. X^{+2}

iii. Which pair of elements will join to form a compound with one to one ratio:

a. Magnesium and Chlorine b. Sodium and Oxygen

c. Potassium and Fluorine d. Lithium and Sulphur

iv. When calcium atom become a Calcium ion(Ca^{+2}):

a. it loses an electron b. it loses two electron

c. it gains electron d. gains two electrons

v. Which two elements will form a covalent compound:

a. Sodium and Oxygen b. Copper and Oxygen

c. Carbon and Oxygen d. Magnesium and Oxygen

vi. In the formation of an ionic bond, the atoms taking part:

a. only gain electrons b. share electrons

c. lose and gain of electrons d. only lose electrons

vii. Fluorine has an electronic configuration 2, 7 and Oxygen 2, 6. The formula of fluorine oxide will be:

a. FO b. F_2O

c. FO_2 d. F_2O_2

viii. Which of these statements about covalent bonds is incorrect:

a. HCl contains one pair of shared electron

b. CCl_4 contains four pairs of shared electrons

c. H_2O contains three pairs of shared electrons

d. NH_3 contains three pairs of shared electrons

ix. Which of the following ions do not have the electronic configuration of an argon atom?

a. Ca^{+2} b. S^{-2}

c. K^{+} d. O^{-2}

x. Which one of the groups below contains only ionic compounds?

a. Calcium Oxide, Hydrogen Chloride, Magnesium Oxide
b. Calcium Oxide, Magnesium Oxide, Sodium Chloride
c. Carbon Dioxide, Copper Sulphate, Hydrogen Chloride
d. Copper Sulpahte, Methane, Sodium Chloride

SHORT QUESTIONS

Answer briefly the following questions.

i. What is electron-sea model of metallic bonding?
ii. Why most atoms are chemically bonded to other atoms in nature?
iii. Identify and define the four major types of chemical bonding.
iv. Arrange the following from strongest to weakest attraction:
a. Covalent bond
b. Dipole-dipole interaction
c. Hydrogen bond
d. Ionic bond
v. Why ionic compounds are good electrolyte is molten and solution form not in solid state.
vi. What types of elements/ atoms tend to form the following types of bonding?
a. ionic b. covalent c. metallic
vii. Give an example of a non-polar molecule with polar bonds. Give reasons.
viii. Predict the bond type (ionic, polar covalent, non-polar covalent) in each of the following:
a. $CaCl_2$ b. H_2O
c. CO_2 d. C_2H_4
ix. Why ionic compounds are good conductor compared to covalent compounds?
x. Give the reason that why dative bond is always polar?

5 5

LONG QUESTIONS

(i) Describe the octet rule in terms of noble-gas configurations and stability.

(ii) (a). What is the main distinction between ionic and covalent bonding? Explain you answer with suitable examples.

(b) How is electronegativity used in determining the ionic or covalent character of the bonding between two elements?

(iii) Draw the Lewis structure for each of the following compounds:

a. CO b. HCl

c. SO_2 d. CCl_4

e. BF_3 f. NH_3

(iv) Explain why most metals are malleable and ductile but ionic crystals are not.

(v) a. What is the meaning of the term polar, as applied to chemical bonding?

b. Distinguish between polar-covalent and non polar-covalent bonds.

Project Work

Prepare a report on the work of Linus Pauling. Discuss his work on the nature of the chemical bond.

Unit 5 Physical States of Matter

After studying this unit, the students will be able to:

- *Explain the effect on the pressure of a gas by a change in the*
  - *a. volume*
  - *b. temperature*
- *Compare the physical states of matter with regard to intermolecular forces present between them.*
- *Account for pressure-volume changes in a gas using Boyle's Law.*
- *Account for temperature-volume changes in a gas using Charles's Law.*
- *Explain the properties of gases (diffusion, effusion and pressure).*
- *Summarize the properties of liquids like evaporation, vapour pressure and boiling point.*
- *Explain the effect of temperature and external pressure on vapour pressure and boiling point.*
- *Describe physical properties of solids (melting and boiling points).*
- *Differentiate between amorphous and crystalline solids.*
- *Explain the allotropic forms of solids.*

Introduction

Anything that has mass and occupies space is called matter. All the substances that you see around are made of matter. Matter can exists in three physical states i.e. gas, liquid and solid. These include the air that we breathe in, the water we drink and the food we eat. The scientific study of matter and their physical states of matter provides an opportunity to understand our world.

All the three states of matter are inter-convertible to one another. When the temperature of solid is increased it is converted into liquid and when the temperature of the liquid is increased it is then converted into gaseous state. So, it is considered mandatory that one should know the matter and its physical states and properties.

In previous grades, you have learnt about the air (gas) and water (liquid). In this unit , you will learn about properties of gases, liquids and solids, gas laws and types of solids.

Scientific Information

There are more states of matter also. We are not as familiar with them, nor do we see them every day. Some exist only in theory, others can be reproduced in laboratories, some are so new that scientists are still figuring out the details and others might exist, but have not yet been found in nature. These include: plasmas, liquid crystals, superfluids, supersolids and the paramagnetic etc.

Gaseous state

The state of matter in which the molecules are far away from each other and there is very weak intermolecular forces of attraction between them. The particles are free to move in all direction with great speed. Gases are restricted only by the walls of the container. Gases have no definite shape and volume and occupy the whole volume of its container.

5.1 Typical Properties of Gases

Some typical properties of gaseous state are following.

i. Indefinite Volume

Gases have no definite volume and they occupy all the available space.

ii. Indefinite Shape

Gases do not have shape but take the shape of the container in which it is placed.

iii. Diffusion

Diffusion describes the movement of molecules from a region of higher concentration to a region of lower concentration. In other words, the molecules tend to move from a more crowded region to a less crowded region. Once the molecules become evenly distributed throughout the medium, there will be no net diffusion in any particular direction.

Ans (2) Gas molecules are in constant random motion. Because of this motion, the gas molecules diffuse very quickly and form a homogenous mixture. **The spontaneous mixing of the molecules of different gases to form a homogenous mixture by random motion and collision of the molecules is called diffusion.** For example, when a gas having characteristic odour is released at one corner of the room, soon its smell will spread in the whole room. Similarly, a volatile liquid (perfume) evaporates and spreads in the air in a very short time. The rate of diffusion of gases varies from gas to gas. In 1833, Thomas Graham an English chemist, discovered that a lighter gas can diffuse much faster than a heavier.

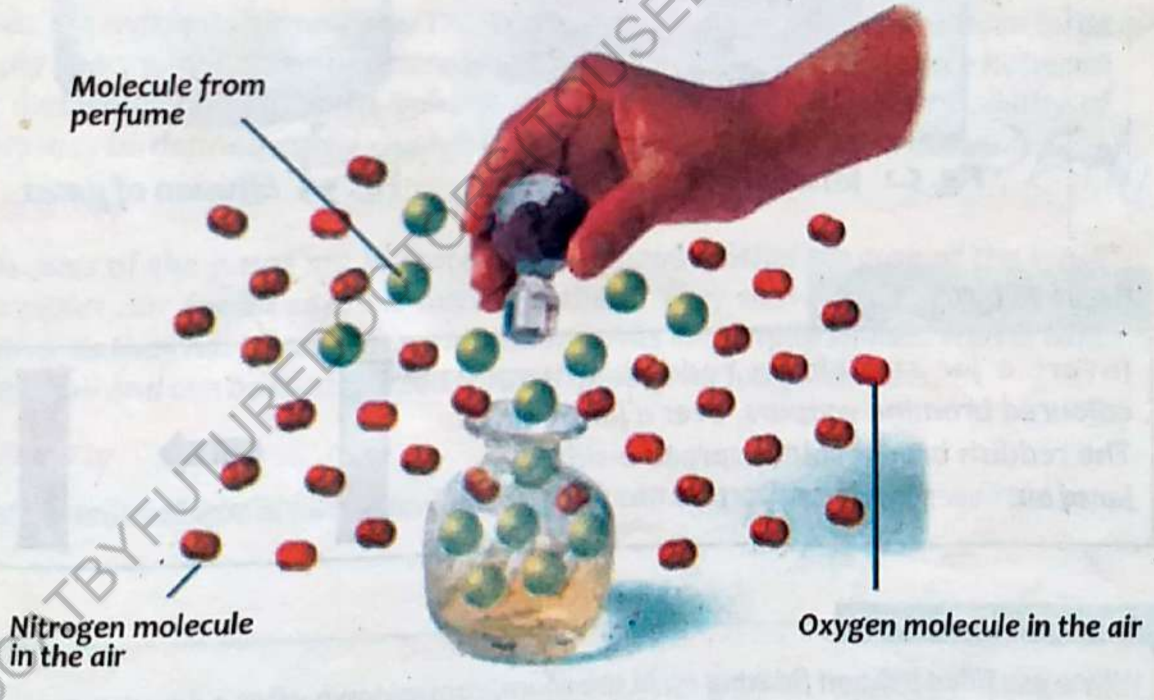


Fig. 5.1 Diffusion of Perfume Molecules In Air

2 Ans

iv. Effusion

When a gas is allowed to pass through a hole of the size of a molecule, only one molecule at a time can pass through the hole. **The escape of molecules in the gaseous state one by one without collision through a hole of molecular dimension is called effusion.** You can smell the onions even when the bag is tightly sealed, due to effusion.

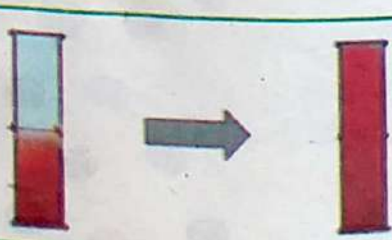
End 2

Fig. 5.2 Effusion of onion smell

Fig. 5.3 Effusion of gases

Activity:5.1

Invert a jar containing reddish brown coloured bromine vaopurs, over a jar of air. The reddish brown colour spread inside the jar of air.



Point to Ponder

Why a gas filled balloon floating up in the room, comes down after a day or two.

v. Pressure

The molecules of the gas are in continuous state of motion. They collide with each other as well as with walls of the container. When the gas molecules collide with walls of the container, they exert pressure on the walls of the container. Force (F) exerted by a gas on per unit area (A) of a container is called pressure (P).

$$P = F/A$$

The SI unit of force is Newton and that of area is m^2. Hence, the SI unit of pressure is $N.m^{-2}$. It is also called Pascal (Pa). Pascal is very small unit in pressure measurement, usually its bigger form is kilo Pascal (kPa) is used.

At sea level at 0°C, the atmospheric pressure is 760 mm of Hg or 760 torr. It is termed as one atmospheric pressure. So,

$$1 \text{ atmosphere} = 760 \text{ mm of Hg}$$
$$= 760 \text{ torr} \qquad (1 \text{ mm of Hg} = 1 \text{ torr})$$
$$= 101.325 \text{ KPa}$$
$$= 1.01325 \times 10^5 \text{ Pa}$$

vi. Compressibility

Gases are highly compressible. This is because that gases molecules have large empty spaces. When the pressure is applied on the gases, the distance between the molecules decreases, its volume also decreases. So the compressibility of gases may be defined as the change in volume per unit change in pressure. End (1)

vii. Mobility

Molecules of the gases are in state of continuous motion because of the weak intermolecular forces present between them. They move from one place to another as they have weak intermolecular forces and empty spaces. That is why, gases flow and can be transported through pipes over long distances.

viii. Density

Mass per unit volume is called density. Density can be mathematically written as,

$$\text{Density} = \frac{m}{v}$$

It is clear from the above formula that density is inversely proportional to its volume. The gases occupy all the space available to it because there are weak intermolecular forces and empty spaces between the molecules of the gases, which increase its volume, so density of a gas is very low as compared to the same amount of a liquid or a solid.

The gases have low density than the liquid and solid state. For example, the gaseous Oxygen has the density of $0.00142 g/cm^3$ at $0^\circ C$ and 1 atmospheric pressure, while the liquid Oxygen has the density of $1.149 g/cm^3$ at $-103^\circ C$ and Solid Oxygen has the density of $1.42 g/cm^3$ at $-252^\circ C$.

5.2 Laws Related To Gases

The behavior of gases towards pressure, volume, temperature and molar amount is governed by the following laws.

The typical laws related to gases are,

5.2.1 Boyle's Law

In 1662, Robert Boyle described the relationship between the volume and pressure of a gas at constant temperature. He observed that **the volume of a given mass of a gas is inversely proportional to the pressure applied, provided the temperature remains constant.** According to this law, volume (V) of a given mass of a gas decrease with the increase of pressure (P) and vice versa.

Mathematically, it can be written as,

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

$$V = K_b \times \frac{1}{P}$$

$$PV = K_b$$

Where, K_b is called constant for Boyle's law.

When the volume of a given mass of gas is changed from V_1 to V_2 and the pressure is changed from P_1 to P_2, then Boyle's law equation can be written as,

$$P_1V_1 = P_2V_2 \quad = K_b \text{ (at constant temperature)}$$

Where, P_1 = initial pressure P_2 = final pressure

V_1 = initial volume V_2 = final volume

According to the above equation, the Boyle's law can also be defined as, **the product of pressure and volume of given mass of gas remains constant provided the temperature is constant.**) End (3)

Experimental Verification of Boyle's Law

The apparatus used for the experimental verification of Boyle's law as shown in the figure 5.4. A certain mass of gas is enclosed in the cylinder. The volume of the gas is changed by increasing and decreasing the pressure. The volume at various pressures is noted. In each case, the product of pressure and volume remains constant at constant temperate and is found according to the Boyle's law.

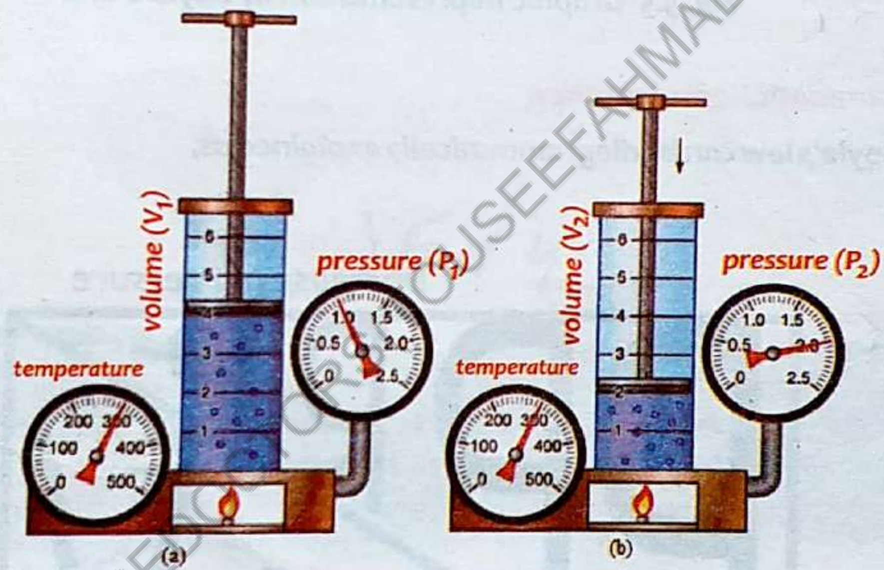


Fig. 5.4 Experimental Verification of Boyle's Law

Society, Technology and Science

Scientists use the power of reasoning to explain their observation. For instance, when a balloon is filled with air, it expands. A scientist would explain it by saying that air molecules are free to move inside their container. There is no attractive or repulsive force between the molecules. As a result, gas expands until it takes the shape of its container. Therefore, air expands to fill the interior of the balloon evenly.

Graphic Representation

If we plot the values of pressure 'P' and volume 'V', curve line is obtained which shows that the volume is inversely proportional to the pressure.

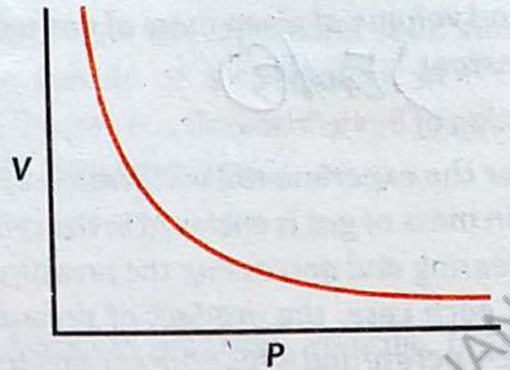


Fig. 5.5 Graphic Representation of Boyle's Law

Diagrammatic Representation

The Boyle's law can be diagrammatically explained as,

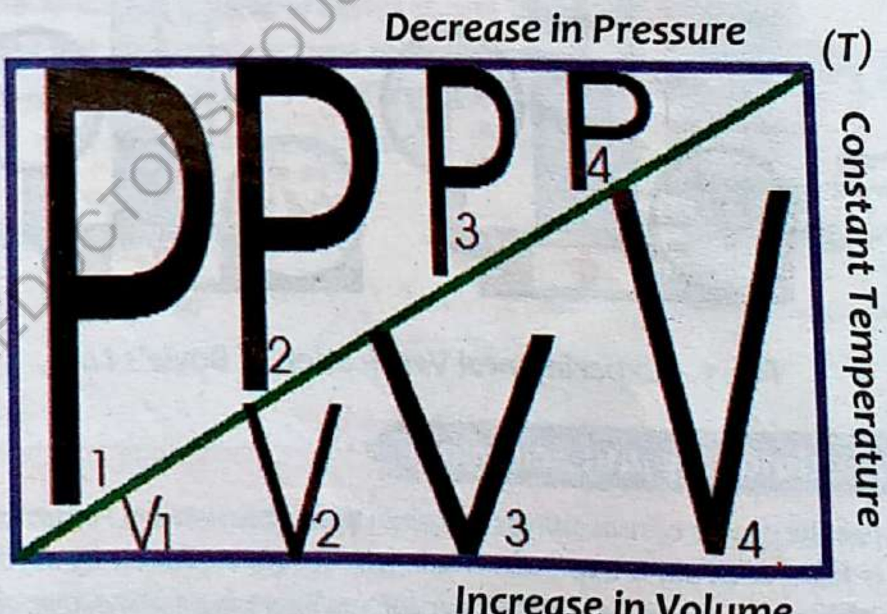


Fig. 5.6 Diagrammatic Representation of Boyle's Law

Ans (4)

EXAMPLE: 5.1

A $530 dm^3$ sample of hydrogen gas was collected in a container at 800 mm of Hg pressure, at room temperature. What volume will the gas occupy at 400 mm of Hg?

Solution:

$V_1 = 530 dm^3$ $\quad P_1 = 800$ mm of Hg

$V_2 = ?$ $\quad P_2 = 400$ mm of Hg

Applying Boyle's law, $P_1V_1 = P_2V_2$

Or

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

$$V_2 = \frac{800mm\ of\ Hg \times 530 dm^3}{400 mm\ of\ Hg}$$

$$V_2 = 1060 dm^3$$

End 4

Practice Problem: 5.1

Calculate the initial volume of a sample of gas at 1.20 atm. If its volume is changed to $70.4\ cm^3$ as its pressure is changed to 3 atm at constant temperature.

Ans (5)

5.2.2 Charles's Law

In 1787, Jacques Charles's gave a relationship between volume and temperature of a gas at constant pressure and formulated a law known as Charles's law. This law states that, the volume of a given mass of gas is directly proportional to the absolute temperature at constant pressure.

Mathematically the law can be expressed as,

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

$$V = K_c T$$

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

Where 'K_c' is called constant of Charles's law.

When the volume is changed from 'V_1' to 'V_2' by changing the temperature from 'T_1' to 'T_2', then the relationship can be written in the following form,

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

Where, T_1 = initial temperature

T_2 = final temperature

V_1 = initial volume

V_2 = final volume

From the above equation, the Charles's law can be defined as, **the ratio between volume and the absolute temperature of the given mass of a gas is constant at constant pressure.**

Experimental Verification of Charles's law

The apparatus used for the experimental verification of the Charles's law consists of a cylinder. The cylinder has a piston. The walls of the cylinder are heat insulator while the base of the cylinder is heat conductor. When the cylinder is heated at constant pressure, the piston move upward and the volume will increase. It is noted from various observations that the ratio between volume and absolute temperature remains constant. This verifies Charles's law.

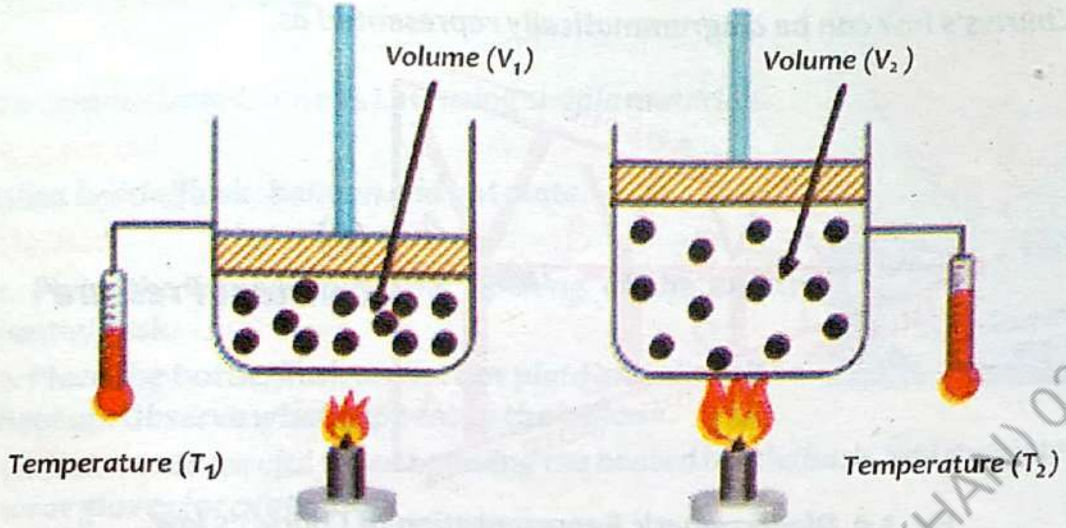


Fig 5.7 Experimental Verification of Charles's Law

Scientific Information

Absolute Temperature mean the temperature on the Kelvin scale. Each temperature on the absolute scale is 273 units greater than the same temperature on the Celsius scale. $K = {}^{\circ}C + 273$

Graphic Representation

If the values of volume 'V' is plotted against temperature 'T', a straight line is obtained, which shows that the volume is directly proportional to the temperature.

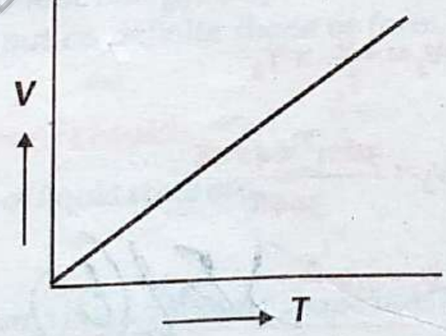


Fig 5.8 Graphic Representation of Charles's Law

The Charles's law can be diagrammatically represented as,

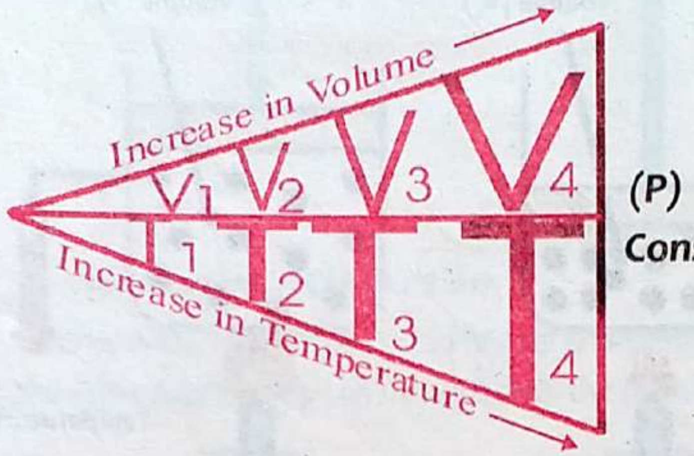


Fig 5.9 Diagrammatic Representation of Charles's law

EXAMPLE: 5.2

Ans (b)

If $3dm^3$ of air is heated from 300K to 400K at constant pressure then what is the volume of the gas at higher temperature?

Solution:

$V_1 = 3dm^3$ $\quad T_1 = 300K$

$V_2 = ?$ $\quad T_2 = 400K$

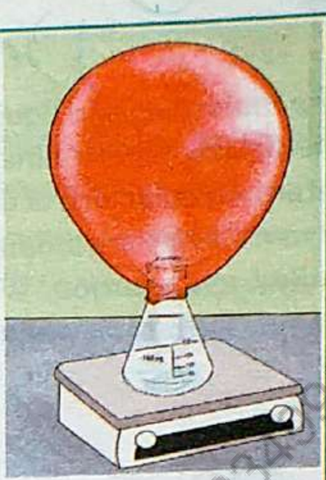
Applying Charles's law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

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

$$V_2 = \frac{3dm^3 \times 400K}{300K}$$

$V_2 = 4dm^3$ End (b)

Activity: 5.2



Aim:

To demonstrate Charles's Law using simple materials.

Apparatus:

Glass bottle/flask, balloon and hot plate.

Method:

1. Place the balloon over the opening of the empty bottle/flask.
2. Place the bottle/flask on the hot plate and allow it to heat up. Observe what happens to the balloon.

WARNING: Be careful when handling the heated bottle/flask. You may need to wear gloves for protection.

Results:

You should see that the balloon starts to expand. As the air inside the bottle/flask is heated, the pressure also increases, causing the volume to increase. Since the volume of the glass bottle/flask can't increase, the air moves into the balloon, causing it to expand.

Conclusion:

The temperature and volume of the gas are directly related to each other. As one increase, so does the other.

Ans (7)

Liquid State

Liquid state is the state of matter in which the intermolecular forces of attraction are stronger than gaseous state. These forces of attraction is strong enough to hold the molecules together but not so strong to stop the molecular motions. Liquid molecules have kinetic energy, they move randomly. A liquid therefore, possess a fixed volume but no definite shape or form. The surface of a liquid is always leveled.

5.3 Typical Properties of Liquids

Some typical properties of liquid state are,

(i) Volume and shape

Liquid has a definite volume but no definite shape that is why it takes the shape of the container in which it is placed.

Ans 8 7

(ii) Evaporation Ans 8

The molecules of a liquid move with different kinetic energies. The molecules of the liquids whose kinetic energies are higher, move faster and overcome the intermolecular attractive forces. These molecules come out of the liquid and convert into the gaseous state. Ans 8 The conversion of liquid state into gaseous state is called evaporation. We can say that the phenomenon in which a liquid is converted into its vapours without external heating is called evaporation.

7

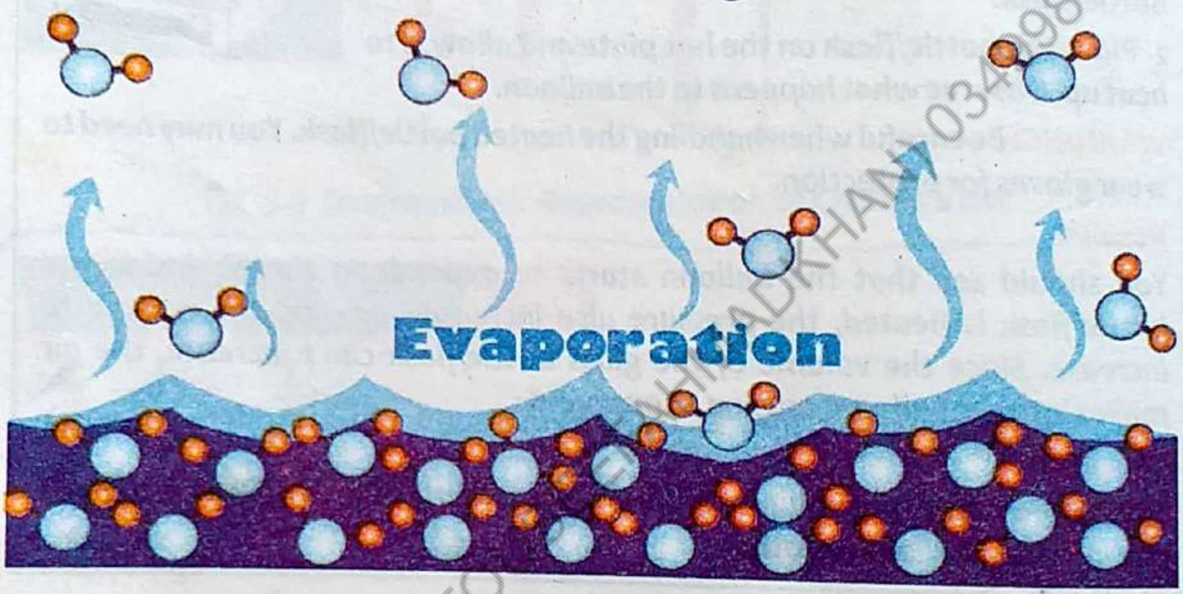


Fig 5.10 Process of Evaporation

8

Evaporation takes places at all temperatures. When the molecules of higher kinetic energies are converted into the gaseous state and escape into the atmosphere; the temperature of the remaining liquid falls. That is why evaporation cause cooling.

The Evaporation depends on the following factors 10

(a) Surface area

The process of evaporation takes place from the surface. Greater the surface area higher will be the rate of evaporation and vice versa.

(b) Temperature

With the increase in temperature, the average kinetic energy of the liquid molecules increases, so the rate of evaporation increases.

(c) Intermolecular force

The evaporation of the liquids depends upon the intermolecular forces. Stronger the intermolecular forces slow will be the rate of evaporation and vice versa. Different liquids have different rate of evaporation at the same temperature, because the attractive forces are different in different liquids. For example, water has stronger intermolecular force than alcohol, therefore alcohol evaporate quickly than water. End (8)

Ans 9 (iii) Vapour Pressure

(7) The pressure exerted by the vapours of the liquid when the rate of evaporation becomes equal to rate of condensation is called vapour pressure. (3)

When a liquid is placed in a closed container, the kinetic energies of all the molecules is not the same. The molecule of the liquid, whose kinetic energies is greater than the average, come to the liquid surface and are converted into the vapours, i.e. evaporation occurs. The vapours gather in the space above the surface of the liquid in closed container and produce vapour pressure. Some of the vapour molecules strike back to the surface of the liquid, lose their energy and becomes liquid again. In other words, condensation takes place.

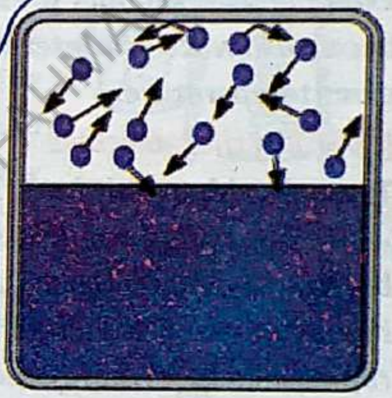


Fig 5.11 Vapours Pressure

As the rate of evaporation increases, the rate of condensation also increases. A state is reached, when the rate of evaporation becomes equal to the rate of condensation. This state is called an equilibrium state.

$$\text{Liquid} \underset{\text{Condensation}}{\overset{\text{Evaporation}}{\rightleftharpoons}} \text{Vapours}$$

Therefore, the vapours pressure can be defined as, the pressure exerted by the vapours of a liquid in equilibrium state.

Factors affecting vapour pressure

The vapour pressure of liquid does not depends upon the amount of liquid or surface area of liquid but depends on the following factors.

(a) Nature of liquid

The vapour pressure of liquid depends on the nature of liquid. Polar liquids having high boiling points and exert little vapour pressure at a given temperature, while non-polar liquids having low boiling liquids and exert more vapour pressure at the same temperature. For example, water has less vapour pressure than that of acetone at same temperature.

(b) Intermolecular Forces

Stronger the intermolecular attractive forces in a liquid lesser will be its tendency to evaporate, so lesser will be its vapours pressure at a given temperature. For example, the vapours pressure of ethyl alcohol is higher than that of Water, because in water the intermolecular forces are stronger than ethyl alcohol at the given temperature.

(c) Size of the Molecules

Those liquids which have small size are easily evaporated than those liquids which have large size. The small sized molecules liquids form more vapour than big sized molecule liquids at the same temperature. For example, pentane (C_5H_{12}) has small sized molecule than decane ($C_{10}H_{22}$), therefore pentane vaporizes rapidly and exert more vapour pressure than decane at the same temperature.

Table 5.1 Some Compounds and their Vapour Pressure at 298K

Name of compound	Formula	Vapour pressure at 298K (25°C) (mm Hg)
Chloroform	$CHCl_3$	170
Carbon tetrachloride	CCl_4	87
Water	H_2O	18
Glycerol	$C_3H_8O_3$	0.00016

(d) Temperature

Increase in temperature of a liquid, increases the average kinetic energy of the liquid molecules. The intermolecular distance among the molecules of the liquid increases, which cause decrease in the intermolecular attractive forces between the liquid molecules, Hence, the rate of evaporation of the liquid increases and thus vapours pressure increases. End (9)

iv. Boiling Point (7)

When a liquid is heated, the kinetic energies of the liquid molecules increases, which increase the vapour pressure of the liquid. A time is reached when the vapour pressure of liquid becomes equals to the atmospheric pressure (external pressure). At this moment, the liquid begins to boil and small bubbles of liquid go to the surface of the liquid, where they burst and vapours are released into the atmosphere. Ans (10) The temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure is called boiling point. End (7)

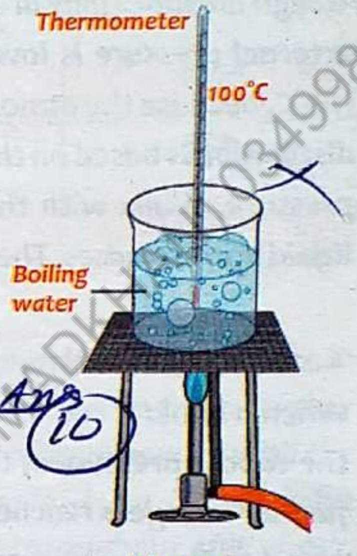


Fig 5.12
Boiling Point of Liquid

During boiling, the temperature of the liquid remains constant. Because the heat provided to the liquid during boiling is used to overcome the intermolecular attractive forces of liquid molecules. The boiling point of liquid depends on the following factors.

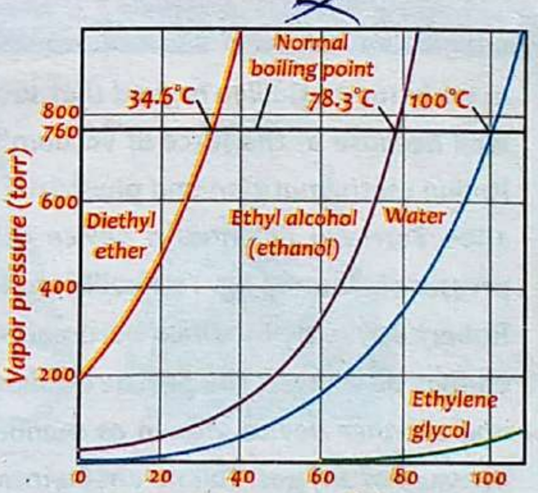


Fig 5.13 Comparison of boiling point of different liquid

(a) Nature of liquid (Intermolecular force)

The polar liquids have higher boiling point than that of non-polar liquids. In polar liquids, there are stronger intermolecular forces of attraction than non-polar liquids.

(b) External pressure

Boling point of liquid depends on the external pressure. With the increase in external pressure, the boiling point also increases. Similarly, the decrease in external pressure causes decrease in boiling point.

At high altitude (hilly areas), the boiling point of a liquid will be less because the external pressure is lower. Therefore, water boils at lower temperature than 100°C, because the atmospheric pressure is less than 760mm of Hg. The vacuum distillation is based on the decrease in boiling point with the decrease in external pressure. While with the increase in external pressure, the boiling point of a liquid also increases. The pressure cooker works on this principle. End (10)

v. Freezing Point

When a liquid is cooled the kinetic energies of the liquid molecules decrease, so the vapour pressure of the liquid also decreases. By decreasing the temperature further a stage is reached when the vapour pressure of the liquid state becomes equal to the vapour pressure of the solid state. The temperature at which the liquid and solid state exists in equilibrium state is called freezing point of that liquid.

Society, Technology and Science

In early 1600s, Galileo argued that suction pumps were able to draw water from a well because of the force of vacuum" inside the pump. After Galileo's death, the Italian mathematician and physicist E. Torricelli proposed another explanation. In 1646 Torricelli invented a device called barometer. His measured atmospheric pressure is 760mm Hg. Torricelli's work soon caught the attention of British scientist Robert Boyle. He modified barometric tube into a J-shaped tube. Boyle's from the studies discovered the pressure-volume relationship. J-tube was further modified and another device known as manometer was developed that can measure the pressure of any gas. This means instrumentation improves as science progresses.

Table 5.2 Freezing Point and Boiling Points of some Liquids

S. No.	Solvent	Formula	Freezing Point (°C)	Boiling point (°C)
1.	Water	H_2O	0.0	100.0
2.	Acetic Acid	CH_3COOH	17.0	118.1
3.	Benzene	C_6H_6	5.5	80.2
4.	Phenol	C_6H_5OH	43.0	181.0
5.	Chloroform	$CHCl_3$	−63.5	61.2
6.	Ethanol	C_2H_5OH	−114.7	78.4

vi. Diffusion

Diffusion can also happen in liquids. This is because the particles in liquids are in continuous state of motion. The molecules of liquid move from higher concentration to lower concentration. So, the molecules of liquid mix up with the other liquids and form a homogenous mixture. For example, if you drop little bit of ink into glass of water, the ink will spread slowly throughout the water.

a. Ink is dropped in. b. Diffusion begins c. Ink is evenly distributed.

Fig 5.14 Diffusion of liquid

Diffusion in liquids is slower than diffusion in gases because the molecules in liquid move more slowly than gas molecules. The rate of diffusion increases with increasing temperature. The increase in temperature increases the kinetic energy of liquid molecules so the rate of diffusion increases.

vii. Mobility

The kinetic theory of matter (particle theory) says that liquids consist of molecules which are constantly moving or in a constant state of motion. This mobility of molecules depends on the energy of molecules. Greater the kinetic energy, the higher will be the mobility of molecules and vice versa. For example, when we increase the kinetic energy (temperature) of liquids, the mobility (fluidity) of liquid increases.

viii. Density

Density is mass per unit volume. Liquids have higher densities than gases because in liquids the molecules are close to each other as compared to gases. Liquids molecules have strong intermolecular forces than gases, hence they cannot expand and have fixed volume.

Table 5.3 Densities of Various liquids

S. No.	Liquid	Density (g/cm^3)	Temperature (°C)
1.	Acetone	0.792	20
2.	Ethyl Alcohol	0.791	20
3.	Methyl Alcohol	0.810	20
4.	Olive oil	0.918	15
5.	Castor oil	0.969	15
6.	Linseed oil	0.942	15
7.	Water	1.00	4

Activity:5.3

Observe the diffusion of liquid bromine

Ans (11)

Solid State

In solid state of matter the particles (molecules, ions) are closely packed in a fixed pattern. In solids there occurs a strong force of attraction between the solid particles, which hold them firmly together, so that they cannot leave their position. Solid particles posses only the vibrational motion. Hence, solid cannot diffuse like gases and liquids.

5.4 Typical Properties of Solids

Some typical properties of solid state are;

(i) Volume and Shape

Solids have strong intermolecular forces present between their particles thus having a definite shape and a definite volume.

(ii) Melting Point

In solid state, particles (molecules, ions) are not free to move like gases and liquids, molecules can only vibrate. When solid is heated, the kinetic energy of molecules increases and the particles of the solid vibrate with higher speed. When the energy is continuously supplied, a stage is reached at which the particles leave their mean positions and start moving freely. The solid starts melting. The temperature at which the solid starts melting and exist in dynamic equilibrium with liquid state is called melting point. At melting point, temperature does not change and remains constant until the whole solid is converted into the liquid state.

(12) Ans

Melting point of a solid depends on the strength of attractive forces that hold particles together in the fixed positions. Melting point is the characteristic property of a crystalline solid by which purity of solid can be checked.

$$\text{Solid} \underset{\text{Freezing Point}}{\overset{\text{Melting Point}}{\rightleftharpoons}} \text{Liquid}$$

End (12)

(iii) Rigidity

The particles in solid are fixed and closely packed. The particles in solid neither move nor slide over their mean positions. Therefore, the solids are rigid in their structure. The solids resist the deforming force due to hard structure and strong intermolecular force.

S.Q

(iv) Density

Solids are denser than liquids and gases. In solid, the particles are closely packed and have no empty spaces between the particles. Their mass per unit volume is greater. Therefore, they have higher densities as compared to the other two states of matter. End (H)

Scientific Information

Comparison of Properties of Solids, Liquids and Gases

S#	Properties	Solids	Liquids	Gases
1.	Mass	Definite	Definite	Definite
2.	Shape	Definite	Acquires the shape of the container	Acquires the shape of the container
3.	Volume	Definite	Definite	Indefinite
4.	Compressibility	Not possible	Almost negligible	Highly compressible
5.	Fluidity	Cannot flow	Can flow	Can flow
6.	Rigidity	Highly rigid	Less rigid	Not rigid
7.	Diffusion	Slow	Fast	Very fast
8.	Space between particles	Most closely packed	Less closely packed	Least closely packed
9.	Inter-molecular force	Strongest	Slightly weaker than in solids	Negligible

5.5 Types of Solids

Solids can be classified into two types, based on the arrangement of particles. These are amorphous solid and crystalline solids.

Ans (13)

5.5.1 Amorphous Solids

Amorphous means shapeless. Solids in which the particles are not regularly arranged or in which the particles are not properly arranged in three dimensions are called amorphous solids. In simple words, we can say that amorphous solids are one that lacks ordered arrangement of its particles. Amorphous solids are hard like true solids but they do not have sharp melting point. They melt over a range of temperature. For example, Glass, wax, butter, plastics, rubber, cotton candy etc.

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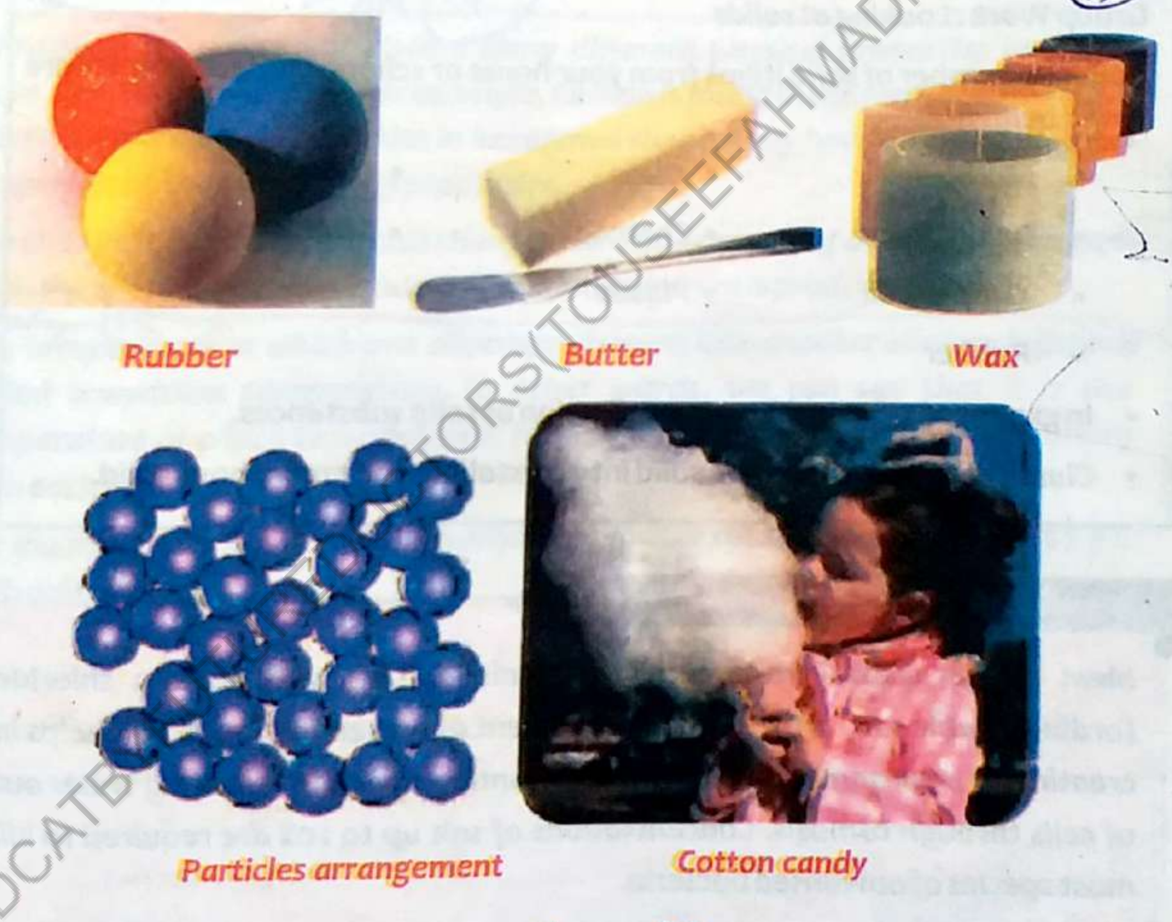


Fig 5.15 Amorphous solids

5.5.2 Crystalline Solids

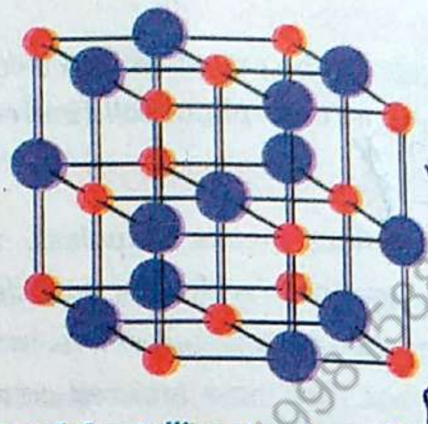


Fig. 5.16 Crystalline structure of NaCl

Crystalline solids are the solids in which particles (ions, atoms, or molecules) are arranged in regular three dimensional patterns. They have definite surfaces or faces. Each face has definite angle with the other face. Pure crystalline solids have sharp melting point. For example sodium chloride (NaCl), Naphthalene etc.

End 13

Activity: 5.4

Group Work: Looking at solids

Collect a number of solid items from your home or school. Some examples are listed below:

- *Marbles*
- *Baking soda*
- *Candle wax*
- *Rubber*
- *Glass*
- *Charcoal*
- *Plastic*
- *Table salt*
- *Lead pencil*
- *Eraser*

- *In groups of 5-8, combine your collection of solid substances.*
- *Classify the above mention solid into crystalline and amorphous solid.*

Society, Technology and Science

Meat or fish preserved or cured with salt. Salting with Sodium chloride (ordinary table salt) is the primary ingredient of preserving meat. Salt helps in creating an environment where bacteria cannot grow and removing water out of cells through osmosis. Concentrations of salt up to 20% are required to kill most species of unwanted bacteria.

5.6 Allotropy

Ans (14)

The existence of an element in more than one crystalline forms is called allotropy. The different forms are called allotropic forms.

The allotropy of an element is due to following reasons.

i. The existence of an element in more than one physical state or form, such as carbon (Kajol, soot, Diamond, Graphite etc.)

ii. The existence of two or more kinds of molecules of an element. In this case, each molecule has different number of atoms such as allotropes of oxygen are oxygen (O_2) and ozone (O_3).

iii. Different arrangement of two or more atoms or molecules in crystal of an element. For Example, Sulphur shows allotropy (monoclinic and orthorhombic forms) due to different arrangement of molecules (S_8) in the crystals. End (14)

Allotropes of an element always show different physical properties but have same chemical properties. For example, Carbon is found in the form of diamond in tetrahedral shape and graphite in hexagonal shape. They have different physical properties but same chemical properties.

The change in temperature also changes the arrangement of atoms in allotropes. With the change in temperature a new allotropic form is produced.

Ans (15)

The temperature at which one allotrope changes into another allotropic form is called transition temperature. In other words, we can say that it is the temperature at which two allotropic forms of an element co-exist in equilibrium with each other.

For example, the Sulphur S_8 orthorhombic form exists in equilibrium at 95.5°C with monoclinic form.

$$\underset{\text{(orthorhombic)}}{S_8} \xrightleftharpoons{95.5^\circ C} \underset{\text{(monoclinic)}}{S_8}$$

$$\underset{\text{(cubic)}}{\text{Grey tin}} \xrightleftharpoons{13.2^\circ C} \underset{\text{(tetragonal)}}{\text{white tin}}$$

$$\underset{\text{(rhombohedral)}}{\text{Potassium nitrate}} \xrightleftharpoons{128.5^\circ C} \underset{\text{(orthorhombic)}}{\text{potassium nitrate}}$$

End (15)

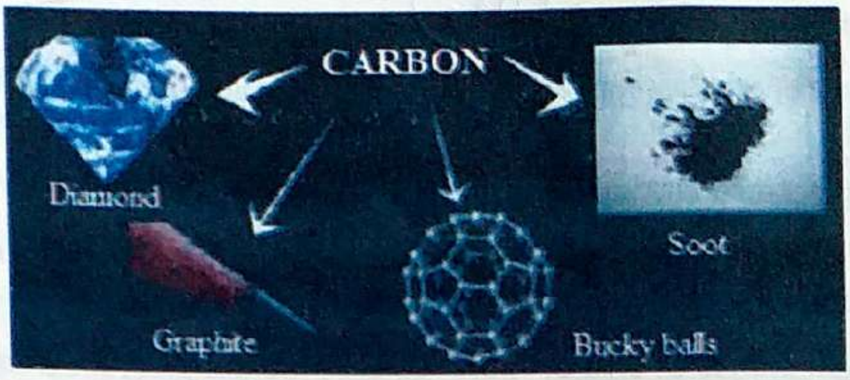


Fig. 5.17 Allotropes of Carbon

Allotropes of Carbon

Carbon exists in crystalline allotropic forms as well as non-crystalline or amorphous allotropic forms.

Carbon exists in three crystalline allotropic forms namely, diamond, graphite and bucky balls.

Carbon also exists in non-crystalline or amorphous allotropic forms such as coal, coke, charcoal, lampblack, etc. We will discuss the crystalline allotropic form in detail here.

Ans 16

(i) Structure of Diamond

Diamond is the crystalline allotropic form of carbon. In diamond, carbon exist in cubic form. Each carbon atom is tetrahedrally bonded by four covalent bonds with other carbon atoms. Covalent bond is very strong, so diamond is very hard and has high melting point. Diamond is bad conductor of electricity because all the four valence electrons are used in formation of covalent bonds, which tightly held with each other.

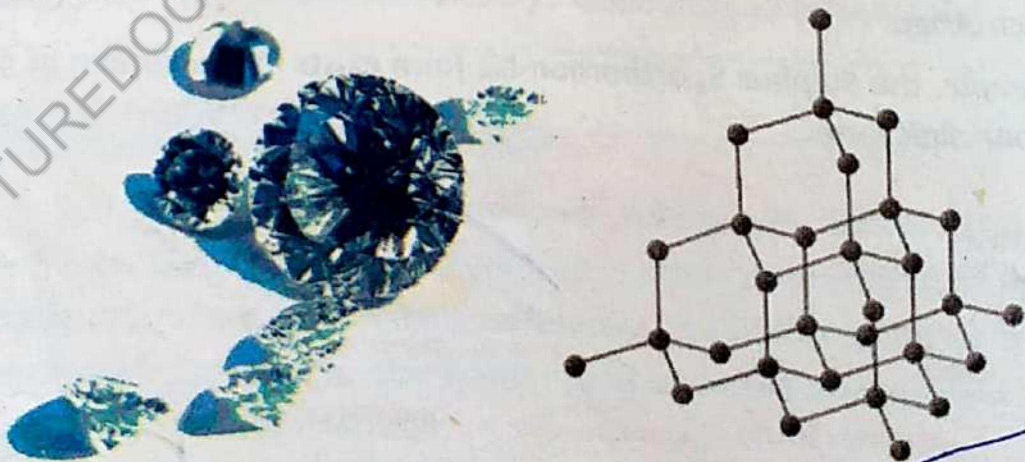


Fig. 5.18 Structure of Diamond

End 16

(ii) Structure of Graphite

In graphite, each carbon atom is only covalently bonded to three other carbon atoms rather than to four atoms as in diamond. Carbon exist in hexagonal form of sheets/ layers. These sheets linked with other carbon atoms by weak attractive forces. These sheets slide over each other.

Graphite is soft. It is used as lubricant in heavy machinery. It is good conductor of electricity because it has free electron available.

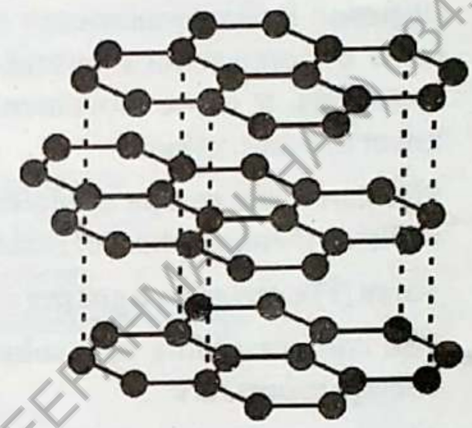


Fig. 5.19 Structure of Graphite

(iii) Structure of Buckyball

Buckyball is the crystalline allotropic form of carbon. It is recently discovered in 1985. Buckyball consists of 20 – 100 carbon atoms. In Buckyballs, the atoms are arranged in a hollow cage like structure. The carbon atoms link with each other and adopt the shape of football. In Buckyball, the carbon atoms joined together making pentagonal, hexagonal etc structures. Buckyballs are used as semiconductors, superconductors and lubricants.

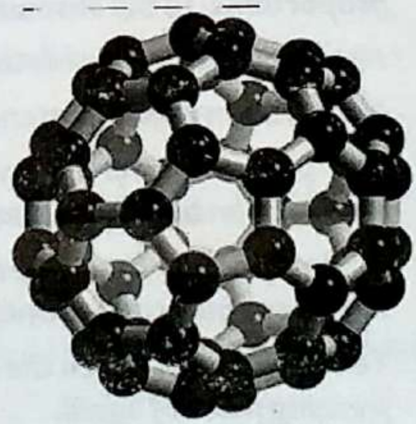


Fig. 5.20 Structure of Buckyball

Key Points

- Matter exist in three states solid, liquid and gas.
- The highest energy state of matter is called gas.
- Diffusion is the spontaneous mixing of the molecules of different gases to form a homogenous mixture by random motion and collision of the molecules. It is the movement of molecules from higher concentration to lower concentration.
- Effusion is the escape of molecules in the gaseous state one by one without collision through a hole of molecular dimension.
- Force (F) exerted by a gas per unit area (A) of a container is called pressure(P).
- The compressibility of a substance is the change in volume with per unit change in pressure
- Mass per unit volume is called density.
- According to Boyle's law volume (V) of a given mass of a gas decrease with the increase of pressure (P) at constant temperature.
- Charles's law states that, the volume of a given mass of gas is directly proportional to the absolute temperature at constant pressure.
- Absolute Temperature is the temperature on the Kelvin scale
- Evaporation is the spontaneous change of a liquid into the gaseous state.
- The pressure exerted by the vapours of the liquid when the rate of evaporation becomes equal to rate of condensation is called vapour pressure.
- Boiling point is the temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure.
- Temperature at which the liquid and solid state exists in equilibrium is called freezing point of liquid.
- Melting point is the temperature at which the solid starts melting and exists in dynamic equilibrium with liquid state.
- Amorphous solid is one that lacks ordered arrangement of its particles.

- *Crystalline solids are the solids in which particles (ions, atoms, or molecules) are arranged in regular three dimensional patterns.*
- *The existence of an element in more than one physical state and phase are called allotropic forms or allotropes and this phenomenon is called allotropy.*
- *The temperature at which one allotrope changes into to another allotropic form is called transition temperature.*

Exercise

Choose the correct option.

i. The spontaneous mixing of particles is called :

a. evaporation b. sublimation
c. diffusion d. boiling

ii. Which statement for the particles of solid is not correct:

a. they move at great speed
b. they are arranged in regular pattern
c. there is a very little space between the particles
d. the force of attraction between the particles are strong

iii. A liquid boils when its vapour pressure becomes equal to :

a. 760cm of Hg b. 1 Pascal
c. 101.325 kilo Pascal d. 0.1 atm

iv. The vapour pressure of a liquid increases with the :

a. increase of pressure
b. increase of temperature
c. increase of intermolecular forces
d. increase of polarity of molecules

v. Water normally boils at 100°C but it is possible to boil at 50°C which variable would you have to change to do this :

a. increase external pressure b. decrease external pressure
c. increase surface area d. decrease surface area

vi. The vapour pressure of a liquid in a closed container depends upon:

a.	amount of liquid	b.	surface area of the liquid
c.	temperature	d.	both (b) and ©

vii. Which one of the following is not an example of amorphous solid :

a.	rubber	b.	glass
c.	glucose	d.	plastic

viii. At freezing point which one of the following coexists in dynamic equilibrium :

a.	gas and solid	b.	liquid and gas
c.	liquid and solid	d.	all of these

ix. Ink spreads in water because of :

a.	vapour pressure	b.	expansion
c.	diffusion	d.	compressibility of water

x. What will be the pressure of a gas, if the volume of the gas at 2 atmospheres is increased from 1.5 dm^3 to 3dm^3:

a.	1atmosphere	b.	1.5 atmosphere
c.	2 atmosphere	d.	2.5 atmosphere

Solve the following numerical questions.

i. Calculate the final pressure of a sample of gas that is changed at constant temperature to 14.3 dm^3 from 7.55 dm^3 at 828 torr. (Ans: 437.160 torr)

ii. Calculate the final volume at 302 K of a 5.41 dm^3 sample of gas originally at 353 K if the pressure does not change. (Ans: 4.62 dm^3)

iii. Calculate the initial volume at 0°C of a sample of gas that is changed to 731 cm^3 by cooling to -14°C at constant pressure. (Ans: 770.513cm^3)

iv. A sample of a gas at room temperature occupies 0.80 dm^3 at 1.5 atm. What will be its volume when the pressure of the gas is raised to 2.1 atm? (Ans: 0.571 dm^3)

v. Calculate the final volume of 319°C of a sample of gas original 5.13dm^3 at 171°C, if the pressure does not change. (Ans: 9.57 dm^3)

SHORT QUESTIONS

Answer briefly to the following questions.

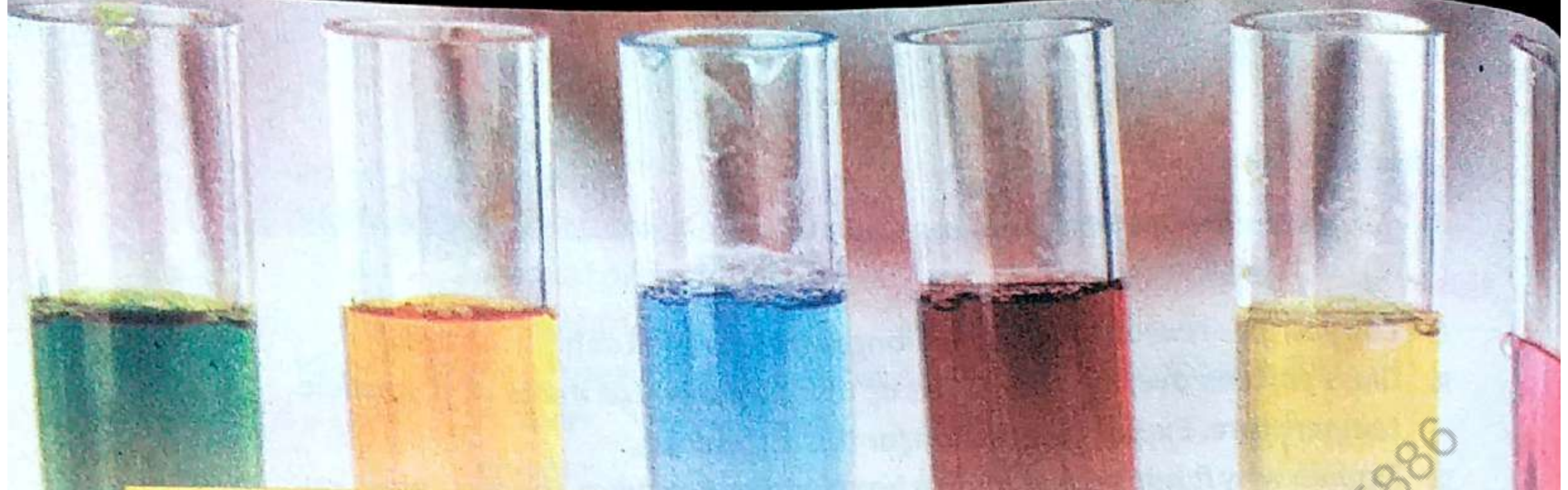
i. *Can you give reason why it takes longer time to cook at high altitude?*
ii. *Glass softens over a wide range of temperature. Ice melts at a specicic temperature. Explain the reason for this difference.*
iii. *Explain why it happens that on a hot summer day when there is sweat on the body of a person, one feels cool under fast moving fan?*
iv. *Why are the densities of gases lower than that of liquids?*
v. *What is the relationship between the atmospheric pressure and boiling point of a liquid?*
vi. *Why a gas is compressible but a solid is not compressible? Give reason(s).*

LONG QUESTIONS

i. *Define Boyle's law and verify it experimentally.*
ii. *Differentiate between*
    a. *Evaporation and Boiling point.*
    b. *Effusion and Diffusion of gases.*
    c. *Condensation and Evaporation*
iii. *Define the term allotropy with examples. Explain the three allotropic forms of carbon in detail.*
iv. *What are solids? Differentiate between amorphous and crystalline solids.*
v. *Define Charles's law and verify it graphically and diagrammatically.*

Project Work

Prepare separate lists of crystalline and amorphous solids found in your home. Compare your lists with those of your classmates and present in the class.



Unit 6 Solutions

After studying this unit, the students will be able to;

- *Define the terms: solution, aqueous solution, solute and solvent and give an example of each.*
- *Explain the difference between saturated, unsaturated and supersaturated solutions.*
- *Explain the formation of solutions (mixing gases into gases, gases into liquids, gases into solids) and give an example of each.*
- *Explain the formation of solutions (mixing liquids into gases, liquids into liquids, liquids into solids) and give an example of each.*
- *Explain the formation of solutions (mixing solids into gases, solids into liquids, solids into solids) and give an example of each.*
- *Explain what is meant by the concentration of a solution.*
- *Define Molarity.*
- *Define percentage solution.*
- *Solve problems involving the Molarity of a solution.*
- *Describe how to prepare a solution of given Molarity.*
- *Describe how to prepare dilute solutions from concentrated solutions of known Molarity.*
- *Convert the Molarity of a solution and its concentration in g/dm^3.*
- *Use the rule that "like dissolves like" to predict the solubility of one substance in another.*
- *Define colloids and suspensions.*
- *Differentiate between solutions, suspension and colloids.*

Introduction

Homogenous mixture of two or more than two components is called solution. Its composition can be changed within certain limits. Solutions are not limited to liquids state only. It is found in three physical states of matter. For example, alloy is a solid solution, sea water is a liquid solution and air is a gaseous solution.

In previous grades, you have learnt about the solutions and suspensions. In this unit, you will learn about different types of solutions, its concentration and units of measurement in detail.

6.1 Solution, Aqueous Solution, Solute and Solvent

(i) Solution

A homogenous mixture of two or more than two substances is called solution. A solution is usually made up of two components i.e. solvent and solute. A solution, which is made of two components, is called binary solution. For example, salt dissolved in water.

Test Yourself.

Suggest a method for differentiation between pure liquid and solution.

(ii) Aqueous Solution

The solution which is formed by dissolving a substance in water is called an aqueous solution. In aqueous solution, water is present in large amount and is termed as solvent. For example, salt in water and sugar in water. Water is polar covalent compound. Water has the ability to dissolve a large number of substances, that is why water is called as universal solvent. It dissolve majority of ionic and polar covalent substances.

(iii) Solute and Solvent

The component of solution which is present in smaller amount is called solute and the component of solution which is present in larger amount is called solvent. A solute is dissolved in solvent to make solution. For example, sugar solution is made by dissolving sugar in water. In this solution, sugar is solute and water is solvent. Beside this, there are a number of solutions which consist of more than two substances. For example, soft drinks, in this water is solvent and sugar, carbon dioxide (CO_2), flavours etc. are the solutes.

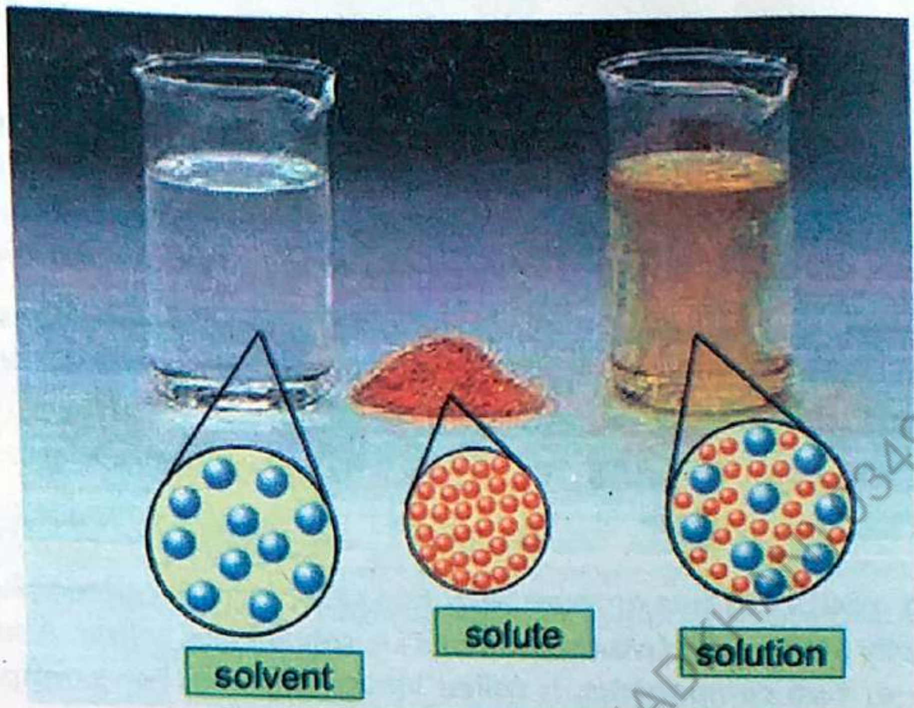


Fig. 6.1 Solvent and solute combine to form solution

6.2 Saturated, unsaturated and super saturated solution

(i) Saturated solution

When a small amount of sugar is added into water in a glass, the sugar dissolve very quickly in it. If some more amount of sugar is added, it will also dissolve in it. If the addition of sugar is kept on, a stage is reached when the further amount of solute will not dissolve in it. But still if we add further amount of sugar, it will remain undissolved and will settle down at the bottom of the glass.

Solution which cannot dissolve further amount of solute at a particular temperature is called saturated solution at that temperature.

Skill Development

(1) Prepare a saturated solution of NaCl in water at room temperature.
(2) Prepare solutions of different strengths

a) 2M solution of NaOH
- Weigh 20 g of NaOH on a physical balance
- Dissolve it in water and make the volume upto 250 mL in a volumetric flask.

b) 0.1 M solution of NaCl
- Weigh 5.85 g of NaCl on a physical balance
- Dissolve it in little water and make the volume upto 1000 mL

(ii) Unsaturated solution

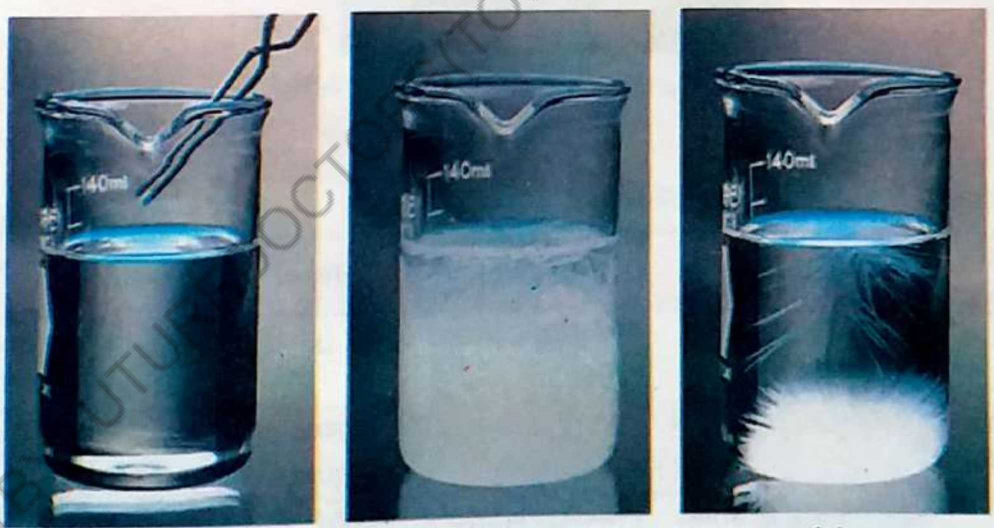
An unsaturated solution is a solution, which has less amount of solute than is required to saturate it at a particular temperature. For example, if we dissolve a small amount of sugar in water, this solution still has the capacity to dissolve more solute. This solution is an unsaturated solution.

A Solution which can dissolve further amount of solute at a particular temperature is called unsaturated solution.

(iii) Supersaturated solution

A solution which contains more amount of solute than the amount required to saturate it at a particular temperature is called supersaturated solution. For example, when a saturated solution is heated, it has the capacity to dissolve more amount of solute. Such solution contains more solute than is required to form a saturated solution and it becomes concentrated solution.

We can easily distinguish between unsaturated, saturated and super saturated solutions. When a crystal of already dissolved solute is added to each of the three forms of solution. If the crystal dissolves, the solution is unsaturated. If it remains un-dissolved, it is saturated solution and if some excess of the dissolved solute deposits on the crystal, then it is supersaturated solution.



(a) (b) (c)

Fig. 6.2 Unsaturated, saturated and supersaturated solutions

6.3 Types of Solution

Solutions exist in gas, liquid and solid states. On this basis, there are nine types of solutions.

6.3.1 Gaseous Solution (Gas in Gas, Gas in Liquid and Gas in Solid)

Different types of gaseous solution are following.

Type of solution	Solution	Solute	Solvent	Example
Gaseous Solutions	i. Gas in Gas	Oxygen(gas)	Nitrogen (gas)	Air (oxygen in Nitrogen)
	ii. Gas in Liquid	Carbon dioxide (gas)	Water (liquid)	Carbon dioxide in water, e.g. carbonated drink
	iii. Gas in Solid	Hydrogen (gas)	Palladium (solid)	Hydrogen gas adsorbed at palladium

Fig. 6.3: Carbon dioxide in water

6.3.2 Liquid Solutions (Liquid in Gas, Liquid in Liquid, Liquid in Solid)

Different types of liquid solutions are following.

Type of solution	Solution	Solute	Solvent	Example
Liquid Solutions	i. Liquid in Gas	Water vapours (liquid)	Air (gas)	Fog, humidity in air
	ii. Liquid in Liquid	Alcohol(liquid)	Water (liquid)	Alcohol in water
	iii. Liquid in Solid	Mercury(liquid)	Gold (solid)	Mercury amalgams, e.g amalgam used by dentist (Hg in Ag)

Fig. 6.4 Mercury Amalgam

6.3.3 Solids Solutions (Solid in Gas, Solid in Liquid, Solid in Solid)

Different types of solid solution are following.

Type of solution	Solution	Solute	Solvent	Example
Solid Solutions	i. Solid in Gas	Carbon particles (solid)	Air (gas)	Smoke (Carbon particles in air)
	ii. Solid in Liquid	Salts (solid)	Water (liquid)	Sea water, solutions e.g. sugar in water
	iii. Solid in Solid	Zinc (solid)	Copper (solid)	Brass (alloy), e.g. copper and zinc and bronze (alloy) of copper and tin

Fig. 6.5 Brass, Sterling silver and Bronze alloys

Society, Technology and Science

The substances we used in our life for existence are solutions. The solutions that we use are divided broadly into three categories, Gaseous solutions, liquid solutions and solid solutions.

Gaseous solution: Air is solution (mixture) of different gases. e.g. N_2, O_2, CO_2 and rare gases.

Liquid solution: We use different liquid solutions in our daily lives. For example, the water we use for drinking, cooking and washing is not in pure form. It contains dissolved minerals and gases that are essential for our health. Beverages, vinegar, soft drinks etc are liquid solutions. In addition to that products such as glass cleaners, sanitary cleaner, shampoo, etc are also liquids solutions. Most medicines that we use are produced in solution form.

Solid solution: Alloys are the solid solutions that we use in our daily lives. Jewelry is an alloy (solid solution) of gold and copper or silver. Because pure gold is very soft, so it cannot be used in its pure form. Gold are mixed with copper or silver in order to use it for jewelry. Brass and steel used for making utensils, surgical instruments, buses, cars trains, etc are solid solution of metals. Parts of airplane are made of solid solution of metals such as Al and Mg. Dental fillings are also solid solutions of metals in mercury.

6.4 Concentration of Solution

Concentration is the proportion of a solute in a solution. **The concentration of a solution is a measure or ratio of the amount of solute in a given amount of solvent or solution.** The larger the amount of solute present in a solution, the higher would be the concentration.

(i) Dilute solution

A solution, which contains small amount of solute dissolved in the solvent, is called a dilute solution.

(ii) Concentrated solution

A solution, which contains large amount of solute dissolved in the solvent, is called a concentrated solution.

Scientific Information

A solution of known concentration is called a standard solution.

For example, a solution in which 5g of sugar is dissolved in 100cm^3 of water is dilute as compared to a solution in which 8g of sugar is dissolved in 100cm^3 of water.

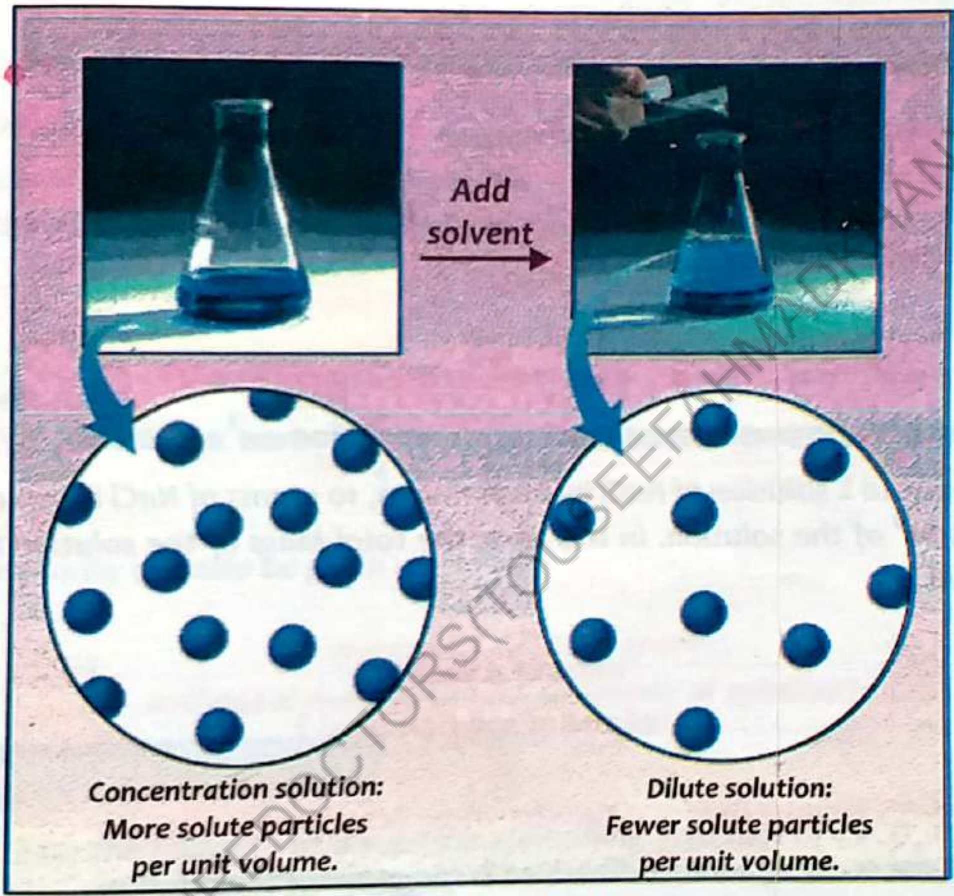


Fig.6.6 Dilution of Solution

There are many units to express the concentration of solutions. In this unit, we will study two units (i) Percentage composition (ii) Molarity.

6.4.1 Percentage composition

It is the number of parts of solute present in 100 parts of solution or **the fraction of a solute in a solution multiplied by 100.**

The percentage compositions of a solution are expressed in four ways:

i. Percentage Mass By Mass (Mass – Mass Relationship m/m):

It is the number of grams by mass of solute present in 100 grams by mass of a solution.

For example, 10 % solution of sugar by mass means, 10 grams of sugar in 90grams of solvent, so that the solution weighs 100 grams.

$$\%\ \text{mass / mass} = \frac{\text{mass of solute (g)}}{\text{mass of solute(g)} + \text{mass of solvent(g)}} \times 100$$

Or

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

ii. Percentage Mass By Volume (Mass – Volume Relationship m/v)

It is the number of grams by mass of solute present in 100 cm^3 of solution.

For example, 10 % solution of NaCl by mass means, 10 grams of NaCl in solvent to make 100cm^3 of the solution. In this case, the total Mass of the solution is not considered.

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

iii. Percentage Volume By Mass (Volume - Mass Relationship v/m)

It is the volume in cm^3 of a solute dissolved in 100grams of the solution.

For example, 10 % solution of Alcohol by volume means, 10 cm^3 of Alcohol in (unknown) volume of water so that the total mass of the solution is 100grams of solvent. In this case, the total volume of the solution is not considered.

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

iv. Percentage Volume By Volume (Volume - Volume Relationship v/v)

It is the volume in cm^3 of a solute dissolved per 100 cm^3 of the solution.

For example, 10 % solution of Alcohol by volume means, 10 cm^3 of Alcohol in sufficient volume of solvent, so that the volume of solution is 100cm^3.

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

6.4.2 Molarity

Molarity is a concentration unit. It is defined as **the number of moles of solute dissolved per dm^3 (liter) of solution.** It is represented by "M". The formula used for the preparation of molar solution is as follows.

$$Molarity(M) = \frac{number\ of\ moles\ of\ solute}{volume\ of\ solution\ in\ dm}$$

By definition

$$Mole = \frac{amount\ of\ solute\ in\ gram}{molecular\ mass\ of\ solute}$$

So molarity can also be given as,

$$M = \frac{amount\ of\ solute\ in\ gram}{molecular\ mass\ of\ solute \times volume\ of\ solution\ in\ dm^3}$$

EXAMPLE: 6.1

Calculate the molarity of a solution containing 7.50 mol of $CaCO_3$ in enough water to make 1.50 dm^3 of solution.

Solution $Molarity = \frac{7.50}{1.50dm^3} = \frac{5.0mol}{1dm^3} = 5.0M$

Scientific Information

$1dm^3 = 1\ liter = 1000cm^3 = 1000ml$

$1cm^3 = 1\ ml$

6.4.3 Problems Involving the Molarity of a Solution

The following examples will help you to understand the molarity and how molar solutions are prepared.

To find the molarity of a solution, you must know the molar mass of the solute.

For example, 1 molar solution of glucose ($C_6H_{12}O_6$) means 1 mole (180grams) of Glucose per dm^3 of solution. Similarly, "one-molar" solution of sodium hydroxide (NaOH), contains one mole of NaOH (40 grams) in one dm^3 of solution. As the symbol for molarity is M, and the concentration of a one-molar solution of sodium hydroxide is written as 1M NaOH.

One mole of NaOH has a mass of 40.0 g. If this quantity of NaOH is dissolved in enough water to make exactly 1.00dm^3 of solution, the solution is a 1M solution. If 20.0g of NaOH, is dissolved in enough water to make 1.00dm^3 of solution, a 0.500M NaOH solution is produced. This relationship between molarity, moles and volume may be expressed in the following ways.

$$M = \frac{\text{Amount of solute in gram}}{\text{molecular weight of solute} \times \text{volume of solution in } dm^3}$$

$$M = \frac{40}{40 \times 1dm^3} = 1mol$$

Or

$$M = \frac{20}{40 \times 1dm^3} = 0.500mol$$

If twice the molar mass of NaOH, i.e. 80 g, is dissolved in enough water to make 1dm^3 of solution, a 2M solution is produced.

EXAMPLE: 6.2

5.85 g of NaCl is dissolved in water so as to make 500 cm^3 solution. Determine the molarity of the solution. (Atomic mass of Na = 23g and Cl=35.5g)

Solution: Amount of NaCl = 5.85 g

Molar mass of NaCl = $23+35.5 = 58.5\, g\, mol^{-1}$

Moles of NaCl (solute) = $\frac{5.85\, g}{58.5\, gmol^{-1}} = 0.1\, mol$

Volume of solution in cm^3 = $500\, cm^3$

Volume of solution in dm^3 = $\frac{500}{1000} = 0.5\, dm^3$

$$M = \frac{\text{Moles of solute}}{\text{Volume of solution in } dm^3}$$

On putting the values,

$$M = \frac{0.1\, mol}{0.5\, dm^3} = 0.2\, M$$

It is 0.2 molar solution.

EXAMPLE: 6.3

Calculate the molarity of 50.0 cm^3 of solution containing 7.50 g of CH_3OH.

Solution

Because molarity is defined in terms of moles of solute and liters of solution, the given quantities can be converted to moles and liters, respectively:

$$\text{No. of moles of } CH_3OH = \frac{7.50g}{32.0g} = 0.2344\, mol\ CH_3OH$$

$$50.0cm^3 \left[\frac{1dm^3}{1000cm^3}\right] = 0.0500dm^3$$

$$M = \frac{0.2344 mol CH_3OH}{0.0500dm^3} = 4.69 M CH_3OH$$

Practice Problem: 6.1

Calculate the molarity of 11.6 cm^3 of solution containing 0.750 g of $CaCl_2$.

Activity: 6.1

The preparation of a 0.5000 M solution of $CuSO_4 \cdot 5H_2O$.

1

Start by calculating the mass of $CuSO_4 \cdot 5H_2O$ needed. Making a liter of this solution requires 0.5000 mol of solute. Convert the moles to mass by multiplying by the molar mass of $CuSO_4 \cdot 5H_2O$. This mass is calculated to be 124.8g.

2

Add some solvent to the solute to dissolve it and then pour it into a 1.0 dm^3 volumetric flask.

3

Rinse the weighing beaker with more solvent to remove all the solute and pour the rinse into the flask. Add water until the volume of the solution reaches the neck of the flask.

4

Put the stopper in the flask, and shake the solution thoroughly.

5

Carefully fill the flask to the 1.0dm^3 mark with water.

6

Restopper the flask and invert it at least 10 times to ensure complete mixing.

7

The resulting solution has 0.5000 mol of solute dissolved in 1.000 dm^3 of solution, which is a 0.5000 M concentration.

Scientific Information

To dilute a stock solution, dilution equation is used: $M_1V_1 = M_2V_2$. Where M_1 and V_1 are the molarity and volume of the concentrated stock solution and M_2 and V_2 are the molarity and volume of diluted solution you want to make.

EXAMPLE: 6.4

What will be the molarity of the solution if $30cm^3$ of 2.5M of $CaCO_3$ solution is diluted to $280cm^3$?

Solution:

$M_1 = 2.5M$, $V_1 = 30cm^3$, $V_2 = 280cm^3$, $M_2 = ?$

Or we can write,

$$M_2V_2 = M_1V_1$$

$$M_2 = \frac{M_1V_1}{V_2}$$

$$M_2 = \frac{2.5 \times 30}{280} = 0.267M$$

EXAMPLE: 6.5

What volume of water must be added to $85\ cm^3$ of 3.5 M Na_2CO_3 to dilute it to 0.41M?

Solution:

$M_1 = 3.5M$ $V_1 = 85cm^3$

$M_2 = 0.41M$ $V_2 = ?$

We can write,

$$M_1V_1 = M_2V_2$$

$$V_2 = \frac{M_1V_1}{M_2} = \frac{3.5 \times 85}{0.41} = 725.60cm^3$$

Volume of solution $= 725.60 - 85$

$= 640.60cm^3$

$640.60cm^3$ needs to be added to get the dilution.

6.5 Solubility

The solubility is defined as **the amount of solute in grams dissolved in 100 grams of the solvent to prepare a saturated solution at a given temperature.** The concentration of a saturated solution is referred to as solubility of the solute in a given solvent.

Different substances have different solubilities in the same amount of solvent at a specific temperature. For example, sodium nitrate ($NaNO_3$) is more soluble than silver chloride ($AgCl$) in water. Generally, the solubility of solute is taken to be the quantity required to make a saturated solution in given quantity of the solvent.

$$\text{Solubility} = \frac{\text{Wt. of the solute}}{\text{Wt. of the Solvent}} \times 100$$

Solubility of solutes depend on the following factors:

(i) Nature of the solvent

(ii) Nature of the solute

(iii) Pressure

(iv) Temperature

(i) Nature of solvent

When the molecules of solute are similar in structure and properties to the molecules of the solvent, the solubility is greater because like dissolve like. For example, sodium chloride is an ionic compound. It has greater solubility in a polar solvent like water but low solubility (insoluble) in non-polar solvent like benzene.

(ii) Nature of the solute

Nature of solute also affects the solubility. If a solute is changed and the solvent remain the same, the solubility of the solute also changes. For example, sodium chloride has high solubility in water and sugar has comparatively low solubility in water.

"LIKE DISSOLVE LIKE" is the general principle of solubility. It means that:

i. The ionic and polar substances are soluble in polar solvents. Ionic solids and polar covalent compounds are soluble in water e.g. $NaCl$, KCl, Na_2CO_3, sugar, glucose and alcohol are soluble in water.

ii. Non-polar substances are insoluble in polar solvents. Non-polar covalent compounds are insoluble in water such as benzene and petrol is insoluble in water.

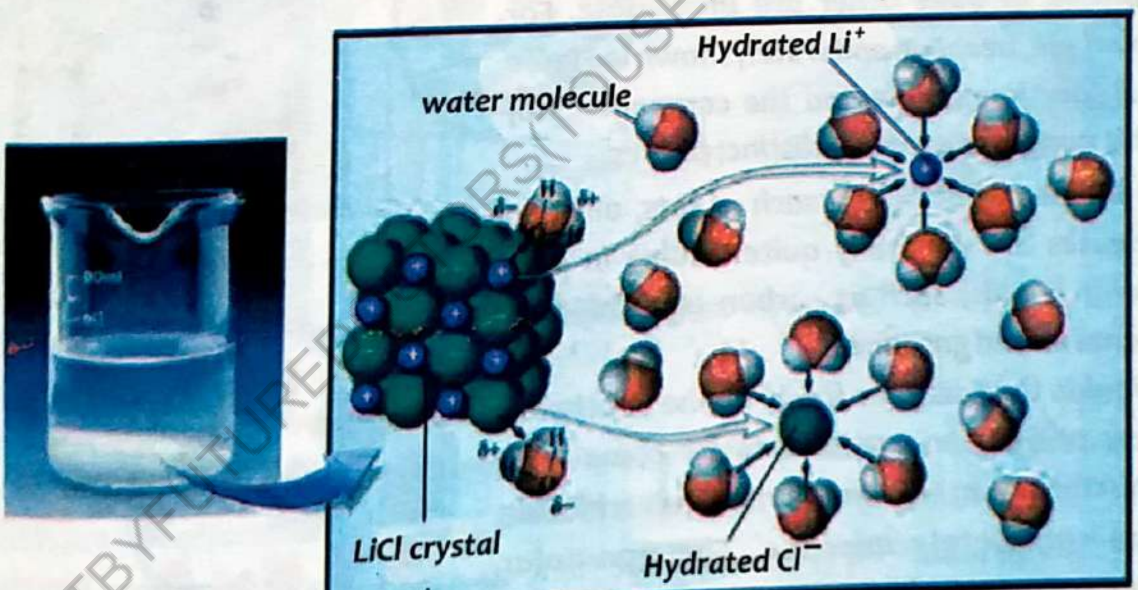
iii. Non-polar covalent substances are soluble in Non-polar solvents. Grease, paints are soluble in petrol, ether or carbon tetrachloride etc.

Solubility and Solute - Solvent Interaction

Lithium chloride is highly soluble in water but gasoline is not. On the other hand, gasoline mixes readily with benzene (C_6H_6), but lithium chloride does not. Why are there such differences in solubility? **"Like dissolves like"** is a rough but useful rule for predicting whether one substance will dissolve in another or not.

(i) Dissolving Ionic Compounds in Polar Solution

The polarity of water molecules plays an important role in the formation of solutions of ionic compounds in water. The charged ends of water molecules attract the ions in the ionic compounds and surround them to keep them separated from the other ions in the solution. For example, we add a few crystals of lithium chloride into a beaker of water. The water molecules come into contact with Li^+ and Cl^- ions. The positive ends of the water molecules are attracted to Cl^-



When LiCl dissolves, the ions are hydrated. The attraction between ions and water molecules is strong enough that each ion in solution is surrounded by water molecules.

Fig. 6.7: Solubility and Solute - Solvent Interaction

ions, while the negative ends are attracted to Li^+ ions. The attraction between water molecules and the ions is strong enough to draw the ions away from the crystal and form solution, as illustrated in Figure 6.7. This solution formation process with water as the solvent is referred to as **hydration**. These ions are said to be hydrated. As hydrated ions diffuse into the solution, other ions are exposed and are drawn away from the crystal surface by the solvent. The entire crystal gradually dissolves and hydrated ions become uniformly distributed in the solution.

(ii) Dissolving Ionic Compounds in Non-polar Solvents

Ionic compounds are generally insoluble in non-polar solvents such as carbon tetrachloride (CCl_4) and benzene (C_6H_6). The non-polar solvent molecules do not attract the ions of the crystal strongly enough to overcome the forces holding the crystal together.

Lithium chloride (LiCl) is not soluble in benzene. Lithium chloride (LiCl) and benzene (C_6H_6) differ widely in bonding, polarity and intermolecular forces.

(iii) Liquid Solutes and Solvents

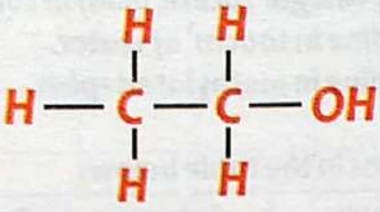
Liquid solutes and solvents that are not soluble in each other are immiscible. For example, benzene and water, shown in Figure 6.8, are immiscible and the components of this system exist in two distinct phases.

Non-polar substances, such as fats, oils and greases are generally quite soluble in non-polar liquids, such as carbon tetrachloride, benzene and gasoline.

Liquids that dissolve freely in one another in any proportion are said to be completely miscible. Benzene and carbon tetrachloride are completely miscible. The non-polar molecules of these substances exert no strong forces of attraction or repulsion and the molecules mix freely with one another.

Fig. 6.8: Liquid Solutes and Solvents - immiscible

Ethanol and water, shown in Figure 6.9, also mix freely. The – OH group on an ethanol molecule is somewhat polar. This group can form hydrogen bonds with water as well as with other ethanol molecules. The intermolecular forces in the mixture are so similar to those in the pure liquids that the liquids are mutually soluble in all proportions. The components of this system exist in a single phase with a uniform arrangement. Hydrogen bonding between the solute and solvent enhances the solubility of ethanol in water.



C_2H_5OH

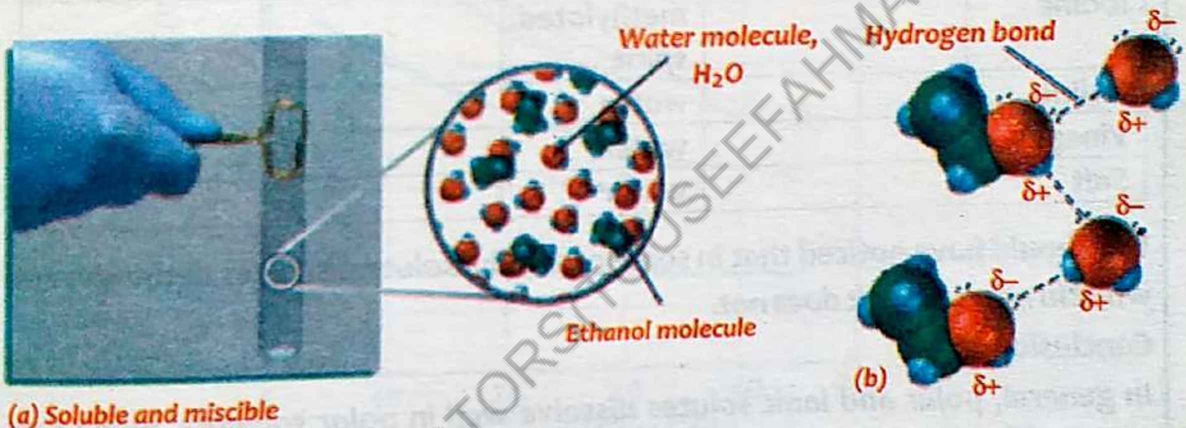


Fig. 6.9: Liquid Solutes and Solvents – miscible

(iii) Effect of pressure on solubility

Solids and liquids are incompressible therefore the solubilities of solids and liquids are not affected by changing the pressure.

Solubility of gases increases with the increase in pressure, for example, carbon dioxide is fiilled in soda water bottles under pressure. When a bottle of soda water is opened, carbon dioxide comes out with effervescence (bubbles) because pressure in the bottle is released resulting in decrease in the solubility of the gas.

Activity:6.2

Aim:

To investigate the solubility of solutes in different solvents.

Chemicals:

Common Salt, vinegar (Acetic Acid), iodine, methylated spirit

Method:

1. Mix half a teaspoon of common salt in 100cm^3 of water.
2. Mix half a teaspoon of vinegar (acetic acid) in 100cm^3 of water.
3. Mix a few grains of iodine in 100cm^3 of water.
4. Mix a few grains of iodine in methylated spirit.

Results:

Record your observations in the table below:

Solute	Polar, non-polar or ionic solute	Solvent	Polar, non-polar or ionic solvent	Does solute dissolve?
Iodine		methylated spirit		
Iodine		water		
Vinegar		water		
Salt		water		

You should have noticed that in some cases, the solute dissolves in the solvent, while in other cases it does not.

Conclusions:

In general, polar and ionic solutes dissolve well in polar solvents, while non-polar solutes dissolve well in non-polar solvents. An easy way to remember this is that 'like dissolves like', in other words, if the solute and the solvent have similar intermolecular forces, there is a high possibility that dissolution will occur.

Point to Ponder

- *If you pour a handful of salt into a full glass of water, the water level will actually go down rather than overflowing the glass. Similarly, if you mix half a liter of alcohol and half a liter of water, the total volume of the liquid will be less than one liter. Explain why?*

(iv) Effect of Temperature on Solubility

The solubility of many substances is affected with temperature. Increasing the temperature, usually decreases gas solubility. As the temperature increases, the average kinetic energy of the molecules in solution increases. A greater number of solute molecules are able to escape from the attraction of solvent molecules and return to the gas phase. At higher temperatures, gases are generally less soluble, as shown in Figure 6.11.

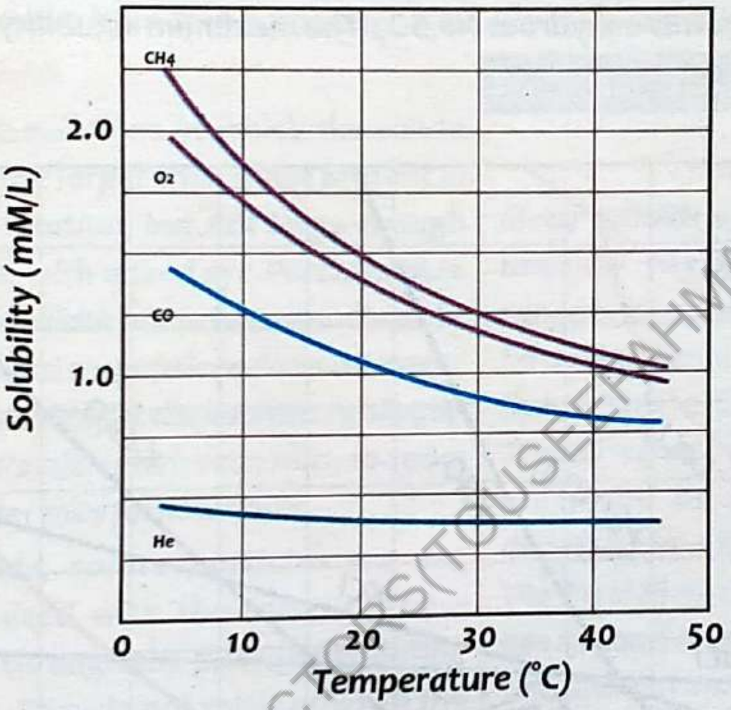


Fig. 6.10: Effect of Temperature on Solubility of Different Gases

The effect of temperature on the solubility of solids in liquids is more difficult to predict. Generally it appears that solubility of solids increases with increase in temperature, but this not always happen. When a solute (salt) is added into solvent, there are different possibilities with reference to effect of temperature on solubility as shown in the figure 6.11. These possibilities are given below.

(i) The solubility of some solutes increases with rise in temperature. For example, the solubility of Potassium Nitrate (KNO_3), $CaCl_2$ and $Pb(NO_3)_2$.

(ii) The solubility of some solutes decreases with increase in temperature, for example, $Ce_2(SO_4)_3$, Li_2CO_3 and CaO.

(iii) The solubility of the NaCl and KBr is not affected by increase or decrease in temperature and remains constant.

(iv) There are some solids whose solubility increases up to a certain temperature and then decreases on further increases in temperature. For example, Sodium Sulphate, forms deca-hydrate ($Na_2SO_4.10H_2O$) at 32.4°C. Above this temperature, it forms anhydrous Na_2SO_4. The maximum solubility of Sodium sulphate is at 32.4°C.

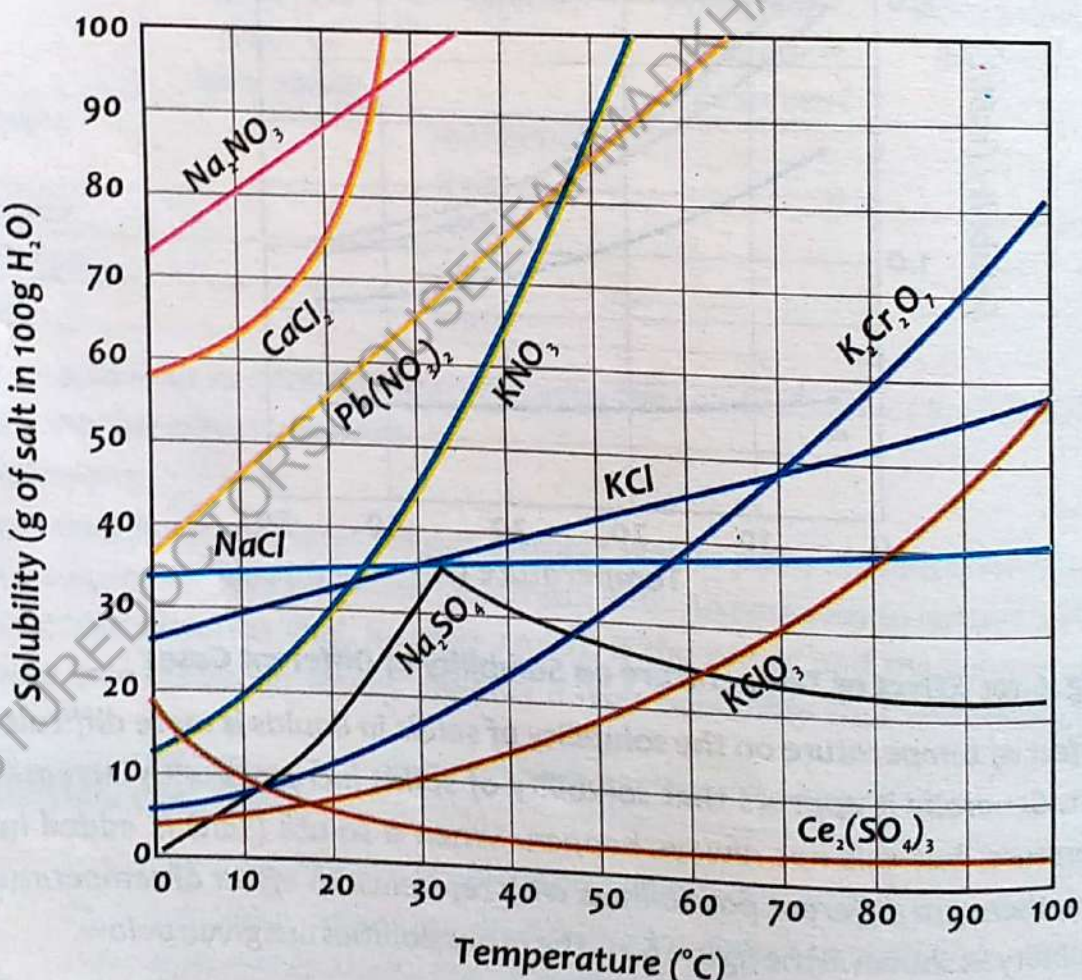


Fig. 6.11: Effect of temperature on solubility of different salts in water

6.6 Comparison of Solution, Suspension and Colloids

6.6.1 Solution

A solution is a homogeneous mixture of two or more components. This is an example of true solution. Each component is mixed in such a way that these components cannot be identified easily. The simplest example is the salt solution. You cannot differentiate between the solute and solvent they are mixed in such a way that they form a uniform mixture. Its particles cannot be seen with naked eye. They can pass through the filter paper during filtration. They do not settle down at the bottom of the container, If allowed to stand even for longer time.

6.6.2 Colloids

These are solutions in which the solute particles are larger than those present in the true solution, but not large enough to be seen with naked eye. **Particles that are intermediate in size between those in solutions and suspensions form mixtures known as colloidal dispersions or simply colloids.** Particles between 1nm–1000nm in diameter may form colloids.

In colloids, solute particles are not homogenized with the solvent. These can pass through the filter paper during filtration. They do not settle down at the

Scientific Information

Tyndall Effect

Many colloids appear homogeneous because the individual particles cannot be seen. The particles are, however, large enough to scatter light. This effect, known as the Tyndall effect, occurs when light is scattered by colloidal particles dispersed in a transparent medium. The Tyndall effect is a property that can be used to distinguish between a solution and a colloid.

Fig. 6.12 Examples of colloidal systems from daily life

bottom of the container, if allowed to stand for some time. For example, a mixture of starch in water is an example of colloid. Examples of the various types of colloids are given in Table 6.1.

Table 6.1 Classes of Colloids

Class of colloid	Phases	Example
Sol	solid dispersed in liquid	paints, mud
Gel	solid network extending throughout liquid	gelatin
Liquid emulsion	liquid dispersed in a liquid	milk, mayonnaise
Foam	gas dispersed in liquid	shaving cream, whipped cream
Solid aerosol	solid dispersed in gas	smoke, airborne particulate matter, auto exhaust
Liquid aerosol	liquid dispersed in gas	fog, mist, clouds, aerosol spray
Solid emulsion	liquid dispersed in solid	cheese, butter

EXAMPLE: 6.6

Which solvent—liquid ammonia (NH_3) or benzene (C_6H_6) is more likely to dissolve each of the following solutes?

(a) H_2O (b) C_6H_{12} (c) AgCl

Solution

a. H_2O is more likely to dissolve in NH_3 because both substances are polar and capable of hydrogen bonding.

b. C_6H_{12} is more likely to dissolve in the nonpolar solvent, (C_6H_6).

c. AgCl is more likely to dissolve in the polar solvent, (NH_3).

Practice Problem: 6.2

Which type of solvent , polar or non polar, is most likely to dissolve methyl alcohol, CH_3OH?

6.6.3 Suspension (Turbidity)

A suspension is a heterogeneous mixture of undissolved particles in a given medium. Its composition is not uniform throughout. In suspension, the solute particles are large. It can be seen with the naked eye. These suspended particles settle down at the bottom if left and allowed to stand for some time. For example, if we stir mud or sand in water. It mixes in water. When left for some time, a layer of particles settle down at the bottom and clear liquid remains at the top. The process of settling of particles at the bottom is known as sedimentation. The clear liquid at the top can be decanted (poured off) from the top. The particles cannot pass through filter paper and stay on paper during filtration. For example, a mixture of chalk water is suspension.

Fig. 6.13: Suspension (Turbidity) and solution

Beside this, other examples of suspension are medicines such as Antiacid, Antibiotic and Paracetamol etc. It is necessary to shake these medicines before use.

The distinctive properties of solutions, colloids and suspensions are summarized in Table 6.2

TABLE 6.2 Properties of Solutions, Colloids and Suspensions

S. No.	Solutions	Colloids	Suspensions
1.	Homogeneous	Heterogeneous	Heterogeneous
2.	Particle size: 0.01–1 nm; can be atoms, ions, molecules	Particle size: 1–1000 nm, dispersed; can be aggregates or large molecules	Particle size: over 1000 nm, suspended; can be large particles or aggregates
3.	Do not separate on standing	Do not separate on standing	Particles settle down
4.	Cannot be separated by filtration	Cannot be separated by filtration	Can be separated by filtration
5.	Do not scatter light	Scatter light (Tyndall effect)	May scatter light, but are not Transparent
6.	Particles are so small that they can't be seen with naked eye	Particles are big but cannot be seen with naked eye	Particles are big enough to be seen with naked eye

Solutions
Table salt in water is an example of solution.

Colloids
Milk in water is an example of colloids.

Suspensions
Flour in water is an example of suspensions.

Fig. 6.14: Comparison of solution, colloids and suspension

Key Points

- A solution is a homogenous mixture of two or more than two substances.
- Aqueous solution is that solution, which is formed by dissolving a substance in water.
- Solute is that component of solution which is present in smaller amount.
- Solvent is that component of solution which is present in larger amount.
- Saturated solution cannot dissolve further amount of solute at a particular temperature.
- Unsaturated solution can dissolve further amount of solute at a particular temperature.
- Supersaturated solution contains more amount of solute than the amount required to saturate it at a particular temperature.
- Solutions can exist in gas, liquid and solid states, on this basis, there are nine types of solution.
- Concentration of a solution is a measure of the amount of solute in a given amount of solvent or solution.
- A dilute solution is that solution which contains small amount of solute dissolved in the solvent.
- A concentrated solution is that solution which contains large amount of solute dissolved in the solvent.
- Percentage is the number of parts of solute present in 100 parts of solution.
- Water is a universal solvent. It dissolves majority of compounds due to its stronger interactions with solute particles.

- Molarity is the number of moles of solute dissolved per dm^3 (liter) of solution.
- Solubility is the amount of solute in grams dissolved in 100 grams of the solvent to prepare a saturated solution at a given temperature.
- Liquid solutes and solvents that are not soluble in each other are called immiscible.
- Liquids that dissolve freely in one another in any proportion are said to be completely miscible.
- Colloidal solution is that solution which has particles that are intermediate in size between solutions and suspensions to form mixtures. Particles between 1 nm – 1000 nm in diameter may form colloidal dispersions or simply colloids.
- Tyndall effect occurs when light is scattered by colloidal particles dispersed in a transparent medium
- A suspension is a heterogeneous mixture of undissolved particles in a given medium. In suspension the particles are large. It can be seen with the naked eye.
- The process of settling of particles at the bottom is known as sedimentation.

Exercise

Choose the correct option.

i. Which of the following solution is more dilute:
   a. 1.0M b. 0.5M
   c. 0.05M d. 0.005M

ii. Milk is an example of :
   a. Solution b. Saturated solution
   c. Colloids solution d. Suspension

iii. Water droplet in air is an example of solution:
   a. Gas in gas b. Gas in liquid
   c. Liquid in gas d. Liquid in liquid

iv. When there is a low concentration of solute in a solution, it is known as:
   a. dilute solution b. saturated solution
   c. concentrated solution d. super saturated solution

v. What is the molarity of a $NaNO_3$ solution made by diluting 250.0 cm^3 of a 1.60 M solution to a final volume of 400cm^3:
   a. 1.20 M b. 1.00 M
   c. 0.200 M d. 0.160 M

vi. What is the concentration, in % mass by volume (m/v) of a solution containing 15.0 g KCl in 600.0 cm^3 solution:
   a. 5.00% b. 2.00%
   C. 0.200% d. 2.50%

vii. When KCl dissolves in water, the following will be produced:
   a. K and Cl b. K^+ and Cl^-
   c. K and Cl_2 d. K^+ and Cl_2

viii. 2 moles of Na_2SO_4 are dissolved in one dm^3 of solution. Molarity of solution is:
   a. 1 M b. 2 M
   c. 3 M d. 0.5 M

ix. Molarity is the number of moles of solute dissolved in:
   a. 1 kg of solvent b. 1 kg of solution
   c. 1 dm^3 of solvent d. 1 dm^3 of solution

x. The molarity of a NaOH solution by dissolving 4 g of it in 250 ml water is:
   a. 0.4 M b. 0.8 M
   c. 0.2 M d. 0.1 M

Solve the following numerical questions.

I. What is the molarity of a solution composed of 5.85 g of potassium iodide (KI) dissolved in enough water to make 0.125 dm^3 of solution? (Ans: 0.281M)

ii. How many moles of H_2SO_4 are present in 0.500 dm^3 of 0.150 M H_2SO_4 solution? (Ans: 0.075 moles)

iii. Suppose you wanted to dissolve 40.0 g NaOH in enough H_2O to make 6.00 dm^3 of solution.

a. What is the molar mass of NaOH? (Ans: 40)

b. What is the molarity of this solution? (Ans: 0.166 M)

iv. What is the molarity of a solution of 14.0 g NH_4Br in enough H_2O to make 150 cm^3 of solution? (Ans: 0.9533 M)

v. Suppose you want to produce 1.00 dm^3 of 3.50 M solution of H_2SO_4.

a. What is the solute? (Ans: H_2SO_4)

b. What is the solvent? (Ans: H_2O)

c. How many grams of solute are needed to make this solution? (Ans: 343 g)

SHORT QUESTIONS

Answer briefly the following questions.

i. Is seawater a solution? How would you prove with a simple experiment whether it is pure water or a solution?

ii. A bottle in a drug store contains a label "3 percent hydrogen peroxide." What does it mean?

iii. Classify the following as a solution, colloid or suspension and explain why:

(i) Milk (ii) Hot cup of tea (iii) Orange juice with pulp
(iv) Mayonnaise (v) Listerine mouthwash (vi) Milk of Magnesia
(vii) Cheese (viii) Mist (ix) Bottled water

iv. Why we stir paints thoroughly before using it?

v. Why suspensions and solutions do not show Tyndall Effect, while colloids do?

LONG QUESTIONS

i. Define solution? Explain types of solution on the basis of states of matter.

ii. (a) Discuss the solubility of a substance?

(b) Explain the factors that are responsible for the solubility of a substance?

iii. (a) What is the difference between a concentrated and a dilute solution? Give examples of each.

(b) Differentiate between unsaturated, saturated and supersaturated solutions.

iv. Describe one way to prove that a mixture of sugar and water is a solution and that a mixture of sand and water is not a solution.

v. Explain the following concentration units.

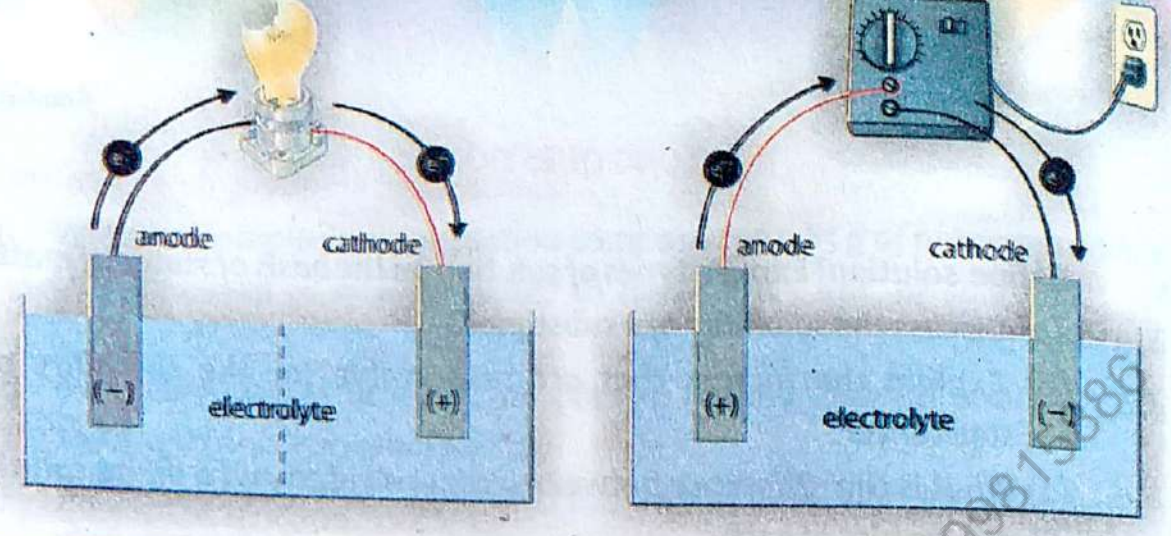
a. Percentage composition

b. Molarity

Project Work

Which of the following substances are pure substances and which are mixtures?

a. Sea water
b. Cooking oil
c. Coal
d. Steel
e. Chilli sance
f. Bronze
g. Sand
h. Gold
i. Orange squash
j. Salt
k. Milk
l. Oxygen gas
m. Air
n. Orange juice
o. Coke
p. Soil



Unit 7 Electrochemistry

After studying this unit, the students will be able to;

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen.
- Define oxidation and reduction in terms of loss or gain of electrons.
- Identify the oxidizing and reducing agents in a redox reaction.
- Define the oxidizing and reducing agent in a Redox Reaction.
- Define oxidation state.
- State the common rules used for assigning oxidation numbers to free elements, ions (simple and complex), molecules, atoms.
- Determine the oxidation number of an atom of any element in a compound.
- Describe the nature of electrochemical processes.
- Sketch an electrolytic cell, label the cathode and the anode.
- Identify the direction of movement of cations and anions towards respective electrodes.
- List the possible uses of an electrolytic cell.

- *Sketch a Danniell cell, labeling the cathode, the anode and the direction of flow of the electrons.*
- *Describe how a battery produces electrical energy.*
- *Identify the half-cell in which oxidation occurs and the half-cell in which reduction occurs given a voltaic cell.*
- *Distinguish between electrolytic and voltaic cells.*
- *Describe the methods of preparation of alkali metals.*
- *Describe the manufacture of sodium metal from fused NaCl.*
- *Identify the formation of by products in the manufacture of sodium metal from fused NaCl.*
- *Describe the method of recovering metal from its ore.*
- *Explain electrolytic refining of copper.*
- *Define corrosion.*
- *Describe rusting of iron as an example of corrosion.*
- *Summarize the methods used to prevent corrosion.*
- *Explain electroplating of metals on steel (using examples of zinc, Tin and chromium plating).*

Introduction

In electrochemistry we study about the chemical changes, which take place when electric current is passed through a particular type of material. We also study those chemical reactions which produce electric current. Electrochemistry deals with the interconversion of electrical energy and chemical energy. The applications of electrochemistry are widespread.

What is common in rusting of iron, combustion of fuel and food metabolism in human and animal bodies? All these processes involve oxidation – reduction reaction. These reactions involve transfer of electrons.

In previous grades, you have learnt about the chemical reactions. In this unit, you will learn about oxidation – reduction reactions. You will also learn that how electrical energy is converted into chemical energy and chemical energy into electrical energy.

7.1 Oxidation and Reduction

Electrons are lost in oxidation and gained in reduction. Therefore, when oxidation occurs during a chemical reaction, reduction must occur simultaneously. The number of electrons lost in oxidation must equal the number of electrons gained in reduction. This name is often shortened to redox reaction.

Oxidation

a) The addition of oxygen to a substance is called oxidation. For example carbon is oxidized to CO_2 which involves the addition of oxygen.

$$C + O_2 \longrightarrow CO_2$$

b) The removal of hydrogen or other electropositive elements from a substance is called oxidation. For example ammonia (NH_3) oxidized to nitrogen by losing hydrogen.

$$2NH_3 + 3Cl_2 \longrightarrow N_2 + 6HCl$$

c) The removal of electron(s) from a substance is also called oxidation. For example, when Fe^{+2} loses electron and oxidizes to Fe^{+3}.

$$Fe^{+2} \longrightarrow Fe^{+3} + e^-$$

Reduction

a) The removal of oxygen from a substance is called reduction. For example HgO is reduced to Hg on heating.

$$2HgO \xrightarrow{Heat} 2Hg + O_2$$

Scientific Information

Fireflies contain special cells in their abdomens that produce light. These cells contain a chemical called luciferin, which undergoes oxidation and produces light when the firefly takes in oxygen.

b) The addition of hydrogen to a substance is called reduction. For example nitrogen is reduced to ammonia (NH_3).

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

c) The addition of electron(s) to a substance is also called reduction. For example Sn^{+4} gain two electrons and reduces to Sn^{+2}.

$$Sn^{+4} + 2e^- \longrightarrow Sn^{+2}$$

Consider the reaction between sodium (Na) and chlorine (Cl_2). This reaction involves the transfer of electrons from sodium atom to chlorine atom. Sodium (Na) atom loses an electron and is said to be oxidized to sodium ion (Na^+), while chlorine (Cl) atom gains an electron and is reduced to chloride (Cl^-) ion. Those reactions in which, gain and loss of electrons takes place simultaneously is called an oxidation-reduction or Redox reaction.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

7.2 Oxidation States and Rules for Assigning Oxidation States

(i) Oxidation States or Oxidation Number

It is the apparent charges, positive or negative, on an atom of an element in a molecule or an ion. Unlike ionic charges, oxidation numbers do not have an exact physical meaning. In fact, in some cases they are quite arbitrary. However, oxidation numbers are useful in naming compounds, in writing formulas and in balancing chemical equations. The color of solutions changes with the change of oxidation state.

Fig. 7.1 Color change of chromium solution with change of oxidation state

(ii) Rules for assigning oxidation number

i. The oxidation number of all elements in the free state is zero. For example, the oxidation number of H_2, Cl_2 Na, etc. is zero.

ii. The oxidation number of a simple ion is the same as the charge on it. For example, oxidation numbers of Na^+, Ca^{+2}, Al^{+3} and Br^- are +1, +2, +3 and -1 respectively.

iii. The oxidation number of hydrogen in its compounds is +1 except in the case of metal hydrides, which is −1 e.g. $Na^{+1}H^{-1}$

iv. The oxidation number of oxygen in its compounds is -2 except in the case of peroxide, where it is -1 for example H_2O_2 and in the case of OF_2, it is +2.

v. The oxidation number of each element of group I, II and III is +1, +2 and +3 respectively.

vi. The oxidation number of each element of group VII (halogens) in their binary compounds is -1.

vii. In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero. For example H_2SO_4

$$2H + S + 4(-2) = 0$$
$$2 \times 1 + S - 8 = 0$$
$$2 + S - 8 = 0$$

we can write $S - 8 + 2 = 0$

$$S - 6 = 0 \qquad S = 6$$

Thus $\overset{+2}{H_2} - S^{+6} - O_4^{-8}$

viii. In ions, the algebraic sum of oxidation numbers equal to the charge on the ion. For example MnO_4^{-1}

$$MnO_4^{-1}$$
$$Mn + 4O = -1$$
$$Mn + 4(-2) = -1$$
$$Mn + 8 = -1$$
$$Mn = -1 + 8 \qquad Mn = 7$$

ix. In any substance the more electronegative atom has the negative oxidation number.

x. The same element may show different oxidation number in different compounds e.g. $CO(C^{+2} - O^{-2})$, $CO_2(C^{+4} - O_2^{-4})$

EXAMPLE: 7.1

Determine the oxidation numbers of phosphorus and chlorine in phosphorus pentachloride, PCl_5.

Solution

In this compound, the oxidation number of each chlorine atom is –1.

To make the total of all the oxidation numbers for PCl_5 equal to zero, the phosphorus must have an oxidation number of +5. To do this calculation systematically,

let x equal the oxidation number of phosphorus in PCl_5. Then

$$x \ +5\,(-1) \ = 0$$

Oxidation number of phosphorus | Number of chlorine atoms | Oxidation number of each chlorine atom | Charge on the molecule

$$x-5=0$$
$$x=+5$$

Thus, the oxidation number of phosphorus is 5 and chlorine is –1.

EXAMPLE: 7.2

What is the oxidation number of nitrogen in the nitrite ion, No_2^-?

Solution

The oxidation number of oxygen in most of its compounds is –2. The oxidation numbers of the two oxygen atoms plus that of the nitrogen atom must total –1, the charge on the ion:

$$x \ + \ 2\,(-2) \ = \ -1$$

Oxidation number of nitrogen | Number of oxygen atoms | Oxidation number of each oxygen atom | Charge on the nitrite ion

$$x-4 = -1$$
$$x = -1+4$$
$$x=+3$$

So, the ber of Nitrogen in nitrite is + 3.

Practice Problem: 7.1

- Determine the oxidation number of carbon in CO.
- What is the oxidation number of sulphur in the sulphite ion, SO_3^{-2}

Scientific Information

Most obvious characteristics chemical property of the transition metals is the occurrence of variable oxidation states. Mostly transition elements have variable oxidation states in their compounds.

For example:

Iron exists most commonly in Fe(II) and Fe (III) oxidation states.

Copper have 3 oxidation states. Cu(I), Cu (II) and Cu (III).

Manganese have 6 different oxidation states, i.e. Mn^{+2}, Mn^{+3}, Mn^{+4}, Mn^{+5}, Mn^{+6} and Mn^{+7}.

7.3 Oxidizing and Reducing Agents

(i) Oxidizing agent

An oxidizing agent is the specie that oxidizes a substance and itself gets reduced.

For example, $KMnO_4$, $K_2Cr_2O_7$, HNO_3 and Cl_2 etc. According to the classical concept, an oxidizing agent may be,

a) The donor of oxygen to a substance.
b) The acceptor of hydrogen from a substance.
c) The acceptor of an electron from a substance.
d) Similarly, the oxidation number of an oxidizing agent is decreased during a redox reaction.

EXAMPLE: 7.4

Consider the reaction,

$$Zn_{(s)} + CuO_{(s)} \longrightarrow ZnO_{(s)} + Cu_{(s)}$$

oxidized

reduced

- Oxidation number of Cu in CuO is +2
- Oxidation number of Cu(s) is = 0
- There is decrease in oxidation number of Cu from +2 to 0. So, CuO is oxidizing agent in the given reaction.

EXAMPLE: 7.5

Consider another example,

$$Sn^{2+} + Cl_2 \longrightarrow Sn^{4+} + 2Cl^-$$

Oxidation number of Cl_2 is zero,

Oxidation number of Cl^- is -1.

There is a decrease in oxidation number of Cl_2, from zero (0) to – 1. So, Cl_2 is an oxidizing agent in the given reaction.

(ii) Reducing Agent

A Reducing agent is the specie that reduces a substance and itself gets oxidized.

For example, H_2S, SO_2, Na, Al, and Mg etc. According to the classical concept, a reducing agent may be,

a) The acceptor of oxygen from a substance.
b) The donor of hydrogen to a substance.
c) The donor of an electron to a substance.
d) Similarly, the oxidation number of a reducing agent is increased during a redox reaction.

Consider the following reaction,

$$\overset{0}{Br_2} + H_2\overset{-2}{S} \longrightarrow 2H\overset{-1}{Br} + \overset{0}{S}$$

oxidized

reduced

Oxidation number of S in H_2S is – 2,

Oxidation number of free S is zero.

There is an increase in oxidation number of S from -2 to zero. So, H_2S is a reducing agent.

Table 7.1 Some examples of oxidising and reducing agents

Oxidising agents	Reducing agents
Bromine (Br_2)	Carbon (C)
Chlorine (Cl_2)	Carbon monoxide (CO)
Concentrated sulfuric acid (H_2SO_4)	Hydrogen (H_2)
Nitric acid (HNO_3)	Hydrogen sulfide (H_2S)
Oxygen (O_2)	Metals
Potassium permanganate ($KMnO_4$)	Potassium iodide (KI)
Potassium dichromate ($K_2Cr_2O_7$)	Sulphur dioxide (SO_2)

Scientific Information

- Ozone is a very strong oxidizing agent and used in manufacturing of decolorizing agents or as oxidants of organic materials.
- When oxygen is used as an oxidizing agent at high temperature and pressure by dissolving it in wastewater, the process is known as wet oxidation. It is considered as an effective method of oxidizing organic materials and removal of toxic compounds.

7.4 Oxidation - Reduction Reactions

Those chemical reactions in which the oxidation-reduction (loss and gain of electrons) takes place simultaneously are called oxidation – reductions or redox reactions.

$$CuSO_{4(aq)} + Zn_{(s)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$$

The overall ionic equation is

$$\overset{+2}{Cu}{}_{(aq)} + \overset{0}{Zn}{}_{(s)} \longrightarrow \overset{+2}{Zn}{}_{(aq)} + \overset{0}{Cu}{}_{(s)}$$

decrease in oxidation state (+2 → 0)

increase in oxidation state (0 → +2)

The redox reactions are made up of two half reactions, for example consider, the reaction

$$2\overset{0}{Na} + \overset{0}{I_2} \longrightarrow 2Na^+ + 2I^-$$

This reaction can be written in the form of two half reactions as,

$$2\overset{0}{Na} \longrightarrow 2Na^+ + 2e^- \quad \text{(Oxidation)}$$

$$\overset{0}{I_2} + 2e^- \longrightarrow 2I^- \quad \text{(Reduction)}$$

In the 1st half reaction, loss of electrons takes place. This reaction is called an oxidation reaction. In the 2nd half reaction, gain of electrons takes place. This reaction is called a reduction reaction.

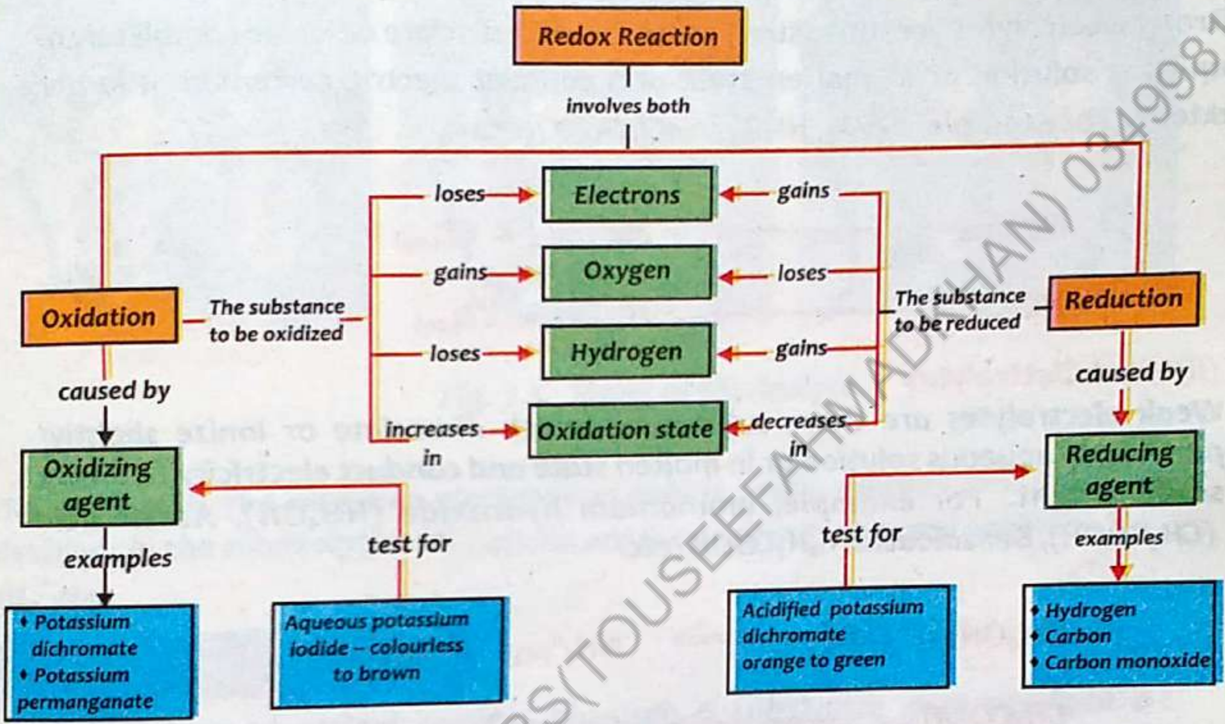


Fig. 7.2 Redox Reaction

7.5 Electrochemical Cell

A device in which interconversion of electrical and chemical energies take place is called electrochemical cell. It is an energy device, in which either a chemical reactions take place by using electric current (such as electrolysis) or chemical reaction produces electric current (such as electric conductance).

Electrochemical cells are of two types.

(i) Electrolytic cell

(ii) Voltaic cells (also called galvanic cell)

7.5.1 Concept of Electrolytes

An electrolyte is a substance in solution or in molten state, which dissociates or ionizes into positive and negative ions and conducts electricity. For example, NaCl, HCl, KCl, NaOH etc. A substance can be strong, weak or non-electrolyte depending upon their degree of dissociation or ionization.

(i) Strong Electrolytes

Strong electrolytes are those substances which dissociate or ionize completely in aqueous solution or in molten state and conduct electric current to a larger extent. For example, H_2SO_4, HNO_3, NaCl, $NaNO_3$ etc

$$NaCl_{(aq)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

(ii) Weak Electrolytes

Weak electrolytes are those substances which dissociate or ionize slightly/partially in aqueous solution or in molten state and conduct electricity to a very small extent. For example, ammonium hydroxide (NH_4OH), Acetic acid (CH_3COOH), Benzoic acid (C_6H_5COOH) etc.

$$NH_4OH_{(aq)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

$$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

(iii) Non – electrolytes

Non – electrolytes are those substances which do not dissociate or ionize in aqueous solution or in molten state, therefore do not conduct electricity. For example, Benzene, Sugar, Urea, Glucose etc.

Electrodes

Electrodes are the conductors i.e. metallic plates, wires or rods, through which electrons enters or leaves the electrolytes in a cell. There are two types of electrodes, anode and cathode.

(a) Anode

The anode is the positive electrode at which the anion gathers and leaves the electron in the electrolytic cell. Anions are the negatively charged particles e.g. Cl^-, OH^- etc.

(a) Noneelectrolyte (b) Weak electrolyte (c) Strong electrolyte

Fig. 7.3 Types of Electrolyte

(b) Cathode

The cathode is the negative electrode at which cation gathers and gains the electrons in the electrolytic cell. Cations are the positively charged particles. Na^+, NH_4^+ etc

Scientific Information

- Passing an electrical current through a substance may produce a temporary change such as lighting up of an electrical bulb or a permanent change like electroplating.
- Metals are good conductors of electricity. Gold and silver are the best electrical conductors but because of their high cost value most of the electrical wires are made of copper and aluminum.

7.5.2 Electrolytic Cells

Some oxidation-reduction reactions do not occur spontaneously but can be driven by electrical energy. The process in which an electric current is used to produce an oxidation-reduction reaction is electrolysis. When electrical energy is required to produce a redox reaction and bring a chemical change in an electrochemical cell, such a cell is called an electrolytic cell. We can also say that it is a device in which non-spontaneous chemical reaction is carried out by passing electric current. Examples of electrolytic cells are Downs cell, Nelson cell etc.

(i) Electrolysis

Electrolysis is defined as the process of chemical decomposition /break down of an electrolyte (ionic compound) in fused /molten state or in solution by passing the electric current.

(ii) Construction of an electrolytic cell

An electrolytic cell consists of a solution of an electrolyte. Two metallic plates called electrodes i.e. anode and cathode are dipped in the electrolytic solution. The electrodes are connected to the terminals of the battery. The electrode which is connected to the positive terminal of the battery is called anode and the electrode which is connected to the negative terminal of the battery is called cathode.

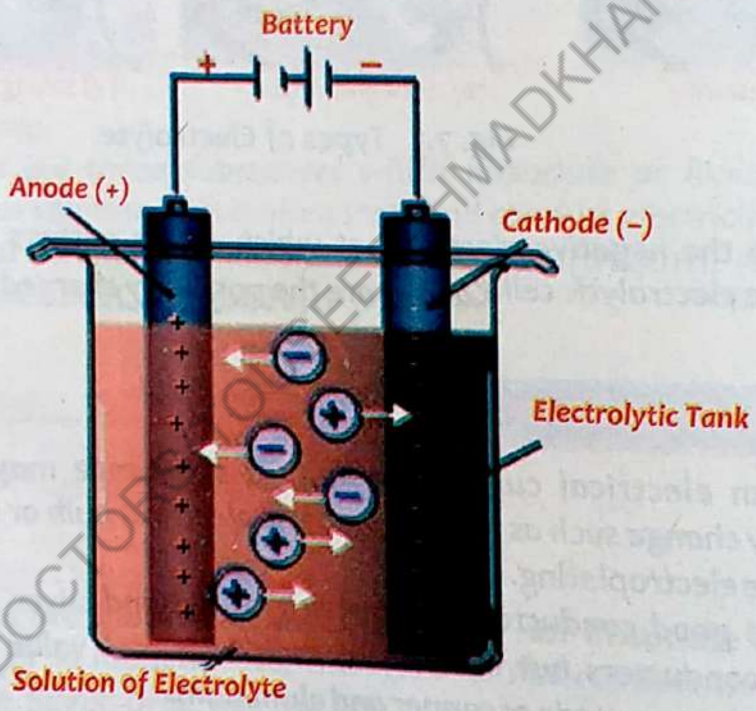


Fig. 7.4 Electrolytic Cell

(iii) Working of an electrolytic cell

When the electrodes are connected to the battery and electric current is passed in the electrolytic cell, the ions in the electrolyte moves towards their respective electrodes. The anions liberate electrons at anode. These electrons pass through outer circuit to the cathode. The cations which surround the cathode, consume those electrons. Hence, the number of electrons lost is always equal to the *number of electrons gained.*

The battery can be thought of as an electron pump, simultaneously supplying electrons to the cathode and receiving electrons from the anode. The anion moves towards anode and discharge their electron(s) there and thus oxidation takes place at anode. The cations move towards cathode and gain the electron(s) there and thus reduction takes place at cathode. For example, when electric current passed from the fused sodium chloride (NaCl), the following reactions take place during the process.

$$2NaCl_{(l)} \longrightarrow 2Na^+_{(l)} + 2Cl^-_{(l)}$$

Oxidation reaction at anode:

$$2Cl^-_{(l)} \longrightarrow Cl_{2(g)} + 2e^-$$

Reduction reaction at cathode:

$$2Na^+_{(l)} + 2e^- \longrightarrow 2Na_{(l)}$$

Overall reaction:

$$2Na^+_{(l)} + 2Cl^-_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2(g)}$$

(iv) Methods of Recovering Metals from Its Ore

Electrolytic Purification of Metals (Refining of Copper)

Pure Copper is very good conductor of electricity and is used in electrical instruments. Copper is purified by electro-refining. Large blocks of blistered (impure) Copper are suspended as anode in the Copper Sulphate ($CuSO_4$) and Sulphuric acid (H_2SO_4) solution between thin sheets of pure Copper which act as cathode. The operation is performed at 50°C and applied voltage of about 0.3 volts and optimum current density used is 160–400 A/m^2.

When the electric current is passed Copper dissolves from the impure Copper anode to give Cu^{+2} ions.

Reaction at anode (oxidation):

$$Cu_{(s)} \longrightarrow Cu^{+2}_{(aq)} + 2e^-$$

At the cathode, all the Cu^{+2} ions from the solution are reduced to metallic copper and get deposited at cathode.

Reaction at cathode (reduction)

$$Cu^{+2}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

As the electrolysis is continued, Copper from the anodes goes into solution. Traces of more active metals like Zn, Fe, etc are also dissolved. The less active metals, for example Au, Ag remain un-dissolved and settled at the bottom of the cell as "Anode Sludge", which is processed to recover these precious metals. The voltage and temperature conditions are such that only Copper is deposited at the cathode. By electrolytic refining up to 99.99% pure copper is obtained.

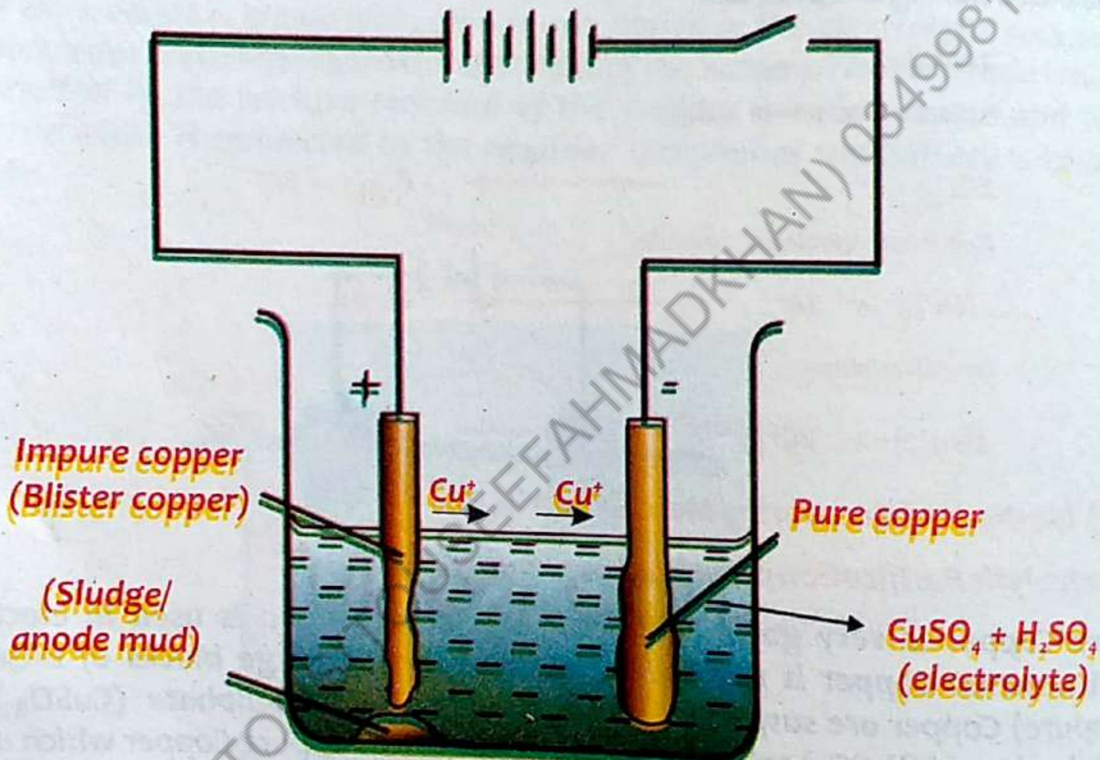


Fig. 7.5: Refining of Copper

7.5.3 Galvanic Cells (Voltaic Cell)

If the redox reaction in an electrochemical cell occurs spontaneously and produces electrical energy, such a cell is called voltaic cell. The device, in which chemical energy (from the reactions between the electrodes and the solution) is converted into electrical energy, is known as the Galvanic cell. The best example of voltaic or galvanic cell is the Daniel cell as shown in fig 7.6.

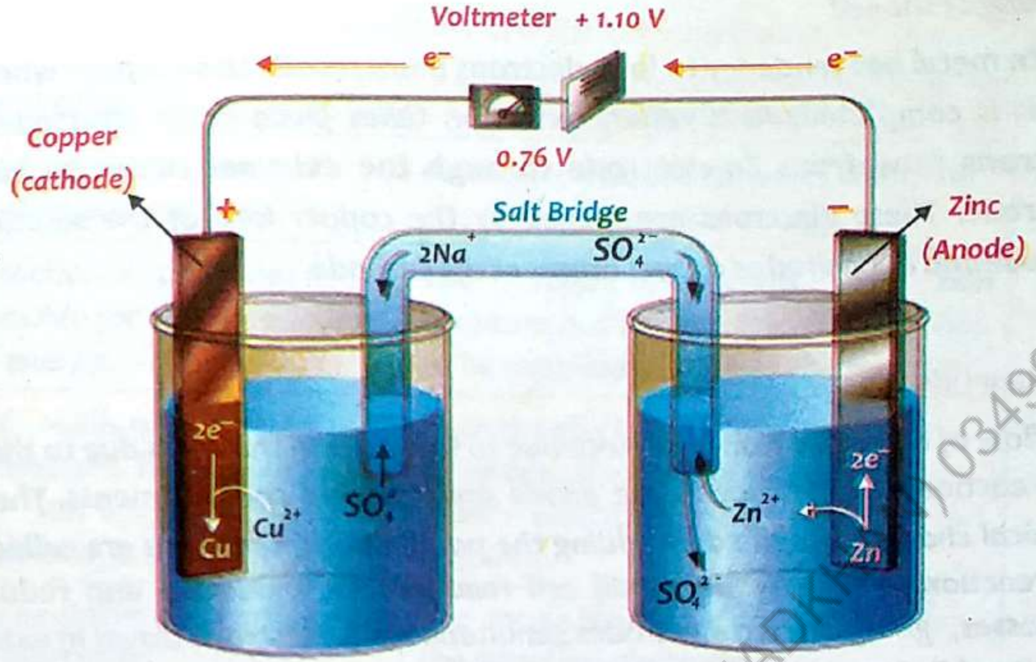


Fig. 7.6: Daniel cell

Construction of a Daniel Cell

A galvanic cell consists of two separate containers, each container is called as half- cell. In each half-cell, an electrode is dipped in 1M solution of its own salt. As shown in figure 7.6. The right half -cell consists of zinc electrode dipped in 1M solution of zinc sulphate($ZnSO_4$). The left half-cell is a copper electrode dipped in 1M solution of copper sulphate ($CuSO_2$) and connected to a wire to an external circuit to which a galvanometer or voltmeter is attached. The solutions in different containers are connected with bridge. This bridge is known as "Salt Bridge". A Salt Bridge is U – shaped tube. This tube is filled with electrolyte gel, such as K_2SO_4 or Na_2SO_4 solution and is called as the "agar". The salt bridge inter-connects the two solutions in the anode container and the cathode container. A Salt Bridge performs three functions,

i. ***It allows electrical contact between the two solutions.***
ii. ***It prevents the mixing of the two solutions.***
iii. ***It keeps electrical neutrality in each half cell.***

Working of the cell

The Zn metal has tendency to lose electrons more readily than copper when the circuit is completed. As a result, oxidation takes place at Zn electrode. The electrons flow from Zn-electrode through the external circuit to copper electrode. These electrons are gained by the copper ions of the solution at cathode and deposited as copper atoms at the cathode.

Cell Reaction

The flow of electrons from one electrode to the other in the cell is due to the half cell reaction-taking place in the anode and cathode compartments. The net chemical changes obtained by adding the two half cell reactions are called the cell reaction. Thus, we have half cell reactions i.e. oxidation and reduction processes, going on two electrodes simultaneously. Electrons travel in external circuit, while ions move through the salt bridge and in this way electric current is produced. These reactions are as follows:

(i) Half –cell reaction at anode (Oxidation)

$$Zn_{(s)} \longrightarrow Zn^{+2}_{(aq)} + 2e^-$$

(ii) Half – cell reaction at cathode (Reduction)

$$Cu^{+2}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

Overall galvanic reaction (the sum of two half –cell reactions)

$$Zn_{(s)} + Cu^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Cu_{(s)}$$

The batteries which are used in automobiles, calculators, toys etc and to light the bulbs work on the same principle.

Table 7.2: Differences between an Electrolytic Cell and Galvanic Cell

Electrochemical cell (Galvanic Cell)	Electrolytic cell
A Galvanic cell converts chemical energy into electrical energy.	An electrolytic cell converts electrical energy into chemical energy.
The redox reaction is spontaneous and is responsible for the production of electrical energy.	The redox reaction is non spontaneous and electrical energy has to be supplied to initiate the reaction.
The two half -cells are set up in different containers, being connected through the salt bridge or porous partition. Salt bridge is required.	Both the electrodes are placed in a same container in the solution of molten electrolyte. Salt bridge is not required.
The anode is negative and cathode is the positive electrode.	The anode is positive and cathode is the negative electrode.
The reaction at the anode is oxidation and that at the cathode is reduction.	The reaction at the anode is oxidation and that at the cathode is reduction.
The electrons are supplied by the species getting oxidized. They move from anode to the cathode in the external circuit.	The external battery supplies the electrons. They enter through the cathode and come out through the anode.

Battery

A group or combination of galvanic cells joined in series is called battery. Car batteries consist of six or more identical voltaic cells connected in series. A battery is a self-contained, chemical power pack that can produce a limited amount of electrical energy, wherever it is needed. A battery converts chemicals energy into electrical energy, for specific period of time.

Fig.7.7 : Many common batteries are simple voltaic dry cells

Dry cell:

The dry cell was prepared by LECLANCHE in 1887.

Construction of the Cell:

The dry cell consists of metallic container. Its container is made up of Zinc (Zn). This container acts as anode. This Zinc casing is consumed during the chemical reaction. A graphite rod is placed in the center of the container. This graphite rod acts as cathode. This container is filled with the mixture of Ammonium chloride (NH_4Cl), Manganese Dioxide (MnO_2) and Carbon (C). This mixture is in the paste form. The cell is water proofed with the wax. The voltage produced by the dry cell is 1.25V.

Reaction in the Cell

Oxidation and reduction reactions occur in the cell to produce the electric current.

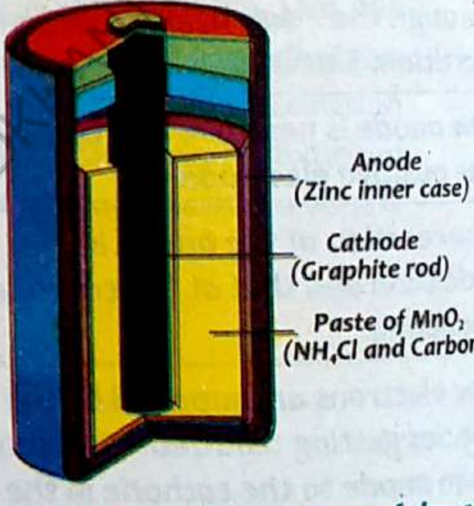


Fig. 7.8 Internal structure of dry cell

Reaction at Anode

The Zinc acts as Anode in the cell. The Zinc is oxidized by losing the two electrons.

$$Zn \longrightarrow Zn^{+2} + 2e^-$$

Reaction at Cathode

The graphite acts as Cathode in the cell. In the cell the NH_4Cl and MnO_2 are reduced to Mn_2O_3 and NH_4.

$$2MnO_2 + 2NH_4^+ + 2e^- \longrightarrow Mn_2O_3 + 2NH_3 + H_2O$$

$$2NH_4Cl \longrightarrow 2NH_4^+ + 2Cl^-$$

Overall reaction

$$Zn^{+2} + 2Cl^- \longrightarrow ZnCl_2$$

7.6 Electrochemical Industries

Electrochemical industries are based on many electrochemical operations. Some of these are as follows.

i. *Electrochemical cells or batteries constructed with different electrodes are available in the market, which are widely used to power toys, flashlights, electronic calculators, pacemakers, radios, tape-recorders, automobiles, etc.*

ii. *Electroplating of metals is the deposition of one metal on other metal electrolytically. This is done for the purpose of its durability, beauty or repair.*

iii. *Electrolytic production of metals (e.g. Na) and electrolytic refining of metals (e.g. Cu) are the popular methods for obtaining metals in their pure form.*

iv. *Many important chemicals are manufactured by electrochemical process, e.g. NaOH.*

7.6.1 Manufacture of Sodium Metal from Fused NaCl

Sodium metal was first discovered by an English Chemist, Sir Humphrey Davey in 1807, by the electrolysis of fused Sodium Hydroxide (NaOH). Now-a-day, the electrolysis of fused compounds is used for the production of most of the metals.

Commercially Sodium metal is obtained from the electrolysis of molten Sodium Chloride (NaCl) in the Down Cell.

Construction of the Down Cell

The down cell consists of steel container lined inside with firebricks. The anode is made of Graphite at the center which emerges from the bottom, above which there is dome for the collection of Chlorine gas. The cathode is circular and made of Copper or Iron. The cathode and anode are separated by an iron screen, so that the two products of electrolysis, namely Sodium and Chlorine gas are kept apart and are collected separately. The molten Sodium collects in the cathode compartment where it rises to the top and is taped off via a pipe. On the other hand, gaseous Chlorine, as by-product is collected, at anode and is collected from the other pipe.

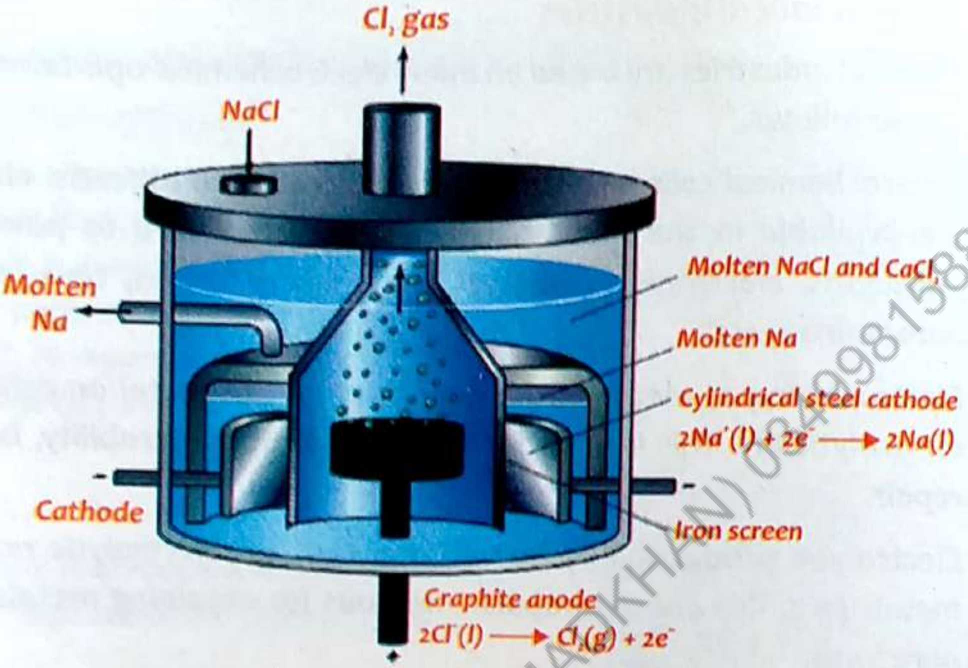


Fig.7.9 Down Cell

Reaction in the Cell

The melting point of Sodium Chloride (NaCl) is 801°C. A mixture of one mole of NaCl and three moles of $CaCl_2$ is added, which reduce the melting point of the Sodium Chloride (NaCl) to 600°C. When the electric current is passed ions are produced and the ions are free to move towards the oppositely charged electrodes. The positively charged Na^+ ions move towards the cathode, picks electron and are deposited as Sodium metal.

Molten NaCl ionizes as

Fused Sodium Chloride contains Sodium and Chloride ions.

$$2NaCl_{(s)} \longrightarrow 2Na^+_{(l)} + 2Cl^-_{(l)}$$

(i) Half cell reaction at Cathode (Reduction)

The Sodium ions receive one electron each, to become Sodium metal atoms. Hence the sodium ion (Na^+) is reduced at the cathode.

$$2Na^+_{(l)} + 2e^- \longrightarrow 2Na_{(l)}$$

(ii) Half cell reaction at Anode (Oxidation)

The Chloride ions (Cl^-) moves towards anode and give its electrons to the electrode and change to neutral chlorine atom. Since chlorine atom cannot exist independently in the atomic state, they combine to form chlorine molecules (Cl_2). Hence, at anode chlorine gas (Cl_2) is liberated.

$$2Cl^- \longrightarrow 2Cl_{(g)} + 2e^-$$

$$Cl_{(g)} + Cl_{(g)} \longrightarrow Cl_{2(g)}$$

Overall reaction is the sum of these two half-cells reactions

$$2NaCl_{(Fused)} \longrightarrow 2Na_{(l)} + Cl_{2(g)}$$

Society, Technology and Science

Aluminum is the most abundant metal in the earth's crust (about 8.2% of the total metals). It is never found in free state in nature. Aluminum is reactive and will react spontaneously with water and air to form aluminum oxide. Aluminum oxide, (Al_2O_3) forms a stable passive layer. This layer is about 4 nm thick and will provide corrosion protection to aluminum as long as this oxide layer is present.

Aluminum is an amphoteric metal and can react with an acid as well as a base. The protective layer of aluminum oxide will deteriorate in environments with high or low pH. The oxide layer is only stable in a pH range of 7.0 to 9.0.

Fe_2O_3 is the chemical formula of Ferric Oxide. Ferric Oxide is one of the three main oxides of iron. The other two types are Ferous Oxide (FeO) and iron Oxide (Fe_3O_4), which is also called Magnetite. When iron comes in contact with oxygen and water, it turns to Iron Oxide (rust). When iron corrodes the color changes and it actually expands. This expanding and color change can produce large red flakes that we all know as rust. Rusting is formed more quickly in the case of saltwater because it contains sodium which is more reactive to iron. This might be one of the major problems which a ship crew has to deal with.

7.6.2 Manufacture of NaOH from Brine

A concentrated aqueous solution of Sodium chloride, NaCl (Brine) is placed in a special apparatus, known as Nelson cell for the manufacture of NaOH.

Nelson cell

It consists of U-shaped tube. This tube is made of steel. It is perforated. This perforated tube acts as cathode. A graphite anode is suspended in the U-shaped tube. The cathode is coated with asbestos. The asbestos separates the anode from cathode.

During the electrolysis, the Chlorine is produced at the anode. It is collected at the Chlorine outlet. Hydrogen gas is produced at the cathode. It is collected at the Hydrogen outlet. During this reaction Sodium hydroxide is also produced. The Sodium hydroxide is collected in the catch basin, placed under the U-shaped tube. In this process, the Hydrogen, Chlorine and Sodium hydroxide is produced at the same time.

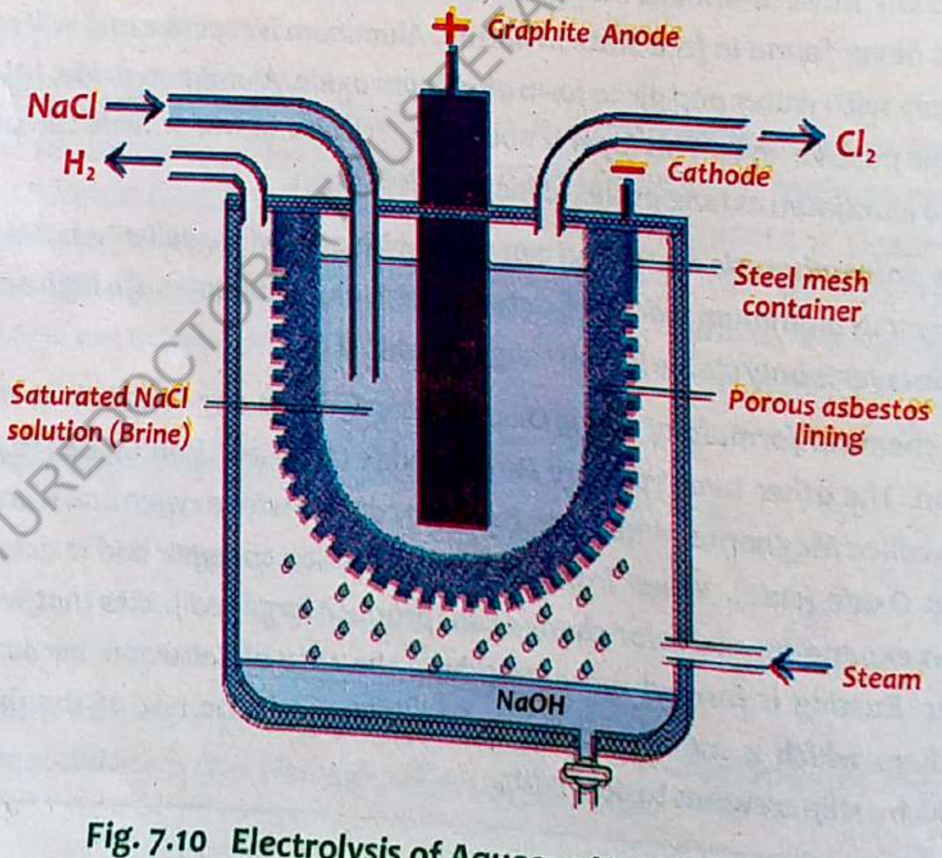


Fig. 7.10 Electrolysis of Aqueous NaCl (Nelson's Cell)

The saturated brine solution ionized as follows:

$$2NaCl_{(aq)} \longrightarrow 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)}$$

$$2H_2O_{(l)} \rightleftharpoons H^{+}_{(aq)} + OH^{-}_{(aq)}$$

When the electrodes are connected to the battery, the positive ions, Na^+ and H^+ move towards the cathode. Since, H^+ have a greater tendency to pickup electrons from the cathode as compared to Na^+, therefore H^+ ions picks up electrons to form H_2 gas. Na^+ ions are not reduced instead they combine with OH^- ions, present in the solution to form caustic soda (NaOH) which make the solution alkaline, while Cl^- ions move towards the anode where they give electrons to the electrode.

(i) Reaction at cathode:

$$2H^{+}_{(aq)} + 2e^{-} \longrightarrow 2H_{(aq)} \quad \text{(Reduction)}$$

$$H + H \longrightarrow H_2$$

$$2Na^{+}_{(aq)} + 2OH^{-}_{(aq)} \longrightarrow 2NaOH_{(aq)}$$

(ii) Reaction at anode:

$$2Cl^{-}_{(aq)} \longrightarrow 2Cl_{(g)} + 2e^{-} \quad \text{(Oxidation)}$$

$$Cl_{(g)} + Cl_{(g)} \longrightarrow Cl_{2(g)}$$

The overall reaction:

$$2NaCl_{(aq)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)} + Cl_{2(g)}$$

Interesting Fact

To separate the chlorine from the sodium hydroxide, the two half-cells were traditionally separated by a porous asbestos diaphragm, which needed to be replaced every two months.

This was damaging to the environment, as large quantities of asbestos had to be disposed. Asbestos is toxic to humans and causes cancer and lung problems. Today, the asbestos is being replaced by other polymers, which do not need to be replaced as often and are not toxic.

7.7 Corrosion and its Prevention

The meaning of corrosion to a number of people is rust. The word "Rust" is more specifically used for iron, whereas corrosion is commonly defined as the deterioration of a substance (usually a metal) or its properties because of a reaction with its environment. The terms corrosion and rust are almost synonymous. **Corrosion is slow and continuous eating away of a metal by the environment or surrounding.** Corrosion usually starts at the exposed surface of the metals. Corrosion is a naturally occurring spontaneous phenomenon and it drives the materials to its lowest possible energy states. **It is an oxidation - reduction process which takes place by the action of air in the presence of moisture with the metals.** The most common example of corrosion is the rusting of iron.

Society, Technology and Science

Photographic processing or development is a chemical process by which photographic film is treated, after exposure to light, to produce an image. All photographic processes use a series of chemical reactions. Processes such as the development stages, requires very close control of light, temperature, shaking and time.

The exposure of light on photographic plate initiated the chemical reaction. The light exposed portion become dark, depending upon the amount or time of exposure. That exposed plate is later on developed to show the image. The film is soaked in water to swell the gelatin layer on it and making it easy for the action of chemical treatments. The developer (silver halide) converts the latent image to macroscopic particles of metallic silver. A bath contains a dilute solution of acetic acid or citric acid; stop the action of the developer. A rinse with clean water may be substituted. The fixer makes the image permanent and light-resistant by dissolving remaining silver halide. A common fixer is hypo, specifically ammonium thiosulphate. Washing in clean water removes remaining fixer. Remaining fixer can corrode the silver image, leading to discolouration, staining and fading. Although, technologically more advanced but the basic procedures which were developed originally are still used in all silver- based photography today. In coloured films, there are three layers of the silver halides which are specifically formulated to be sensitive to red, green and blue light.

7.7.1 Rusting of Iron

The corrosion of iron is commonly known as rusting. The necessary condition for rusting is moisture and air. There will be no rusting in water vaopurs free of air or air free from water vapours. Iron rusts by combining with oxygen in the presence of water to form brown hydrated mass, ferric oxide ($Fe_2O_3.xH_2O$).

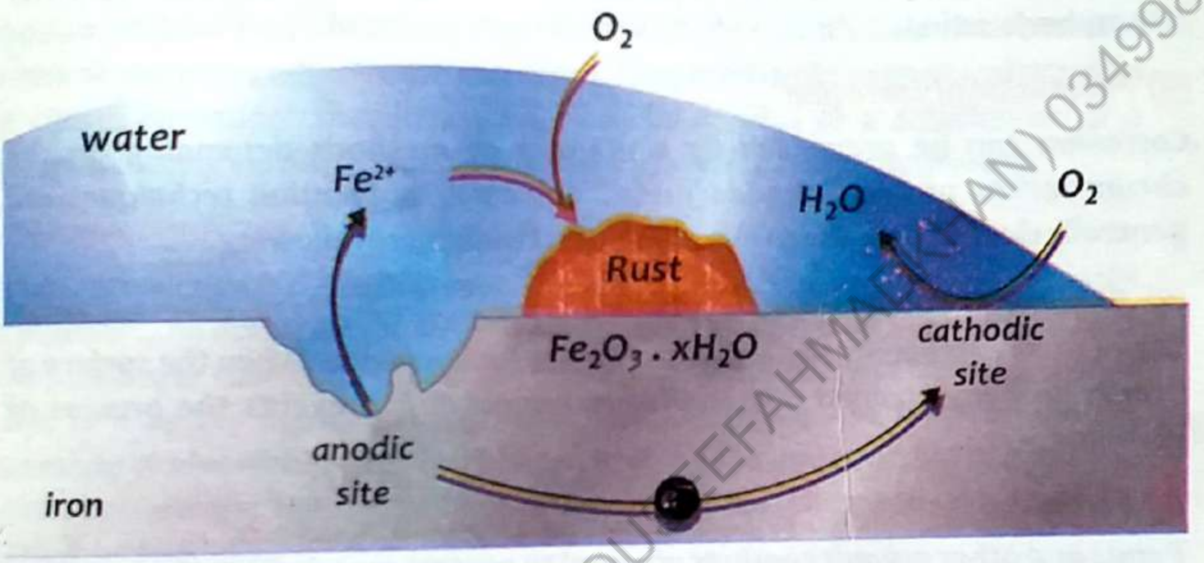


Fig. 7.11 Rusting of Iron

(i) The chemistry of rusting

Dents and stains are the site for the process of rusting. These sites are called anodic region and following oxidation reaction takes place on these sites.

At anode

$$2Fe_{(s)} \longrightarrow 2Fe^{+2}_{(aq)} + 4e^-$$

This loss of electron damage the metals. These free electrons move in the metal sheet, to a region of high concentration of oxygen (O_2) and water near the surface. These sites acts as cathode where electrons react with water and oxygen to form hydroxide ion.

At cathode

$$2H_2O + O_2 + 4e^- \longrightarrow 4OH^-$$

Fe^{+2} is furthers oxidized by atmospheric oxygen to form hydrated Fe^{+3} oxide (rust)

$$Fe^{+2}_{(aq)} \longrightarrow Fe^{+3}_{(aq)} + e^-$$

$$Fe^{+3}_{(aq)} + 3OH^-_{(aq)} \longrightarrow \underset{\text{Rust}}{Fe(OH)_{3(s)}}$$

The rust mass is soft and porous in nature and therefore cannot prevent further atmospheric action.

(ii) Prevention of Corrosion

Corrosion can be prevented by a number of methods depending on the circumstances of the corroded metal. Corrosion prevention techniques are generally classified into six groups. These techniques are following.

(a) Removal of Stains

The regions of stains in an iron act as the site for corrosion. When the surface of iron is properly cleaned and stains are removed, it prevents the process of rusting.

(b) Paints and Coatings

Paints and other organic coatings are used to protect metals from the corrosion effects. Beside this modern paints contain a combination of chemical called stabilizers. These stabilizers provide prevention against not only corrosion but also against weathering and other atmospheric effects.

(c) Alloying

Alloying also helps to protect the corrosion of metals. The best example of alloying is the stainless steel, which is a solid mixture of iron, chromium and nickel. Stainless steel strongly resists the corrosion. The development of new alloys are constantly under production.

Scientific Information

An alloy is a homogenous mixture of metals or a mixture of metals and another element (non-metals).

(d) Metallic Coating or Plating

Metallic coatings or plating, can be applied to inhibit corrosion as well as provide aesthetic and decorative finish. A thin coating of one metal on another can be applied by spraying, galvanizing(deposition of Zinc (Zn) on the metal by dipping) or electroplating, for example Iron articles are protected from rusting by Nickel (Ni), Chromium (Cr) or tin (Sn) plating.

(e) Corrosion Inhibitors

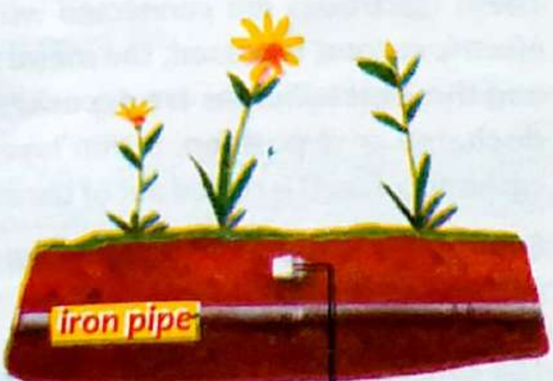
Corrosion inhibitors are chemicals that react with the metal's surface or with the environmental gases which cause corrosion. They interrupt the chemical reaction that causes corrosion. These chemicals can be applied as a solution or as a protective coating via dispersion techniques e.g. glycine, polyethylene etc.

(f) Cathodic Protection

Cathodic protection is a method usually used to protect iron in buried fuel tanks and pipelines. An active metal, such as magnesium or zinc, is connected by a wire to the pipeline or tank to be protected (Figure 7.12). It is because the magnesium or zinc is a better reducing agent than iron, electrons are supplied by the magnesium or zinc rather than by the iron, keeping the iron from being oxidized. As oxidation occurs, the magnesium or zinc anode dissolves and so it must be replaced periodically.

Ships' hulls are protected in a similar way by attaching bars of titanium metal to the steel hull (Figure 7.12). In salt water the titanium acts as the anode and is oxidized instead of the steel hull (the cathode).

Bars of titanium can be fixed to a ship's hull to prevent the ship's steel body from rusting.



Heavy blocks of magnesium attached to underground pipes made of iron to protect the pipes from rusting.

Fig. 7.12 Cathodic Protection

7.7.2 Electroplating of Tin, Zinc and Chromium on Steel

An electrolytic process in which a thin layer of one metal is deposited on another metal surface. Metal ion is reduced by passing electricity and a solid metal is deposited on a surface is called electroplating. In an electrolytic cell, the metal which is to be deposited is made anode and cathode is made of metal on which deposition takes place. The electrolyte is an aqueous solution of a salt of the respective metal which is to be deposited on the metal.

The purpose of electroplating are following.

i. **Protection:** To protect the inner metal from the atmospheric effect, that is from corrosion. From example, Ni and Cr are deposited over iron to prevent it from corrosion.

ii. **Repair:** To weld the broken parts of the machinery by depositing the metal on it.

iii. **Decoration:** To deposit the noble metals like gold and silver on an inferior metals to enhance its beauty.

Procedure of electroplating

In electroplating, the metallic substance to be electroplated is cleaned with sand, washed with caustic soda (NaOH) and at last, thoroughly washed with water. This cleaned substance is made cathode. A sheet or rod of pure metal to be deposited is made anode. Salt of that metal, which is to be deposited, is taken as an electrolyte. This process is carried out in an electrolytic cell. The electrolytic cell is made up of glass, cement or wood, in which the electrodes are immersed. These electrodes are connected with the terminals of the battery. When the electric current is passed, the metal from anode converted into ions in solution and these metallic ions are deposited on the cathode (object). As a result of this discharge or deposition, a thin layer of metal is deposited on the object. The cathode (object) is pulled out of the electrolytic cell, cleaned and dried.

Some examples of electroplating are given below in detail.

(i) Electroplating of Tin

The target metal is cleaned with caustic soda, treated with acids, in order to remove the rust and oils/greases if any present on it. Then it is washed thoroughly with water. The Electroplating of Tin is carried out in electrolytic cell. In this process, pure piece of Tin act as anode and is dipped in sodium stannate

($Na_2SnO_3 \cdot 3H_2O$) used as electrolytic solution. The cathode is the object to be coated with Tin. When the electric current is passed through the cell, the anode starts dissolving and converted into Sn^{+2} ions. These Sn^{+2} ions move towards the cathode. At cathode they discharged and deposited on the object. The chemical reaction can be represented as,

At anode:

$$Sn_{(s)} \longrightarrow Sn^{+2}_{(aq)} + 2e^-$$

At cathode:

$$Sn^{+2}_{(aq)} + 2e^- \longrightarrow Sn_{(s)}$$

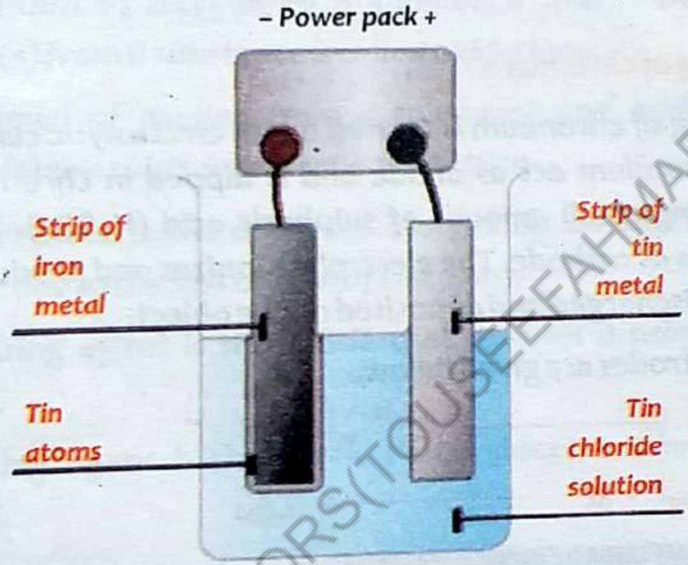


Fig. 7.13 Electroplating of Tin

Society, Technology and Science

Pure silver is very soft metal, also called fine silver. It is relatively soft, very malleable. Silver atoms have weak interactions and are loosely packed together. Silver tarnishes in air when it comes in contact with trace amount of H_2S or SO_2 in the air and turn blackish. Due to this reason decorative and practical objects which are made of solid silver gradually turn black and lose their shining appearance. It is easily damaged so it is commonly combined with other metals to produce a more durable product. The most popular of these alloys is sterling silver, which consists of 92.5% silver and 7.5 % copper.

(ii) Electroplating of zinc

The target metal is cleaned and washed thoroughly with water. This object is dipped in zinc sulphate ($ZnSO_4$) containing a small amount of sulphuric acid (H_2SO_4) solutions which act as an electrolyte. The object to be electroplated acts as cathode, while anode is made of zinc plate or rod. When the electric current is passed, zinc anode dissolve and converted into Zn^{+2} ions. The electrons move to the cathode, where Zn^{+2} ions discharged and deposited as Zn metal. The chemical reaction can be represented as,

At anode: $Zn_{(s)} \longrightarrow Zn^{+2}_{(aq)} + 2e^-$

At cathode: $Zn^{+2}_{(aq)} + 2e^- \longrightarrow Zn_{(s)}$

(iii) Electroplating of chromium

The Electroplating of chromium is carried out in electrolytic cell. In this process, pure piece of chromium act as anode and is dipped in chromic acid (H_2CrO_4) solution containing small amount of sulphuric acid (H_2SO_4). The object to be electroplated acts as cathode. The electrolyte ionizes and produces the Cr^{+3} ions, at cathode they discharged and deposited on the object.

Reactions on electrodes are given below,

At anode: $Cr_{(s)} \longrightarrow Cr^{3+}_{(aq)} + 3e^-$

At cathode: $Cr^{3+}_{(aq)} + 3e^- \longrightarrow Cr_{(s)}$

Society, Technology and Science

To preserve the food and taste of the food, preservatives are added. These preservatives are usually organic acids such as acetic acid and benzoic acids etc, or their salts. If the food and beverages are caned in iron or in any other corrosive (reactive) metal, these acids will react with cans. It will corrode the cans and will destroy the food and beverages and may form toxic substances by reacting with iron or other reactive metal.

To deal with these problems, food and beverages industries use tin plated steel cans. Tin plating is nonpoisonous and prevents corrosion. It also preserves the taste of food and beverages by preventing the reaction with cans.

Key Points

- *Electrochemistry is the branch of chemistry which deals with the interconversion of electrical energy and chemical energy.*
- *The addition of oxygen to a substance and removal of hydrogen or electron(s) from a substance is called oxidation.*
- *The removal of oxygen from a substance and addition of hydrogen or electron(s) to a substance is called reduction.*
- *Oxidation state is the apparent charges on an atom whether positive or negative due to the loss or gain of electrons.*
- *An oxidizing agent is the specie that oxidizes a substance and itself gets reduced.*
- *A Reducing agent is the specie that reduces a substance and itself gets oxidized.*
- *Those chemical reactions in which the oxidation-reduction takes place simultaneously are called oxidation – reduction or redox reactions*
- *A device in which interconversion of electrical and chemical energies take place is called electrochemical cell.*
- *An electrolyte is a substance in solution or in molten state, which dissociates or ionizes into positive and negative ions and conducts electricity.*
- *Non – electrolytes are those substances which do not dissociate or ionize in aqueous solution or in molten state, therefore, do not conduct electricity.*
- *The process of chemical decomposition of an electrolyte in fused (molten) state or in solution by passing the electric current is called electrolysis*

- Electrodes are conductors i.e. plates, wires or rods, through which an electrons enters or leaves the electrolytes in a cell. The positive electrodes are called anode and the negative electrodes are called cathode.
- If the oxidation and reduction reaction occurs spontaneously and produces electrical energy in an electrochemical cell, the cell is called voltaic cell.
- Down Cell is used for extraction of Sodium metal commercially, by electrolysis of molten Sodium Chloride (NaCl)
- Nelson cell is used for preparation of sodium hydroxide (caustic soda), by electrolysis of concentrated aqueous solution of Sodium chloride, NaCl (Brine).
- Electroplating is depositing of one metal over the other by means of electrolysis.
- Corrosion is slow and continuous eating away of a metal by the environment or surrounding.

Exercise

Choose the correct option.

i. The oxidation number of nitrogen N_2 is:

a. +1 b. +3

c. 0 d. −3

ii. What is the oxidation number of C in CH_3OH:

a. −2 b. −1

c. 0 d. +1

iii. A cation is :

a. neutral b. negatively charged

c. positively charged d. no charges

iv. Electrons are lost by the:

a. reducing agent as it undergoes oxidation

b. reducing agent as it undergoes reduction

c. oxidizing agent as it undergoes oxidation

d. oxidizing agent as it undergoes reduction

v. What is the oxidation number assigned to manganese in $KMnO_4$:

a. +7 b. +3

c. +2 d. +4

vi. In a particular redox reaction, the oxidation number of phosphorus changed from -3 to 0. From this information, it may be concluded that phosphorus:

a. Lost 3 electrons and was reduced.

b. Lost 3 electrons and was oxidized.

c. Gained 3 electrons and was reduced.

d. Gained 3 electrons and was oxidized.

vii. Which statement is true for an electrochemical cell:

a. Oxidation occurs at the anode only.

b. Reduction occurs at the anode only.

c. Oxidation occurs at both the anode and cathode.

d. Reduction occurs at both the anode and cathode.

viii. In which of the following does sulphur have an oxidation number of +7:

a. HSO_3^- b. SO_3

c. H_2SO_4 d. $H_2S_2O_8$

ix. What happens to the reducing agent in an oxidation-reduction reaction:

a. It is oxidized as it gains electrons.

b. It is oxidized as it loses electrons

c. It is reduced as it gains electrons.

d. It is reduced as it loses electrons.

x. In an electrochemical cell, electrons travel in which direction:

a. From the anode to the cathode through the external circuit

b. From the anode to the cathode through the salt bridge

c. From the cathode to the anode through the external circuit

d. From the cathode to the anode through the salt bridge

SHORT QUESTIONS

Answer briefly the following questions.

(I) Indicate which element is reduced in the following reactions:

a. $Ca_{(s)} + Br_{2(g)} \longrightarrow CaBr_{2(s)}$

b. $\overset{+}{H}Cl^{-} + \overset{0}{Zn} \longrightarrow \overset{+2}{Zn}Cl^{-2} + \overset{0}{H}_2$

(ii) What is the oxidation number of silver on each side of the following equation?

$$4\,Ag(s) + O_2(g) \xrightarrow{\text{Heat}} 2\,Ag_2O(s)$$

(iii) Why NaOH is a strong but NH_4OH is weak electrolyte?

(iv) How to prevent corrosion? Enlist few of the methods.

(v) Write chemical reactions that occur in Nelson's cell.

(vi) Write one example from daily life which involves the oxidation – reduction reaction.

(vii) Assign oxidation numbers to each atom in the following compounds.

a. HI b. PBr_3 c. $CaCO_3$ d. H_3PO_4

e. As_2O_3 f. H_2SO_4

(viii) Why O_2 is necessary for rusting?

(ix) Sketch the Daniel cell, labeling the cathode, anode and the direction of flow of the electrons.

(x) Write down some possible uses of an electrolytic cell.

LONG QUESTION

I. a. What is electroplating?

b. Distinguish between the nature of the anode and cathode in such a process.

II. Differentiate between the processes of oxidation and reduction. Write an equation to illustrate each.

III. a. What is corrosion? Explain rusting of iron as an example of corrosion.

b. Differentiate between electrolytic cell and voltaic cell.

c. Discuss the method of recovering/extracting of metal form its ore.

V. Discuss the preparation of Sodium Hydroxide (NaOH) from Brine along with diagram and reactions at cathode and anode.

Project Work

Investigate the types of batteries being considered for electric cars. Write a report on the advantages and disadvantages of these types of batteries.

Unit 8 Chemical Reactivity

After studying this unit, the students will be able to;

- Show how cations and anions are related to the terms metals and non-metals.
- Explain why alkali metals are not found in the Free State in nature.
- Identify elements as an alkali metal or an alkaline earth metal.
- Explain the differences in ionization energies of alkali and alkaline earth metals.
- Describe the position of sodium in Periodic Table, its simple properties and uses.
- Describe the position of calcium and magnesium in Periodic Table, their simple properties and uses.
- Differentiate between soft and hard metals (Iron and Sodium)
- Describe the inertness of noble metals.
- Identify the commercial value of Silver, Gold and Platinum.
- Compile some important reactions of halogens.
- Name some elements, which are found in uncombined state in nature.

Introduction

The science of chemistry revolves around the chemical reactions. Chemical are considered the soul of chemistry. When sodium metal is placed in air it catches fire, when iron is placed in moist air, it rusts but when metals like gold and platinum are placed in air they neither catches fire nor rusts. Why ? It is common observation that some metals are found free in nature and some are found in the combined state in the form of compounds, can you give the reason?

In previous grades, you have learnt about the fundamentals of metals and non-metals. In this unit, you will learn in detail about the metals such as alkali and alkaline earth metals and non-metals such as halogen e.g. Cl_2, Br_2 etc.

8.1 Metals

Metals are those substances which are good conductor of heat and electricity. They are malleable and ductile. Their oxides and hydroxides are bases. All of the metals, except Mercury (Hg) exist in solid state at room temperature.

A metal is one which react with oxygen and produce a basic oxide. When it is dissolve in water forms an alkaline solution which turns red litmus paper blue.

Scientists redefined the metals on the basis of loss and gain of electron. A metal is an element which loses an electron and forms a cation. For example, sodium metal loses an electron and form sodium cation (Na^+), etc. Metals are electropositive and so they have the tendency to lose electrons. The only exception to this definition is the non-metal, Hydrogen which is usually an electron donor.

Scientific Tidbit

In the periodic table,
Group 1 = All are metals except Hydrogen
Group 2 = All are metals
Group 3 = All are metals except Boron
Group 4 = Tin and Lead are metals
Group 5 = Antimony and Bismuth are metals
Group 6 = All are Non-metals
Group 7 = All are Non-Metals
Group 8 = All are non-Metals
d- Block Elements = All are metals
Lanthanides = All are metals
Actinides = All are metals

Scientific Information

The earth is said to be made up of about 8% aluminium, 5% iron, 4% calcium and less than 4% of potassium, sodium and magnesium.

Characteristics of Metals and Non-Metals

8.1.1 Electropositive or Metallic Character

The tendency of an element to lose electrons and forms positive ions (cations) is called electropositive or metallic character. A more electropositive element has more metallic character.

$$Na \longrightarrow Na^{+} + e^{-}$$

$$Ca \longrightarrow Ca^{+2} + 2e^{-}$$

$$Al \longrightarrow Al^{+3} + 3e^{-}$$

The elements having lower ionization energies have higher tendency to lose electrons, thus they are electropositive or metallic in their behaviour. The elements having low values of ionization energies are metals. Alkali metals are the most highly electropositive elements, whereas elements having high values of ionization energies are non-metals. The border line elements behave as metalloids.

Non-metals have the ability to accept electron(s) in their valence shell to form negatively charged particles called anions.

$$F + e^{-} \longrightarrow F^{-}$$

$$Cl + e^{-} \longrightarrow Cl^{-}$$

$$O + 2e^{-} \longrightarrow O^{2-}$$

(i) Variation of Metallic Character across a Period

Metallic character decreases across a period from left to right. On the other hand, non-metallic character increases with increase in atomic number across a period. For example, let us consider elements of second and third periods.

	Metals		Metalloid	Non-metals				
Period 2	Li	Be	B	C	N	O	F	Ne
Period 3	Na	Mg	Al	Si	P	S	Cl	Ar
	Metals			Metalloid	Non-metals			

In the second period, lithium and beryllium are metals; boron is a metalloid while carbon, nitrogen, oxygen, fluorine and neon are non-metals.

In the third period, sodium, magnesium and aluminum are metals; silicon is a metalloid while phosphorus, sulphur chlorine and argon are non-metals.

(ii) Variation of Metallic Character along a Group

On going along a group from top to bottom, the metallic character of elements increases. In each group, the first element is least metallic while the last element is most metallic. For example, let us consider the elements of groups 4 and 5.

In group 4, the first element carbon is a non-metal; silicon and germanium are metalloids while tin and lead are metals.
In group5, the first two elements, nitrogen and phosphorus, are non-metals while arsenic and antimony are metalloids and bismuth is a metal.

	Group IV	Group V	
Non-metals	C Carbon	N Nitrogen	Non-metals
Metalloid	Si Silicon	P Phosphorus	Non-metals
Metalloid	Ge Germanium	As Arsenic	Metalloid
Metals	Sn Tin	Sb Antimony	Metalloid
Metals	Pb Lead	Bi Bismuth	Metals

Scientific Information

Conductance:

Metals are good conductor of electricity while non-metals are insulator except carbon in the form of graphite. In order to decide whether an element is metallic or non-metallic, the electrolysis test is conducted. This test consists of dissolving the element in an acid and running an electric current through the solution. If the element is metallic, its atoms will become positively charged and they will move towards the negative pole or cathode of the electrolytic cell. For example,

$$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2}(g)$$

$$ZnCl_{2(aq)} \longrightarrow Zn^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)}$$

8.1.2 Alkali Metals

The elements of the IA group except Hydrogen are called Alkali metals. The name alkali came from Arabic. It means "the Ashes". These metals were first found in the ashes of plants.

Some chemist has the opinion that the word alkali is given due to the fact that these elements react with water and form the strong Alkalies. Alkali metals include the elements Lithium (Li), Sodium (Na), Potassium (K), Rubidium (Rb), Cesium (Cs) and Francium (Fr).

These metals have only one electron in their valence shell. Their valence sub-shell is "s". They are highly electropositive elements. The alkali metals lose their one electron and form mono-positive ions. The ionization energy of alkali metals is low. The electron thus removed is provided to an electronegative element, to form ionic compounds. Elements of group I form ionic compounds with elements of group VI and group VII.

Fig.8.1 Alkali Metals

Occurrence of Alkali Metals

Alkali metals have low ionization energies. They are very reactive metals in nature, that is why they do not occur in the Free State. Lithium found in the form of complex minerals. It mostly occurs in the form of Spodumene, $LiAl(SiO_3)_2$. Sodium and Potassium are abundantly (2.4%) found on the earth crust. Rubidium and Cesium occurs in small amounts in the Potassium salts deposits. Francium is not found in nature. It is prepared in the laboratory.

Table 8.1 Reactivity of Alkali Metals

Order of reactivity: lithium <sodium< potassium

Reactions of alkali metals			
Alkali	**Reaction with air (Oxygen)**	**Reaction with water**	**Reaction with chlorine**
Lithium (Li) At. No = 3	Burns with red flame to give lithium oxide, which is white solid $4Li_{(s)} + O_{2(g)} \longrightarrow 2Li_2O_{(s)}$	Floats on water and reacts quickly to produce lithium hydroxide and hydrogen gas. $2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$	Burns with a bright flame to give a white solid of lithium chloride. $2Li_{(s)} + Cl_{2(g)} \rightarrow 2LiCl_{(s)}$
Sodium (Na) At. No = 11	Burns with a bright yellow flame to produce white sodium oxide $4Na_{(s)} + O_2(g) \rightarrow 2Na_2O_{(s)}$	Floats on water and reacts very quickly to produce sodium hydroxide and hydrogen gas. $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$	Burns with a bright flame to give a white solid of sodium chloride. $2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$
Potassium (K) At. No = 19	Burns violently with a little colored flame to produce white potassium oxide. $4K_{(s)} + O_{2(g)} \rightarrow 2K_2O_{(s)}$	Floats on water and reacts very violently (explodes) to produce potassium hydroxide and hydrogen gas. $2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$	Burns vigorously in chlorine with a bright flame to give a white solid of potassium chloride. $2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$

Scientific Information

(a) (b)

Like other alkali metals, potassium reacts so strongly with water that, It must be stored in kerosene oil to prevent it from reacting with moisture in the air.

8.1.3 Alkaline Earth Metals

The elements of the group IIA are called Alkaline earth metals. The name of this group is due to they produce the alkalies and are widely distributed in the earth crust. The Alkaline earth metals have two electrons in their valance shells. Their valence sub-shell is "s". They are electropositive metals. They lose the two valence electrons and form M^{+2} ions. Their ionization energies are low.

There are six alkaline earth metals, including Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). They become stable by attaining the electron configuration of noble gasses by losing their outermost electrons. These metals are often found in the form of sulphates in nature. Examples include the minerals such as gypsum (calcium sulphate), epsom (magnesium sulphate) and barite (barium sulphate).

Occurrence of Alkaline Earth Metals

Alkaline earth metals have low ionization energies, so they are very reactive metals. That is why they do not occur free in nature. Beryllium occurs in nature in small amount in the form of Beryl. Magnesium and Calcium are very abundant in the earth crust. Magnesium and Calcium are present with Sodium and Potassium in rocks as cations. Magnesium halides are found in seawater. Magnesium is an important constituent of Chlorophyll. Calcium is found in nature in the form of Calcium phosphate and Calcium Fluoride. Calcium is the important constituent of living organism. It occurs as skeletal materials in bones, teeth, egg shells, etc. Radium is a rare element. It is radio active in nature.

Difference between Ionization Energies of Alkali and Alkaline Earth Metals

The minimum energy required to remove the valence electron from the outermost shell of gaseous atom to form the positive ion is called ionization potential or ionization energy. Ionization energy of an element is measured experimentally in joules or kilojoules per mole.

(a) Ionization Energies of Alkali metals

The elements of group I, except hydrogen are called alkali metals, and consist of the elements Li, Na, K, Rb, Cs and Fr. They are most electropositive elements. These metals have one electron in their outer most shell. The removal of this valence electron is very easy, which make it the most reactive metals.

$$Na \longrightarrow Na^{+} + 1e^{-} \qquad ?E=496Kj.mole^{-1}$$

In alkali metals, lithium is at the top of the group, has the highest ionization energy due to smallest atomic size. It has only two electronic shells. The distance between the nucleus and the valence shell is very small. So the removal of electron is very difficult and has higher ionization energy. Going down in the group, there is increase in atomic number, which result in the increase in atomic radius. Hence, the distance between the nucleus and valence shell increases due to which the valence electron is less firmly held by the nucleus and the removal of electron becomes easier. Thus the ionization energies decreases down the group, as shown in table 8.2.

Table 8.2 Ionization energies in kJ/mol of Group I (Alkali Metals)

Elements	Atomic Number	Atomic Radius Å	Ionization Energy (kJ/mol)
Li	3	1.55	520
Na	11	1.86	496
K	19	2.27	419
Rb	37	2.48	403
Cs	55	2.65	376

(b) Ionization Energies of Alkaline Earth Metals

The elements of group IIA, are called alkaline earth metals. These are the second most electropositive metals of all the elements (the alkali metals are the most electropositive). The alkaline earth metals have two electrons in their valence shell, so they lose two electrons to form M^{+2} cations.

If the alkaline earth metals are compared to the alkali metals, many dissimilarities are noticeable. The main difference is the electron configuration, which is nS^2 for alkaline earth metals and nS^1 for alkali metals. For the alkaline earth metals, the nucleus also contains an additional positive charge. Also, the elements of Group II (alkaline earth metal) have much higher melting and boiling points compared to those of Group I (alkali metals). The alkali metals are softer and more lighter weight than the alkaline earth metals.

The second valence electron is very important when it comes to comparing chemical properties of the alkaline earth and the alkali metals. The second valence electron is in the same "sublevel" as the first valence electron. This means that the elements of Group II have a smaller atomic radius due to more nuclear attraction and much higher ionization energy than those of Group I. Even though the Group IIA contains much higher ionization energy, they still form ionic compounds containing M^{+2} cations.

$$Mg \xrightarrow{Energy} Mg^{+1} + 1e^- \qquad \Delta E = 737.7\ kJmol^{-1}$$

To remove a second electron from the valence shell of an ion, more energy is required because of increased nuclear force of attraction.

$$Mg^{+1} \xrightarrow{Energy} Mg^{+2} + 1e^- \qquad \Delta E = 1450.7\ kJ\ mol^{-1}$$

Beryllium, however, behaves differently. This is due to the fact that it is at the top of the group and has the highest ionization energy in the group due to the small size.

Table 8.3 Ionization energies in kJ/mol of Group II (Alkaline earth metals)

Elements	Atomic Number	Atomic Radius Å	First Ionization Energy (kJ/mol) $M_{(g)} \rightarrow M^+_{(g)} + e$	Second Ionization Energy (kJ/mol) $M^+_{(g)} \rightarrow M^{2+}_{(g)} + e$
Be	4	1.12	899	1757.1
Mg	12	1.60	738	1450.7
Ca	20	1.97	590	1145
Sr	38	2.15	549	1064.2
Ba	56	2.24	503	965.1

Table 8.3, shows that the ionization potential values decreases down the group in each case. The atomic number increases down the group with the addition of one more shell for each period due to which atomic size increases. As a result, the distance between nucleus and the valence electrons increases and thus a regular decrease in ionization energies of elements occur.

A comparison of ionization energies of group I and group II elements show that the lower ionization energy values of group I elements make the ion formation easier in these elements and thus alkali metals are comparatively more reactive than the group II.

Moving from alkali metals to alkaline earth metals (in a period) the melting and boiling points, densities and hardness is increased. However, a decreasing trend in electropositivity, conductivity, reducing power and atomic radii is occurred.

Sodium

Sodium does not occur as a free metal in nature because it is too reactive and readily combines with other elements and compounds. It is found in the sea as sodium chloride, sodium bromide and sodium iodide. It is also found in deposits as rock salt.

Fig.8.2 Sodium metal

Note: When freshly cut, alkali metals shows a shiny and silvery surface that rapidly tarnishes in air.

(a) Position of sodium in the periodic table

Sodium belongs to alkali metals. Sodium atomic number is '11' and its symbol is "Na". It occupies first position in 3rd period and 3rd position in group I as it has three electronic shells and only one electron in the valence shell.

Table 8.4 Physical properties of sodium

S. No	Characteristic	Property
1.	Appearance	Silvery white solid
2.	Softness	Soft metal, can be cut with knife
3.	Density	0.971 g/cm^3
4.	Melting point	97.6°C
5.	Boiling point	880°C
6.	Tensile strength	Has relatively low tensile strength
7.	Lightness	Lighter than water, it floats on the surface of water
8.	Malleable and ductile	Malleable and ductile
9.	Conductivity	Good conductor of heat and electricity

(b) Chemical properties of sodium

Sodium is very reactive alkali metal. It is a strong reducing agent. Some of its simple reactions are given below.

i. Reaction with hydrogen

Sodium reacts with hydrogen to form sodium hydrides.

$$2\,Na_{(s)} + H_{2(g)} \longrightarrow 2NaH_{(s)}$$

ii. Reaction with oxygen

Sodium reacts with oxygen to form basic oxide (Na_2O), which on reaction with water forms alkali (NaOH). This is a characteristic reaction for metals, that is why the elements in this group are called alkali metals.

$$2\,Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{(s)}$$

$$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$

iii. Reaction with water

Sodium reacts with water vigorously, liberating hydrogen. The reaction is exothermic because of the heat produced, the liberated hydrogen catches fire.

$$2\,Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)} + Heat$$

Due to the high reactivity of sodium it is kept in kerosene oil or liquid paraffin to avoid its contact with air (oxygen and water).

iv. Reaction with halogens

Sodium reacts with halogens to form sodium halide.

$$2\,Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

$$2\,Na_{(s)} + Br_{2(l)} \longrightarrow 2\,NaBr_{(s)}$$

v. Reaction with sulphur

Sodium reacts with sulphur to form sodium sulphide.

$$2\,Na_{(s)} + S_{(s)} \longrightarrow Na_2S_{(s)}$$

vi. Reaction as reducing agent

Sodium is a powerful reducing agent. It reduces most of the oxides and halides.

$$2\,Na_{(s)} + MgO_{(s)} \longrightarrow Na_2O_{(s)} + Mg_{(s)}$$

$$4Na_{(s)} + TiCl_{4(s)} \longrightarrow 4NaCl_{(s)} + Ti_{(s)}$$

(c) Uses of sodium

Some of the uses of sodium are following.

I. It is used in the preparation of important compounds such as sodium carbonate(Na_2CO_3), sodium bicarbonate($NaHCO_3$), Sodium hydroxide ($NaOH$), sodamide($NaNH_2$) etc.

ii. It is used in sodium vapour lamps (which give a bright orange – yellow light) for street lighting.

iii. It is used as coolant in nuclear reactors.

iv. It is used in the purification of petroleum, in order to remove sulphur from it. This process is called desulphurization.

v. It is used as reducing agent to prepare metals such as Titanium, Zirconium etc. from their chloride and oxides.

vi. It forms alloys with other metals. It's most useful alloy is with mercury called sodium amalgam and with silver metal.

Magnesium

Magnesium is the member of alkaline earth metals. It occurs in nature only in combined state, as Dolomite ($CaCO_3.MgCO_3$), Kieserite ($MgSO_4$), Epsom salt ($MgSO_4.7H_2O$), in many silicates including talc and asbestos. Magnesium is present in sea water as chlorides and bromides. It is responsible for permanent hardness of water. It is also an essential constituent of chlorophyll in green plants.

(a) Position of Magnesium in the periodic table

Magnesium atomic number is 12 and its symbol is "Mg". It occupies second position in 3rd period and 2nd position in group IIA as it has three electronic shells and two electrons in the valence shell.

Table 8.5 Physical properties of Magnesium

S. No	Characteristic	Property
1.	Appearance	Silvery grey solid
2.	Density	1.74 g cm^3
3.	Melting point	651 $^{\circ}C$
4.	Boiling point	1106 $^{\circ}C$
5.	Malleable and ductile	Malleable and ductile
6.	Conductivity	Good conductor of heat and electricity

Calcium

Calcium is too reactive to occur as free metal in nature. It occurs abundantly in the combined state in minerals such as calcium carbonate ($CaCO_3$), in lime stone, marble, chalk and as calcium sulphate ($CaSO_4$) in gypsum etc.

Pure calcium

Marble

Fig.8.3 Calcium forms

(a) Position of calcium in the periodic table

Calcium atomic number is 20 and its symbol is "Ca". It occupies second position in 4^{th} period and 3^{rd} position in group II, as it has four electronic shells and two electrons in the valence shell.

Table 8.6 Physical properties of calcium

S. No	Characteristic	Property
1.	Appearance	Silvery white solid
2.	Density	$1.55\ g/cm^3$
3.	Melting point	$851^{\circ}C$
4.	Boiling point	$1487^{\circ}C$
5.	Malleable and ductile	Malleable and ductile
6.	Conductivity	Good conductor of heat and electricity

(b) Chemical properties of magnesium and Calcium

i. Magnesium and calcium combines directly with hydrogen forming hydrides.

$$Mg_{(s)} + H_{2(g)} \xrightarrow{\text{Heating}} MgH_{2(s)}$$

$$Ca_{(s)} + H_{2(g)} \longrightarrow CaH_{2(s)}$$

ii. Both magnesium and calcium burns in air, magnesium burns with dazzling flame forming magnesium oxide (MgO), commonly known as magnesia, while calcium forms calcium oxide (CaO), with a characteristic brick red color flame.

$$2Mg_{(s)} + O_{2(g)} \xrightarrow{\text{Heating}} 2MgO_{(s)}$$

$$2Ca_{(s)} + O_{2(g)} \xrightarrow{\text{Heating}} 2CaO_{(s)}$$

When these oxides, magnesium oxide (MgO) and calcium oxide (CaO) are dissolved in water, they form basic solutions.

$$2MgO_{(s)} + H_2O_{(l)} \longrightarrow 2Mg(OH)_{2(aq)}$$

$$2CaO_{(s)} + H_2O_{(l)} \longrightarrow 2Ca(OH)_{2(aq)}$$

iii. Magnesium and calcium reacts with nitrogen (N_2) forming their respective nitrides.

$$3\,Mg_{(s)} + N_{2(g)} \xrightarrow{\text{Heating}} Mg_3N_{2(s)}$$

$$3\,Ca_{(s)} + N_{2(g)} \xrightarrow{\text{Heating}} Ca_3N_{2(s)}$$

iv. Magnesium and calcium reacts with acids liberating hydrogen gas.

$$Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(s)} + H_{2(g)}$$

$$Mg_{(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(s)} + H_{2(g)}$$

$$Ca_{(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + H_{2(g)}$$

$$Ca_{(s)} + H_2SO_{4(aq)} \longrightarrow CaSO_{4(aq)} + H_{2(g)}$$

v. Magnesium and calcium reacts with halogen to form their respective halides.

$$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$$

$$Mg_{(s)} + Br_{2(l)} \longrightarrow MgBr_{2(s)}$$

$$Ca_{(s)} + Cl_{2(g)} \longrightarrow CaCl_{2(s)}$$

$$Ca_{(s)} + Br_{2(l)} \longrightarrow CaBr_{2(s)}$$

(c) Uses of Magnesium and Calcium

(i) Uses of Magnesium

Some of the uses of Magnesium are following.

i. Magnesium is low density metal, so it is used in formation of light but tough alloys, such as Duralumin (a mixture of Al, Cu, Mg and Mn) and magnalium (a mixture of Al, Mg). These alloys are used for construction of aircrafts, cars and moving parts of machines.

ii. It is also used in photographic flashlight powder, flames and fireworks.

iii. It is used as deoxidant in metallurgy and in the extraction of Titanium and Uranium.

iv. Its compounds such as magnesium oxide (MgO) are mixed with clay, to make refractory bricks for furnace lining.

v. Magnesium sulphate ($MgSO_4$) is used in textile, paper industry, soap formation and pharmaceutical industries etc.

(ii) Uses of Calcium

Some of the uses of Calcium are following.

i. Calcium is used as a deoxidant in steel castings and copper alloys.

ii. It is used in the making of calcium fluoride and calcium hydride and in the extraction of uranium.

iii. Their compounds such as lime (CaO) is added to soil as fertilizer to decrease its acidity. It is also used for water softening, pollution control and in pulp, paper, sugar and glass manufacturing industries.

iv. It is used in steel making.

8.1.4 Soft and Hard Metals

Elements of group I, group II, transition metals, lanthanides and actinides all are metals. But there is much difference in the properties of these metals. The metals present in group I and II are soft, very reactive, having low ionization energies and very electropositive are known as soft metals. Metals like copper, silver, iron etc are less reactive, having high ionization energies, less electropositive are known as hard metals/noble metals. Their melting points, boiling points and densities show much higher values. A comparison of properties of sodium (soft metal) and Iron (hard metal) representative is given below in table 8.7.

Table 8.7 Comparison of properties Sodium (Na) and Iron (Fe)

Sodium (Na)	Iron (Fe)
Sodium is an alkali metal, with atomic number 11.	Iron is a transition metal, with atomic number 26.
One electron in its outer most shell and is very soft and can be cut with knife.	Hard and requires great energy to break.
Weak attractive force between the atoms of sodium.	Strong attractive force between the atoms of iron.
The melting point of sodium is 97.6°C.	Melting point of iron is 1538°C.
Boling point of sodium is 880°C.	Boling point of iron is 2862°C.
Low density $0.927 g/cm^3$.	Higher density $7.874 g/cm^3$.
Lighter and floats on the surface of water.	Heavy and settles at the bottom of water.
Low tensile strength, cannot be used where stress is required.	High tensile strength can be used in construction of building and bridges. It is also used to prepare steel.
Very reactive, stored in kerosene oil.	Less reactive than sodium.

8.1.5 Inertness of Noble Metals

Group I and group II metals are fairly reactive because of their low ionization potential but the transition metals have a tendency of low reactivity. Their low reactivity is due to their high ionization energies.

Noble metals are those metals, which resist oxidation and corrosion in moist air (unlike most base metals). Chemically noble metals are Ruthenium (Ru), rhodium (Rh), palladium (Pd), silver (Ag), osmium (Os), iridium (Ir), platinum (Pt) and gold (Au). Some chemists include mercury (Hg), rhenium (Re) and copper (Cu) as noble metals. These metals are relatively inert and found free in nature.

The noble metals are valuable because of their inertness, rarity in the Earth's crust and their usefulness in areas like metallurgy, high technology and ornamentation (jewellery, art, etc.).

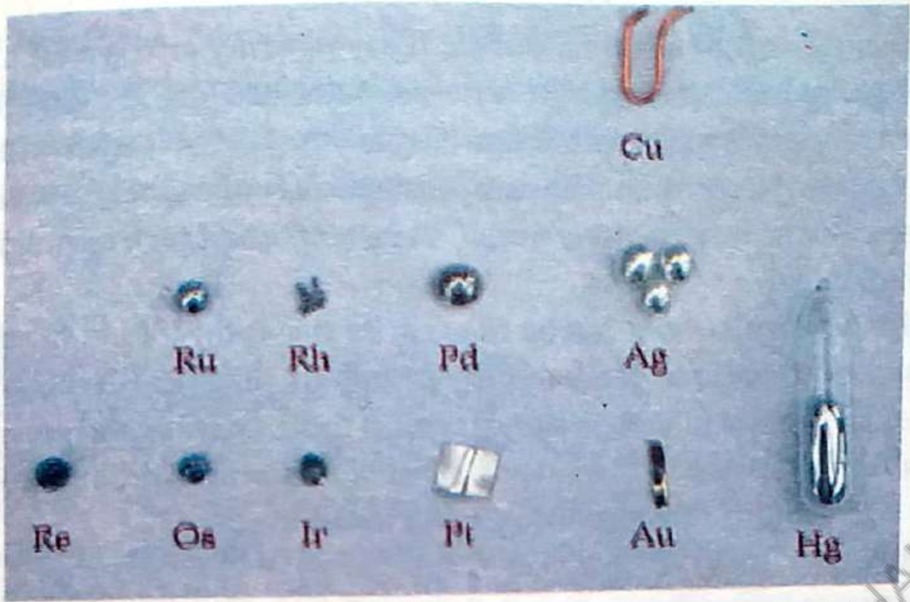


Fig.8.4 Noble metals

Commercial value of Silver (Ag), Platinum (Pt) and Gold (Au)

Those metals which are expensive and have great commercial and economic values are precious metals. In Nobel metals particularly Silver (Ag), platinum (Pt) and gold (Au) are considered as precious metals.

(i) Silver (Ag)

Silver is a soft, white metal that usually occurs in nature in one of four forms, a) as a native element, b) as a primary constituent in silver mineral, c) as a natural alloy with other metals, and d) as a minor constituent in the ores of other metals.

Silver is known as a precious metal because it is rare and has a high economic value. It is valuable because it has a number of physical properties that make it the best possible metal for many different uses.

Pure silver is very soft. It is usually mixed with copper to form an alloy for making commercial articles. This alloy is used to make coins, jewellery and tableware. Silver chloride combined with silver bromide is used in photography. Silver is drawn into sheets and wires. It has higher electrical and thermal conductance and reflectivity than any other metal.

(ii) Platinum (Pt)

The name platinum comes from the Spanish word platina meaning little silver. Platinum is the 72[nd] most common element in Earth's crust. That is why, platinum is an expensive metal.

Platinum is heavy, soft, malleable, ductile (easy to draw into wires) and has a fairly high melting point (1770°C). It is noble metal because it is un-reactive. It does not even react with oxygen in air and it is resistant to react with acids.

Platinum is used in the catalytic converters to remove pollutants from car engine exhaust gases. But as an expensive metal, so other metals such as palladium etc are used in its place.

The ease with which platinum can be shaped, its strength, colour, hardness and inertness makes it suitable for jewellery and gem setting. Un-reactivity also makes it useful in dental fillings, making surgical tools, coating and apparatus for scientific laboratories. Apart from that, platinum is also used in the electrical industry, in lasers and in making photographic materials.

(iii) Gold (Au)

Gold has been used to make ornamental objects and jewellery for thousands of years. Special properties of gold like very high luster, attractive colour, inertness, tarnish resistivity, ability to be drawn into wires, hammered into sheets or cast into shapes etc. make it perfect for manufacturing of jewellery

Pure gold is too soft to resist the stresses applied to many jewellery items. Alloying gold with other metals such as copper, silver and platinum increase its durability. In olden days the coins were made of gold. Gold coins were commonly used in transactions up to paper currency became a more common form of exchange.

Gold is used in standard desktop or laptop computers. The rapid and accurate transmission of digital information from one component to another requires an efficient and reliable conductor. Gold meets these requirements better than any other metal. The importance of high quality and reliable performance justifies the high cost. Gold alloys are used for dental fillings, tooth crowns and orthodontic appliances. Gold is used in dentistry because it is chemically inert, non-allergic and easy for the dentist to work.

8.2 Non-Metals

Non-metals are those substances which are nonconductor of heat and electricity. They are neither malleable nor ductile. Their oxides and hydroxides are acids. Non-metals exist in all the three states of matter at room temperature.

A non-metal is one which burns in oxygen and produce an oxide which is usually acidic. When dissolve in water form an acidic solution that turns blue litmus paper red.

Scientists redefined the non-metals on the basis of loss and gain of electron. **A non-metal is an element which gains an electron and forms an anion.** For example chlorine, a non-metal gains an electron and form chloride ion (Cl^-). Non-metals are electronegative and so they have the tendency to gain electron(s).

$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$

Society, Technology and Science

A noble metal is a rare element of high durability and economic value. The best known noble metals are gold and silver. Although both have industrial uses, but they are better known for their uses in art, jewellery and coinage. Other precious metals include the platinum group metals: Rhenium, Rhodium, Palladium, Osmium, Iridium and Platinum, of which Platinum are the most widely traded. Plutonium and Uranium may also be considered precious metals. Chemically, the noble metals are less reactive than most of other metals.

Historically, metals such as Copper, Silver and Gold have been used for coins and jewellery. These metals have high luster and malleability. Copper is common coinage metal. Coins are usually made from Cu – Ni alloy.

Scientific Information

- The non-metals include hydrogen of group I, boron of group III, carbon and silicon of group IV, nitrogen and phosphorus of group V, all members of group VI, VII and VIII, which make a total of 27. Most are gases, (Argon, chlorine, fluorine, hydrogen, helium, krypton, neon, nitrogen, oxygen, radon and xenon). One is liquid (bromine) and few are solids (e.g. carbon, iodine, phosphorous, selenium, silicon and sulphur).
- Three metals mercury (Hg), M.P 38.87°C, Galium (Ga) M.P. 29.76°C and cesium M.P. 28.4°C occur in liquid form. Only one non-metal, Bromine (Br) M.P. 7.2°C occurs in the liquid state.

Table 8.8 Comparison of metals and non-metals

Metals	Non-metals
1. They are electron donor during chemical reactions. They are reducing agents.	They are electron acceptors during chemical reactions. They are oxidizing agents.
2. They become positively charged ion in solutions.	They become negatively charged ion in solutions.
3. Some metals can replace hydrogen from acids to form salts.	Non-metals cannot replace Hydrogen from acids.
4. They form basic oxide. The soluble basic oxides form alkalis when they dissolve in water.	They form mainly acidic oxides which dissolve in water to form acids.
5. They form electrovalent (ionic) chlorides.	They form covalent chlorides.
6. They do not combine easily with Hydrogen. Few hydrides formed are electrovalent.	They combine easily with Hydrogen to form many stable covalent hydrides.
7. Usually solids at room temperature (except Hg).	Often gases (except Br(l), S(s), P(s), I(s), C(s), B(s) and Si(s).
8. Good conductor of heat and electricity.	Poor conductors of heat and electricity (except graphite)

8.2.1 Electronegative Character

All non-metallic elements have the ability to gain electrons from other elements and hence are electronegative in nature. **Electronegativity is a measure of an atom's ability to attract the shared pair of electrons of a covalent bond to itself.**

In short, we can say that the elements placed on the right corner of the periodic table will be the most electronegative elements, which are nonmetals. At the right corner of the periodic table is the group VII, which is the most electronegative in the non-metals.

8.2.2 Halogens

Group VII elements are collectively called halogens. Halogen family consists of five elements fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (At). The word 'halogen' is derived from Greek word means 'salt-forming'. Halogens are found in the environment only in the form of ions or compounds, because of their high reactivity. They do not exist in the elemental form in nature.

Halogens are nonmetals. At room temperature, fluorine and chlorine are gases, bromine is a liquid, iodine and astatine are solids. Halogens are very reactive, the reactivity decreases from fluorine to astatine. Fluorine is the most electronegative element. Astatine is radioactive with short half-lives.

Halogens Reactivity

The electronic configuration of group VII (Halogens) is nS^2nP^5. They need one electron to complete their octet, by gaining it from less electronegative atom. By gaining electron they form anions, commonly known as halide ions.

$$X_2 + 2e^- \longrightarrow 2X^-$$

Where $\bar{X}$ is halide ion. For example, chloride ion is formed as follows,

$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$

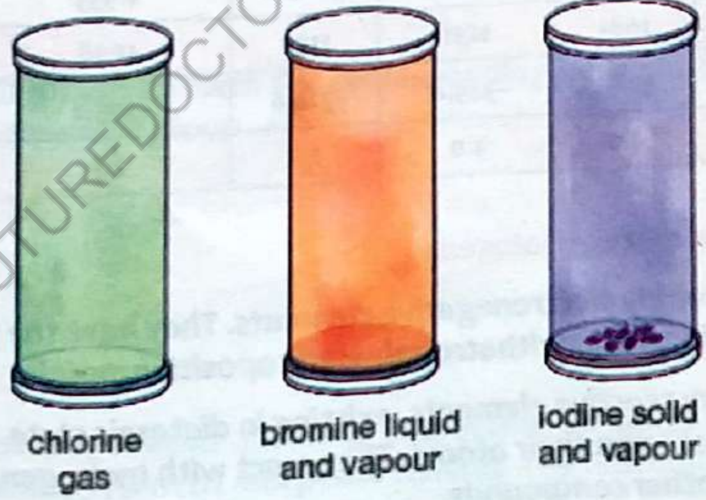


Fig.8.5 Different halogens

(a) Physical Properties of Halogens

The group of halogens is the only group in periodic table which contains elements in all three states of matter at standard temperature and pressure. They are poor conductors of heat and electricity. Some of the characteristics properties of halogens are given in table8.9.

Table 8.9 Characteristics Properties of Halogens

Name and Symbol	Fluorine (F)	Chlorine (Cl)	Bromine (Br)	Iodine (I)	Astatine (At)
Atomic No.	9	17	35	53	85
Electronic arrangement	$1s^2\ 2s^2\ 2p^5$	$[Ne]3s^2\ 3p^5$	$[Ar]3d^{10}\ 4s^2\ 4p^5$	$[Kr]4d^{10}\ 5s^2\ 5p^5$	$[Xe]4f^{14}\ 5d^{10}\ 6s^2\ 6p^5$
Physical state	pale yellow gas	greenish gas	dark reddish liquid	black solid and when heated it forms a purple vapour	black solid
Melting point(°C)	−219.6	−101.5	−7.2	113.7	301.8
Boiling point (°C)	−188.2	−34.04	58.8	184.3	336.8
Atomic radius (pm)	71	99	114	133	150
Density (g/dm^3)	0.0017	0.0032	3.1028	4.933	-
I.P ($kj.mol^{-1}$)	1681	1251	1140	1008	890±40
E.A ($kj.mol^{-1}$)	-328.0	−349.0	−324.6	−295.2	−270.1
Electronegativity	4.0	3.0	2.8	2.5	2.2

(b) Chemical Reactivity of Halogens

Halogens are strongly electronegative elements. They have the tendency to form negative ions, when react with strongly electropositive metals.

Halogens are very reactive elements, existing in diatomic state. They form single covalent bond between their atoms. They react with hydrogen, oxygen, metals, non-metals and other compounds.

i. Reaction with Hydrogen

Halogen reacts with hydrogen to form the respective hydrides. Fluorine reacts violently, chlorine reacts in the presence of sunlight, bromine reacts on heating, and iodine reacts reversibly with hydrogen.

HF differs from other hydrogen halide as it is liquid at room temperature due to strong hydrogen bonding. These hydrides act as a strong reducing agents.

$$H_{2(g)} + F_{2(g)} \longrightarrow 2HF_{(l)}$$

$$H_{2(g)} + Cl_{2(g)} \xrightarrow{\text{Sunlight}} 2HCl_{(g)}$$

$$H_{2(g)} + Br_{2(l)} \xrightarrow{\Delta} 2HBr_{(l)}$$

$$H_{2(g)} + I_{2(s)} \rightleftharpoons 2HI_{(l)}$$

ii. Reaction with Oxygen

Fluorine reacts with oxygen to form fluorine monoxide and dioxide, while other halogens form oxides like, Dichlorine heptoxide (Cl_2O_7), tribromine tetraoxide (Br_3O_4) and Iodine pentoxide (I_2O_5).

$$O_{2(g)} + 2F_{2(g)} \longrightarrow 2OF_{2(g)}$$

$$O_{2(g)} + F_{2(g)} \longrightarrow O_2F_{2(g)}$$

iii. Reaction with Metals

Halogens reacts with metals to form the corresponding ionic halides. The ionic character decreases down the group.

$$Cu_{(s)} + Br_{2(l)} \longrightarrow CuBr_{2(s)}$$

$$2K_{(s)} + I_{2(s)} \longrightarrow 2KI_{(s)}$$

iv. Reaction with Non-metals

Halogens react with non-metals forming respective covalent halides, for example, with phosphorus form tri and penta-halides. The reactivity of reaction decreases from fluorine to iodine.

$$2P_{(s)} + 3Cl_{2(g)} \longrightarrow 2PCl_{3(aq)}$$

$$2P_{(s)} + 3Br_{2(l)} \longrightarrow 2PBr_{3(aq)}$$

$$2P_{(s)} + 5Cl_{2(g)} \longrightarrow 2PCl_{5(aq)}$$

$$2P_{(s)} + 5Br_{2(l)} \longrightarrow 2PBr_{5(aq)}$$

v. Reaction with other Compounds

Halogens oxidize certain compounds and itself reduced during the reaction.

$$H_2S_{(aq)} + Cl_{2(g)} \longrightarrow 2HCl_{(aq)} + S_{(s)}$$

$$2NH_{3(aq)} + 3Cl_{2(g)} \longrightarrow 6HCl_{(aq)} + N_{2(g)}$$

Chlorine reacts with lime water, $Ca(OH)_2$, to form bleaching powder.

$$Ca(OH)_{2(aq)} + Cl_{2(g)} \longrightarrow CaOCl_{2(aq)} + H_2O_{(l)}$$

Thus they act as an oxidizing agent but oxidation power decreases down the group with increasing size of the halogens.

vi. Displacement reaction

A more reactive halogen will displace a less reactive halogen from its halide solution. The reactivity of halogen decreases down the group.

Order of reactivity F > Cl> Br > I > As.

$$2NaBr_{(aq)} + Cl_{2(g)} \longrightarrow 2NaCl_{(aq)} + Br_{2(l)}$$

Activity:8.2

Make a list of elements which are found in uncombined state in nature.

Test Yourself.

Describe and explain what will happen when bromine water is added to pottasium iodide solution.

$$Br_{2(aq)} + 2\,KI_{(aq)} \longrightarrow I_{2(aq)} + \underline{\hspace{2cm}}$$

B. *Infer the identity of an element.*

(I) Element found in period 3.

(Ii) Will displace iodine from sodium iodide.

(Iii) Greenish gas, which is denser than air.

(Iv) Used in swimming pool to kill germs.

Key Points

- *Elements of group I (except hydrogen) group II, group III (except boron) three elements from each of IV and V, lanthanide, actinide and d-block / transition elements are all metals.*
- *Metals are electropositive in nature.*
- *Non-metals are electronegative in nature.*
- *Elements of group I (except hydrogen) are called alkali metals, elements of group II are called alkaline earth metals.*
- *Alkali metals are highly electropositive elements.*
- *Alkali metals are very reactive, found in combined state.*
- *Elements of group VII are called halogens.*
- *Halogens are highly electronegative elements.*
- *Halogens are very reactive, found in combined state.*
- *Sodium is soft and iron is hard metal.*
- *Noble metals are Ruthenium, rhodium, palladium, silver, osmium, iridium, platinum, gold, mercury, rhenium and copper.*
- *Silver, platinum and gold are precious metals.*

Exercise

Choose the correct option.

i. **Bromine is a non-metal in:**
   a. Solid state b. Liquid state
   c. Gaseous state d. Plasma state

ii. **Halogens react with metals to form:**
   a. Halides b. Oxides
   c. Halogen sulphides d. Hydrogenated compounds

iii. **Alkali metals are:**
   a. Oxidising agents b. Dehydrating agents
   c. Reducing agents d. All the above

iv. **Among the alkali metals, the metal with the highest ionization potential is:**
   a. Na b. Li
   c. Rb d. Cs

v. **The halogen present in the solid form is :**
   a. Chlorine b. Fluorine
   c. Iodine d. Bromine

vi. **Tendency of a metal to lose electron is called:**
   a. Electronegativity b. Electropositivity
   c. Electroplating d. Electrolysis

vii. **The word Alkali means:**
   a. Base b. Basic salt
   c. Acid d. Ashes

Viii. **The oxide of calcium CaO is:**
   a. Acidic b. Basic
   c. Amphoteric d. Neutral

ix. **Which one of the following is not an alkali metal:**

a. Francium b. Cesium

c. Rubidium d. Radium

x. **Group II elements are named as Alkaline Earth Metals because:**

a. Their oxides are basic

b. Their oxide and hydroxides are water soluble

c. Both a and b

d. They are found in earth

SHORT QUESTIONS

Answer briefly the following questions.

i. Identify at least two groups which contain only metallic elements.

ii. Write down the reaction of group -I metals with oxygen, with balance equations.

iii. State the physical properties of metals.

iv. How does sodium act as reducing agent and write down its reactions also.

v. Ionization energy of alkaline earth metals is higher than alkali metals, why?

vi. Pure gold is not used for ornaments, why?

vii. What are the uses of magnesium?

viii. Write down the reaction of chlorine with sodium hydroxide with balance equation.

ix. How does ionization energies values vary in group?

x. What happens during displacement reaction in halogens?

LONG QUESTIONS

i. Compare and contrasts the properties of alkali and alkaline earth metals, with reactions.

II. a. Differentiate between hard and soft metals.

b. Give the reaction of magnesium with,

(i) H_2 (ii) HCl (iii) O_2 (iv) H_2O (v) Cl_2

III. Discuss the reasons why some elements exist as free elements in nature while other occurs in combined states as compounds. Give two examples of each type.

IV. Define metal and non-metal and compare the properties (both physical and chemical) of metals and non-metals.

V. Halogens are very reactive elements, write down halogen's reactions with hydrogen, oxygen, metals, non-metals and other compounds.

Project Work

Prepare separate lists of compounds made of alkali and alkaline earth metals. Identify their uses and importance in daily life and as raw material in industries. Present it in the form of charts in your classroom.

GLOSSARY

Absolute temperature	*A temperature, measured on the Kelvin scale that must be used for gas law and certain other types of calculations.*
Alcohol	An organic compound with the - OH functional group e.g. CH_3OH.
Alkali Metals	Metals in group IA of the periodic table e.g. Li, Na, K, Rb, Cs, and Fr.
Allotrope	One of two or more forms of an uncombined element; for example, diamond and graphite are allotropes of carbon.
Alpha Particle	A helium nucleus generated in a nuclear reaction; a stream of such particles is referred to as an alpha ray.
Analytical Chemistry	The branch of chemistry dealing with the composition of samples.
Angular Molecule	A three-atom molecule in which the atoms do not lie on a line e.g. H_2O molecule.
Anion	A negatively charged ion.
Anode	The electrode where oxidation occurs in an electrochemical reaction.
Aqueous Solution	Any solution in which water is the solvent.
Atmosphere	The envelope of air surrounding the Earth.
Atmospheric Pressure	The pressure of the atmosphere. A unit equal to 760 torr and abbreviated as atm that is the pressure of the atmosphere on a "normal" day at sea level.
Atom	The smallest particle that retains the characteristic composition of an element.
Atomic Mass	The weighted average of the masses of the naturally occurring isotopes of an element, compared with one twelfth of the mass of a ^{12}C atom.
Atomic Mass Unit	A mass equal to one twelfth of the mass of a ^{12}C atom; abbreviated as amu.
Atomic Number	The number of protons in the nucleus of each atom of an element.

Atomic Size	The size of an atom.
Avogadro's Number	6.022×10^{23}, is the number of C-12 atoms in exactly 12 g of C-12.
Battery	A combination of two or more galvanic cells.
Binary Compound	A compound composed of two elements e.g. NaCl.
Biochemistry	The branch of chemistry dealing with living things.
Bohr theory	The first theory of the atom to propose that electrons in atoms were in definite energy levels.
Boiling Point	The temperature at which a liquid changes to a gas at the prevailing pressure.
Boyle's Law	At constant temperature, the volume of a given sample of gas is inversely proportional to its pressure.
Catalyst	A substance that affects the speed of a chemical reaction without any permanent change in its own composition.
Cathode	The electrode where reduction occurs in an electrochemical reaction.
Cation	A positively charged ion.
Cell	An electrochemical apparatus in which an oxidation half-reaction occurs at the anode and at the same time a reduction half-reaction occurs at the cathode resulting in (1) a chemical reaction creating electricity (voltaic cell), or (2) electricity creating a chemical reaction (electrolytic cell).
Charge	A quantity of electricity, measured in coulombs or faradays.
Charles' Law	At constant pressure, the volume of a given sample of gas is directly proportional to its absolute temperature.
Chemical Change	A chemical reaction. Irreversible change.
Chemical Property	A property related to changes in the composition of a substance.
Chemical Reaction	A change in which the composition (or structure) of one or more substances is changed.
Chemistry	The study of the interaction of matter and energy and the changes that matter undergoes.

Compound	A chemical combination of elements that has a definite composition and its own set of properties.
Concentration	The quantity of solute per unit volume of solution or per unit mass of solvent.
Condensation	A change of phase from gaseous to liquid or solid.
Covalent Bond	A bond resulting from electron sharing.
Crystalline Solid	A solid with a regular internal structure of repeating units and a definite melting point.
Dalton's Atomic Theory	The theory that matter is made up of small particles (atoms) that have properties which are characteristic of an element.
Daniel cell	A cell with zinc and copper electrodes.
Definite composition	The given ratio by mass of each element in a compound to any other element in the compound.
Density	The mass per unit volume of a sample of matter.
Diatomic Molecule	A molecule containing two atoms e.g. H_2.
Direct Proportionality	The relationship in which one variable changes by the same factor as another.
Dissociation	Separation of ions from their close proximity in a solid lattice to a distance when dissolved in a solvent.
Dissolve	To go into solution, making a homogeneous mixture.
Double Bond	The sharing of two pairs of electrons between two atoms.
Electrochemistry	The branch of chemistry dealing with the interaction of chemical reactions and electricity.
Electrode	Electrodes are the conductors i.e. metallic plates, wires or rods, through which electrons enters or leaves the electrolytes in a cell.
Electrolysis	Production of a chemical reaction by means of an electric current.
Electrolytic Cell	An apparatus in which electric energy can be used to produce chemical reactions.

Electron	A negatively charged subatomic fundamental particle of an atom revolves around nucleus in orbit.
Electron Affinity	The energy liberated when a gaseous atom acquires an electron to form a gaseous anion.
Electron Sharing	The sharing of electrons between atoms to form covalent bonds.
Electronegativity	The tendency of an atom to attract electron pair in the formation of a covalent compound.
Electronic Configuration	The arrangement of the electrons in an atom, ion or molecule.
Element	A substance that cannot be broken down into simpler substances by chemical means, one of the basic building blocks of which all matter is composed.
Family	In the periodic table, a column that includes elements with similar chemical properties. Also known as periodic group.
Formula	A combination of symbols and subscripts that identifies the composition of an element, compound or ion.
Formula Mass	The relative mass of one formula unit compared to the mass of a ^{12}C atom, which is defined as exactly 12 amu.
Formula Weight	Some of masses of all the atoms present in one formula unit.
Freezing	Changing from a liquid to a solid state.
Galvanic Cell	A voltaic cell.
Gas	A state of matter; a substance that attain the volume and shape of its container.
Group	In the periodic table, a column that includes elements with similar chemical properties; a family.
Halogen	An element of periodic group VIIA: F, Cl, Br, I and At.
Heterogeneous Mixture	A physical combination of substances having distinguishable parts.
Homogeneous Mixture	A physical combination of substances whose parts are not distinguishable, even with the best optical microscope; a solution.

Hydrate	A compound that has water molecules bonded in it.
Hydrogen Bonding	The intermolecular force resulting from the attraction of a hydrogen atom on one molecule to a small and highly electronegative atom (F, O, or N) on another molecule (or the same molecule).
Inorganic Chemistry	The branch of chemistry dealing with compounds other than those containing C – C or C – H bonds or both.
Intermolecular Force	An attraction between molecules.
Inverse Proportionality	A relationship in which one variable gets smaller by the same factor as another gets larger or vice versa.
Ion	A charged atom or group of atoms.
Ionic Bond	The attractive force between oppositely charged ions.
Ionic Size	The size of an ion.
Ionic Solid	A solid consisting of ions.
Ionization	Reaction of a substance with solvent to produce ions, as, for example, the reaction of a strong acid with water.
Kelvin Scale	The temperature scale with 273 K as the freezing point of water and 373 K as the boiling point of water; the scale required for gas law and certain other scientific calculations.
Law	A generalized statement that summarizes a collection of observations.
Linear Molecule	A molecule whose atoms all lie on a line like CO_2 molecule.
Liquid	A state of matter; a sample of matter that has a definite volume but assumes the shape of its container.
Lone Pair	An unshared pair of electrons on a bonded atom.
Malleable	Able to be hammered or pressed into shape without breaking or cracking
Mass Number	The sum of the number of protons and the number of neutrons in an atom: the distinguishing difference among isotopes of a given element.
Matter	Anything that has mass and occupies space.

Melting	*A change of phase from solid to liquid by heating.*
Metal	An element on the left side of the periodic table. An electron donors specie.
Metallic Solid	A solid consisting of atoms of one or more elemental metals.
Mixture	A physical combination of substances having properties of its components e.g. air.
Molar	The unit of molarity; abbreviated M.
Molar Mass	The mass in grams of 1 mol of a substance.
Molarity	A measure of concentration defined as the number of moles of solute per liter of solution; its symbol is M.
Mole	The chemical unit of quantity for any substance; equal to 6.02×10^{23} particles (atoms, molecules or formula units) of the substance; abbreviated as mol.
Molecular Formula	The formula of a molecular substance that gives the ratio of atoms of each element in a substance molecule.
Molecular Mass	The relative mass of a molecule of a substance compared to the mass of a C-12 atom.
Molecular Shape	The spatial arrangement of the atoms in a molecule.
Molecular Weight	Molecular mass of all the elements present in the molecule.
Monatomic Ion	An ion consisting of one atom only.
Neutron	A subatomic particle that has no charge and a mass slightly greater than 1 amu. Present in the nucleus along proton.
Noble Gas	An element of periodic group zero: He, Ne, Ar, Kr, Xe and Rn.
Noble gas Configuration	Electronic configurations like that of a noble gas, with 8 electrons (or 2 for very light elements) in the outermost shell.
Non Electrolyte	A compound that is not ionized at all, even in aqueous solution.
Nonmetal	Hydrogen or any element on the right hand side of the periodic table.
Nucleus	The center of an atom, consisting of the protons and neutrons.
Octet	A set of 8 electrons in the outermost shell of an atom or ion.
Octet Rule	Atoms or ions with an octet are stable.

Orbit	*As described by Bohr, the circular path for electrons in an atom around nucleus having specific energy value.*
Orbital	*A part of a sub-shell of an atom.*
Outermost Shell	*The highest shell that contains electrons. Also known as valence shell.*
Oxidation	*An increase in oxidation number.*
Oxidation Number	*A number assigned to an element in chemical combination which represents the number of electrons lost (or gained, if the number is negative), by an atom of that element in the compound.*
Oxidation-Reduction Reaction	*A reaction in which the oxidation number of one element is raised and the oxidation number of another element (or the same element) is lowered.*
Oxidation State	*The oxidation number.*
Oxidizing Agent	*A species that can increase the oxidation number of another reactant.*
Partial Pressure	*The pressure of one gas in a mixture of gases.*
Percent	*Parts per hundred parts.*
Percent by Mass	*One hundred percent times the ratio of the mass of an element in a compound divided by the mass of the compound.*
Period	*The seven horizontal rows of the periodic table.*
Periodic Table	*The arrangement of elements in order of their increasing atomic number, with elements having similar chemical properties together in vertical columns.*
Phase	*A state of matter: solid, liquid or gas.*
Phase Change	*A transition from one to another of the three states of matter, for example, from gas to liquid.*
Physical Change	*A process in which no change in composition occurs. Reversible change.*
Physical Chemistry	*The branch of chemistry dealing with the properties of substances.*
Polar Bond	*A covalent bond in which there is unequal sharing of electrons.*
Polar Molecule	*A molecule that has a permanent dipole.*
Polyatomic Ion	*An ion composed of two or more atoms.*

Pressure	Force divided by area.
Property	A characteristic of a substance.
Proton	A subatomic particle with a mass slightly greater than 1 amu and a charge of +1 allocated in the nucleus.
Pure Substance	An element or compound.
Reactant	Any substance that undergoes a chemical reaction and thus appears on the left-hand side of a chemical equation.
Reactive	Having a high tendency to undergo chemical reaction.
Redox Reaction	An abbreviation for oxidation reduction reaction.
Reducing Agent	A species that reduces the oxidation number of another reactant.
Reduction	The decrease in oxidation number, addition of hydrogen or electron.
Salt Bridge	A solution of an inert electrolyte that connects two half-cells in order to complete the circuit in a voltaic cell.
Saturated Solution	Solution which cannot dissolve further amount of solute at a particular temperature is called saturated solution at that temperature.
Single Bond	A covalent bond formed by a single pair of shared electrons.
Solid	A state of matter; a substance that has definite shape and volume.
Solubility	The concentration of a saturated solution at a given temperature.
Solute	The component of a solution that is dissolved in another component, the solvent. It is always in less amount.
Solution	A homogeneous mixture.
Solvent	The component of a solution that does the dissolving. It is always in larger amount.
Standard	A basis for comparison, such as the mass of C-12.
Standard Atmosphere	A pressure of 760 mm of Hg.

Standard State	*The normal state for a substance at 1 atm and the temperature involved; for example, the standard state for oxygen at 25°C is gaseous O_2 molecules.*
State	*A phase in which matter exists: solid, liquid or gas.*
Strong Electrolyte	*A compound that is fully ionized in aqueous solution e.g. NaOH.*
Subatomic Particle	*A proton, neutron, electron, deuteron, positron, alpha particle, etc.*
Sub-shell	*The portion of a shell characterized by the same principal quantum number and the same angular momentum quantum number.*
Supersaturated Solution	*A solution which contains more amount of solute than the amount required to saturate it at a particular temperature is called supersaturated solution.*
Symbol	*A one or two letter representation of an element.*
Temperature	*The intensity of heat in a body.*
Tetrahedral Molecule	*A molecule with bonded atoms oriented toward the corners of a tetrahedron (a solid object with four sides, all of which are identical equilateral triangles).*
Theory	*A generally accepted explanation for a law (or series of observations).*
Torr	*A unit of pressure equal to 1 mm of Hg.*
Trigonal Planar molecule	*A molecule with atoms oriented toward the corners of an equilateral triangle, with the central atom in the same plane.*
Trigonal Pyramidal Molecule	*A molecule with atoms oriented toward the corners of an equilateral triangle, with the central atom out of that plane.*
Triple Bond	*A covalent bond consisting of three pairs of electrons shared between two atoms.*
Tritium	*The isotope $^3{}_1H$.*
Unit	*A standard division of measure having a certain value; for example, the meter is the primary metric unit of length.*
Unsaturated Solution	*A Solution which can dissolve further amount of solute at a particular temperature is called unsaturated solution.*

Unshared Pair	A pair of electrons in a molecule or ion that is not shared between atoms.
Valence Electron	An electron in or from the outermost electron-containing shell of an uncombined atom.
Valence Shell	The outermost shell containing electrons in an uncombined atom, or that same shell even when the atom is combined in a compound.
Vapour Pressure	The pressure of the vapor in equilibrium with its liquid (or solid).
Vaporization	A phase change from liquid to gas (vapour).
Voltage	Electric potential.
Voltaic Cell	An apparatus that provides a combination of half reactions that can produce an electric current.
Volume	The extent of space occupied by a sample of matter.
Weak Electrolyte	A compound that is only slightly ionized in aqueous solution.

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Neil Bohr's atomic theory

Ionic Bond

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Solibility and effects.

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hydrogen bonding

1to 6 chapter M.C

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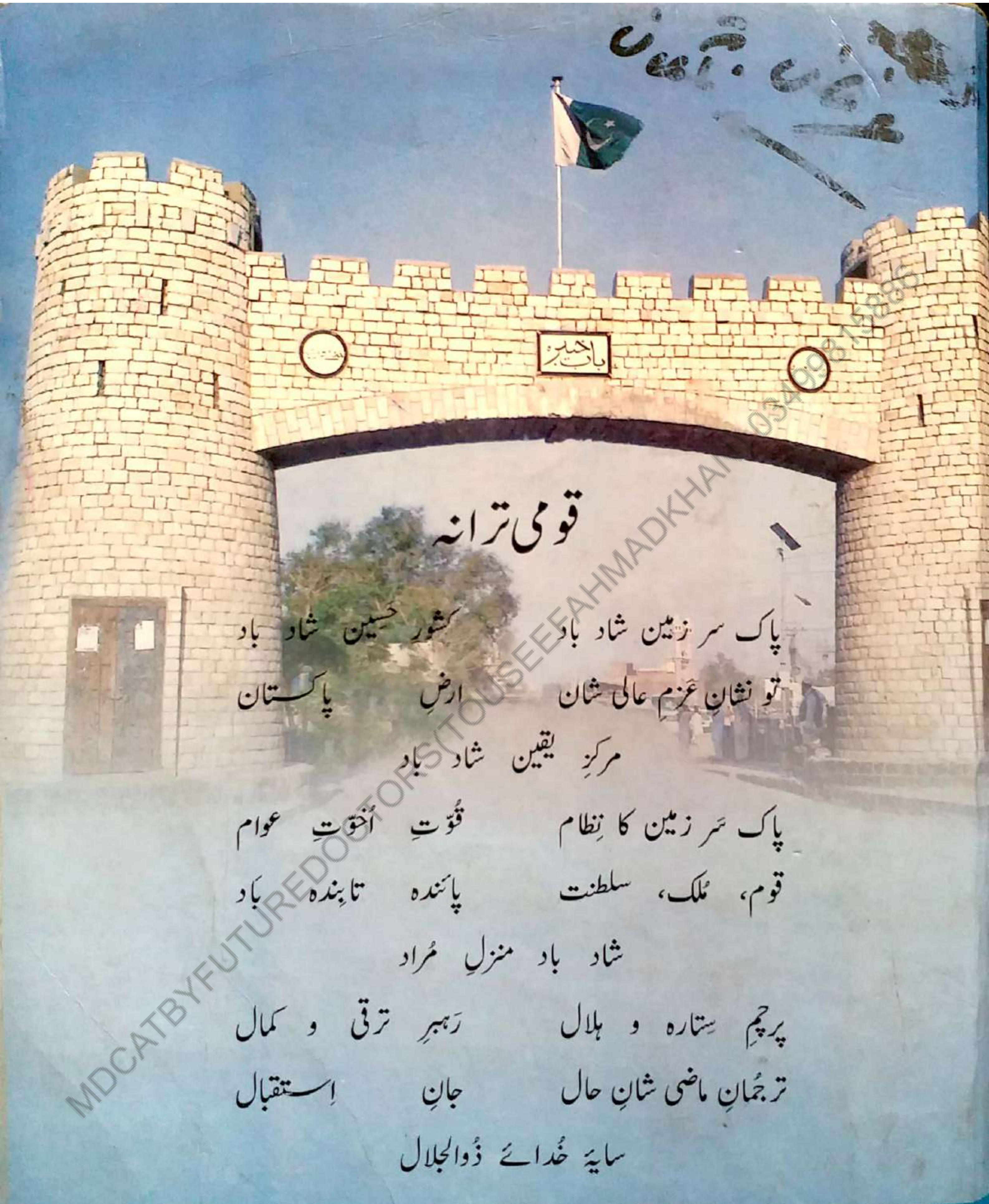
Ruthers ford atomic theory



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