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CHEMISTRY

Class 9th (KPK)

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Chapter # 01

Fundamentals of Chemistry
(TOPIC WISE QUESTIONS)

Q1: Discuss the History of Chemistry?

Ans: History of Chemistry:

The history of chemistry is as old as civilization it grew and flourished in the early civilization of the world. The Egyptians, The Greeks, the Romans and Muslims contributed too much, to the development of chemistry.

Derivation of word Chemistry:

The word chemistry is derived from the word “Kheem” which is the old name of Egypt, due to black colour of its soil. As the time passed the word “Keem” changed into Arabic word “Al-Kimiya” and then changed into English word “Chemistry”.

Purpose of Chemisty:

The purpose of chemistry is to known about the matter, its properties and chemical changes which take place in it. In this regard men kept on learning about many things in the universe.

All these development were improved and achieved by trail and error basis and not on the basis of any systematic study.

Q2: Write a note on the Greek period.

Ans: Greek Period (500 B.C):

The Greek Philosopher were the first to develop ideas related to Chemistry. They introduce the concept of atoms, shape of atoms and chemical combination.

Belief of Greeks:

Greeks believed that all matters were derived from four elements.

i. Earth (Soil)
ii. Fire
iii. Water
iv. Air

According to them one thing or matter could be changed into another if these four elements are used in different properties.

Development of Chemistry in Greek Period:

The theories and the thoughts of the Greek philosopher prevailed upon science for longer time but chemistry could not developed during this period. Because the Greek believed in theoretical ideas not in experiments.

Q3: Write a note on Muslim Period?

Ans: The Muslims Period (600 – 1600 A.D)

The period from 600 to 1600 A.D in A.D in the history of chemistry is known as the Muslims period or Alchemist period.

This period created many talented and genius scientists who observed the matter and conduct experiments to test the observation.

Major aims of Alchemists:

The major aims of al-chemist were:

i. To change base metal into gold.
ii. To discover methods of prolong human life.

iii. To find physical evidence to support religious and philosophical belief.

Contributions:

i. The Muslim scientist discovered metals like arsenic, antimony and bismuth.
ii. They invent and performed different chemical process like sublimation, filtration, calcinations, distillation and fermentation etc.
iii. They also invent different instruments like beaker, funnels, crucibles, furnaces, retorts etc.
iv. They also describe the methods for preparation of chemicals and chemical compounds such as acid like hydrochloric acid (HCl), white lead and alcohols etc.
v. The Muslims also made drugs for various disease.
vi. They also developed methods for the extraction of metals and dying of clothes, leather and varnish making.

Hence in the view of above facts the period of practical chemistry is rightly called the period of Muslims alchemists.

Q4: Discuss the contributions of some prominent Al-Chemists in the development of Chemistry.

Ans: **Contribution of Al-Chemists:**

Muslims scientist (Al-chemist) contributed a lot of knowledge in the field of chemistry. The name and achievements of some of the prominent alchemist are given below.

JABIR IBN-E-HAYYAN (721 – 803 A.D)

Contributions:

1) Jabir ibn-e- Hayyan is generally known as the father of chemistry.
2) He was probably the first scientist who had a well-established chemical laboratory.
3) He invented experimental methods such as distillation, filtration, extraction of metals etc.
4) He prepared Hydrochloric acid, Nitric acid and white lead.

MUHAMMAD IBN-E-ZIKRIYA AL-RAZI (864 – 930 A.D)

Contributions:

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

AL-BERUNI (973 – 1048 A.D)

Contributions:

1) He had a sound knowledge of chemistry, chemical procedures and chemical combinations.
2) He determined the densities of different substances.
3) He also contributed in physics, mathematics, geography and history.

ABU-ALI -IBN-E-SEENA (980-1037 A.D):

Contributions:

1) He is known as the Aristotle of the Muslim World.
2) He is famous for the contribution in the field of medicines, medicinal chemistry, philosophy, mathematics and astronomy.
3) He was the first chemist who rejected the idea that any base metal could be changed into gold.

4) His books were taught in universities of Europe for centuries.

Q4: Define chemistry, state and explain the main branches of chemistry.

Ans: Chemistry:

Chemistry is the branch of science, which deals with the study of composition, structure, properties of matter, the changes occurring in matter and the laws and principles which governs these changes.

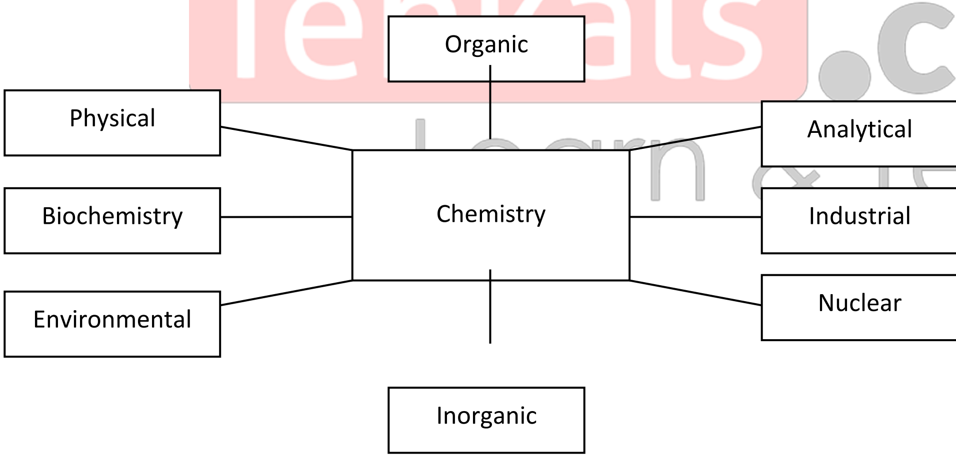
(Or)

The study of matter is called chemistry.

Branches of Chemistry:

Following are the main branches of chemistry:

i. Physical Chemistry
ii. Organic Chemistry
iii. Inorganic Chemistry
iv. Analytical Chemistry
v. Biochemistry
vi. Nuclear Chemistry
vii. Industrial Chemistry
viii. Environmental Chemistry.



i. Physical Chemistry:

The branch of chemistry which deals with the relation b/w physical properties structure, the forces and principles involved in the combination of atoms and molecules.

ii. Organic Chemistry:

The study of carbons and hydrogen containing compound called hydrocarbons and their derivative is called organic chemistry e.g. Methane, Alcohol, Petroleum products etc.

iii. Inorganic Chemistry:

The study of elements and their compound except the hydrocarbon and their derivatives is called inorganic chemistry. E.g. Fe, Cu, Zn, Pb, NaCl, $CaCo_3$ etc.

iv. Analytical Chemistry:

The study of methods and techniques used to determine the kind and quantity of various components in a given substance is called analytical chemistry.

v. Biochemistry:

The study of synthesis and decomposition of compounds and chemical reaction occurring in the living organisms, such as plants, animals and human beings is called biochemistry.

vi. Nuclear Chemistry:

The study of changes occurring in the nuclei of atoms accompanied by the emission or absorption of radiation is called nuclear chemisty.

vii. Industrial Chemistry:

The study of techniques and chemical processes used for the preparation of different industrial products like cement, glass, plastics, fertilizers etc is called industrial chemistry.

viii. Environmental Chemistry:

The study of interaction of chemical substances / processes with environment and their effects on it is called environmental chemistry. Air pollution and water pollution are the two main areas of environmental chemistry.

Q5: Define the following with examples.

(a) Element (b) Compound (c) Mixture

Ans: **Element:**

Element is a pure substance that cannot be broken down into simpler substance by ordinary physical and chemical method. An element is composed of atoms chemically and physically identical in size shape and all other properties.

Therefore element should retain their original properties. There are approximately 118 elements in which 92 are naturally occurring while the rest have been prepared artificially in laboratory.

Symbols of Elements:

In 1884, Berzelius suggested the system for representing elements symbols. The shortest name of an element is called symbol. In most cases, the first letter of name of an element is taken in capital letter as the symbol. In some cases, where the first has been already used, then the initial letters in capital together with a small any other letter is used.

Example:-

Name	**Symbol**
Boron	B
Calcium	Ca
Hydrogen	H
Magnesium	Mg

Some element's symbol starts with their Latin or Greek language name.

Examples:

1	Copper	Cuprum	Cu
2	Gold	Aurum	Au
3	Mercury	Hydragyrum	Hg
4	Tungsten	Wolfram	W

Compound:

Compounds are pure substance which is made up of two or more elements chemically combined together in a definite proportion by mass.

All the compounds which are formed as a result of chemical combination must have completely different physical and chemical properties from the elements. A compound is a pure substance and the components cannot be separated by physical method. Chemical process are necessary to separates its components and the product formed will lose its original shape and properties.

Fixed ratio by mass of a compound is basis component in compound. For example water (H_2O) is a compound. The ratio of hydrogen and Oxygen is always 2:1. Changing this ratio will give a different compound. For example add one more oxygen the ratio becomes 2:2 and the resulting is Hydrogen Peroxide (H_2O_2).

Formula of Compound:

The composition of a compound is 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.

Example:

Water	=	H_2O
Benzene	=	C_6H_6
Sucrose	=	$C_{22}H_{22}O_{22}$

Mixture:

Mixture is an impure substance containing two or more than two elements or compounds physically combined together having no fixed ratio.

A mixture is obtained by mixing two or more elements or compounds in any ratio and the constituent of mixture retain their original properties. The constituent of mixture can easily be separated from each other by various physical methods e.g. filtration, distillation, sublimation, crystallization etc.

Example:

NaCL in water

Iron in Sulphur

Types of Mixture:

There are two types of mixture.

1. Homogenous mixture
2. Heterogeneous mixture

Homogenous Mixture:

Homo means "same" and generous mean "form" so mixture having uniform composition, throughout their mass is called homogenous mixture e.g. Air, Salt in water, Sugar in water. Etc.

Heterogeneous Mixture:

Hetero means "different" and generous means "form" so mixture having different and visible composition and the component can be seen with naked eyes is called heterogeneous mixture e.g. Ice-cream, cooking meal, muddy water etc.

Q6: Write note on the following.

a. Relative atomic mass b. Atomic mass unit c. Average atomic mass

Ans: a. Relative atomic mass:

Definition: The relative atomic masses of an element is the mass of an atom of an element relative to the mass 1/12th mass of carbon-12.

Explanation:

Atom is extremely the smallest particle of matter. It is impossible and find out the accurate and exact mass of an atom of element by using a very sensitive or accurate balance. It is therefore preferable to measure the atomic masses of an atom by comparing masses of an atom with the mass of standard atom. The standard chosen for this purpose is the lightest isotope of carbon, which has a mass exactly 12 a.m.u.

Example:

One atom of H = 1.008 a.m.u

One atom of He = 4.00 a.m.u

One atom of Na = 23 a.m.u

b. Atomic mass unit:

The mass equal to 1/12th of the mass of one carbon 12 atom is called atomic mass unit. The mass of 1 atom of C_{12} = 12 a.m.u.

1 mol of C_{12} = 12g

1 mol of $C_{12} = 6.023 \times 10^{23}$

1 mol of $C_{12} = 12g = 6.023 \times 10^{23}$

1/12 x mass of one atom of C_{12} taken exactly as 12 = 1 a.m.u

1 amu = 1/12 g mol / 6.023×10^{23} mol

1 amu = $1/12 \times 1.99 \times 10^{23} = 1.66 \times 10^{24}$ g = 1.66×10^{-27} kg

The mass of one proton or one neutron is equal to one amu.

c. Average Atomic mass:

Average atomic mass is the weighted average of the atomic masses of the naturally occurring of the isotopes of an element.

The atomic masses 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.

$$\text{Average atomic mass} = \frac{\text{atomic mass of 1st isotope x its \% abundace}}{100} + \frac{\text{atomic mass of 2nd isotope its \%age abundace}}{100}$$

Q7: Define (1) Formula unit (2) chemical species (3) ions (4) molecular ions (5) free radical.

Ans: **1. Formula Unit:**

The simplest ratio between the ions of an ionic compound which are present in giant structure is
called formula unit. OR

Formula unit is the smallest repeating unit of an ionic compound showing the simple ratio between the ions.

For example. The simplest relation b/w Na and Cl ions in the whole crystal lattice of NaCl is 1:1 so the formula unit of sodium chloride is NaCl. Similarity KCl is the formula unit of potassium Chloride.

2. Chemical Species:

An atom or group of atoms which can take part in a chemical reaction is called chemical species. A chemical species may be neutral or it may carry a charge. Such chemical species are classified into ions, free radicals and molecular ions.

3. Ions:

Electrically charged particles are ions. (OR)

The particles that carries a positive or negative charge by the loss or gain of electrons is also called ions.

There are two types of ions.

i. Cation:

The positively charged ions that are formed by the gain of electrons is called cation. Positive ion always have less number of electrons than number of protons.

Examples:-

Na^+, K^+, Ca^{2+}, Mg^{2+} etc.

ii. Anion:

The negatively charged ions that are formed by the gain of electron is called anion. Negative ions have always have more number of electrons than number of protons.

Example:-

F^-, Cl^-, O2, etc.

4. Molecular Ions:

Electrically charged molecules formed by the lost or gain of electrons is called molecular ions. Positive molecular ions are formed by the loss of electrons from neutral molecules and negative molecular ions are formed by the gain of electrons from neutral molecules. Positive molecular ions are called molecular cation.

Example:-

O_2^+, CO^+, CH_4^+, $C_2H_5^+$, NH_4^+, H_3O^+ etc.

Negative molecular ions are called molecular anions.

Examples:-

SO_4^{-2}, CH_3COO^-, OH^-, CO_3^{-3} etc.

5. Free Radical:

An atom or molecule having single (an unpaired) electron in the outer shell with no charge is called a free radical.

Explanation:-

Free radicals are highly reactive species formed by the bond breaking (hemolytic fission) of stable molecules in such a manner that the resulting reactive specie get separated with unpaired electron. A free radical has no change and are represented by dot(.) which is written on the upper side of an atom or molecule. A free radical is reactive specie which does not exist independently E.g H, Cl, CH_3.

Example:- During the reaction b/w chlorine molecule (Cl_2) and methane (CH_4) in the presence of diffused sunlight the chlorine molecules first form chlorine free radical which then ultimately result in a chain reactions.

Cl_2 sunlight 2CL (Chlorine free radical)

The chlorine free radical (Cl) react with CH_4 to form methyl free radical

CH_4 + Cl ➔ CH_3 + (Methyl free radical) + cl

The CH_3 react with another Cl_2 molecule forms chloromethane and Cl

$CH_3 + Cl_2$ $CH_3 - CL + CL$

Q8: What is difference between an atom and ion?

Ans: Difference between an atom and ion:

ATOM	ION
An atom is the smallest particle of an element that can take part in a Chemical reaction.	The particle that carries a positive or negative charge by the loss or gain of electrons is also called ions.
It is a 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.
It is the smallest particle of an element.	It is the smallest unit of ionic compound.
It can or cannot exist independently.	It cannot exist independently.
Example He, Na, Fe, Cl	Examples Na^+, Fe^{+2}, Cl

Q9: Define molecule and there types:

Ans: **Molecule:**

The smallest particle of an element or compound which can exist independently and do not take part in chemical reaction is called molecules.

A molecule may be Mono atomic and poly atomic.

Mono Atomic Molecules:

(Mono=one) This type of molecule is made up of only one atom, Examples of such molecules are the molecules of noble gases such as He, Ne, Ar, Kr, Xe, and Rn.

Polyatomic molecules (Poly=many

The molecules made up of more than one atom are termed as poly atomic molecules. This may be diatomic which is made up of two atoms, triatomic made up of three atoms and tetra atomic made up of four atoms.

Examples:-

Di atomic = CO_2, CO

Tri atomic = CO_2, H_2O

Tetra atomic = NH_3

Penta Atomic = CH_4

Q10: Define gram atomic mass, gram molecular mass and gram formula mass of the element

and compounds give at least two examples in each case.

Ans: **Gram Atomic Mass:**

When atomic mass of an atom of element expressed in gram is called gram atomic mass. It is also called gram atomic mass.

Examples:

Gram atomic mass of H atom = 1.008 gram

Gram atomic mass of O atom = 16 gram
Gram atomic mass of C atom = 12 gram

Gram Molecular Mass:

When molecular mass of molecules of covalent compounds are expressed in grams is called gram molecular mass. It is also called gram molecules.

Examples:

Gram molecular mass of H_2O = 2 x 1 + 1 x 16 = 18 gram
Gram molecular mass of CO_2 = 1 x 12 + 16 x 2 = 44 gram
Gram molecular mass of CH_4 = 1 x 12 + 1 x 4 = 16 gram

Gram formula Mass:

When molecular mass of all the ionic compounds of all the ions present in a formula unit expressed in gram is called gram formula mass. It is also called gram formula.

Examples:

Gram formula mass of Nacl = 23 + 35.5 = 58.5 gram
Gram formula mass of $CaCl_2$ = 40 + 2 x 35.5 = 111 gram
Gram formula mass of KCL = 39 + 35.5 = 74.5 gram

Chapter # 01

Fundamentals of Chemistry

(LONG QUESTION ANSWER EXERCISE)

Q1. State and explain with examples.

a. The empirical formula of compound

b. The molecular formula of compound

Ans: **(a) The empirical formula of compound:**

The formula which shows the simplest ratio b/w the atoms of different elements present in one molecules of a compound is called empirical formula. It is also called simple formula.

Explanation:

Empirical formula does not tell us about actual numbers of atoms present in the compound.
E.g. Benzene has a formula C_6H_6 the simplest ratio b/w carbon and hydrogen is 1:1. Therefore the empirical formula of benzene is CH. Some other examples of empirical formula are as follows:

Name	FORMULA	EMPIRICAL FORMULA
Glucose	$C_6H_{12}O_6$	CH_2O
Acetic Acid	$C_2H_4O_2$	CH_2O
Hydrogen peroxide	H_2O_2	HO
Acetylene	C_2H_2	CH

Sometime different compounds have some empirical formula e.g. Benzene (C_6H_6) and acetylene (C_2H_2) have same empirical formula CH. Similarly all the ionic compounds are also represented by their empirical formula which shows the simplest ratio between them.
e.g. NaCl contain NA^+ and Cl^+ ion in a ratio of 1:1 $CaCl_2$ have a ratio 1:2.

(b) Molecular Formula:

The formula which shows the actual and exact number of atoms of different elements present in one molecule of a compound is called molecular formula e.g. molecular formula of a glucose is $C_6H_{12}O_6$, acetic acid is CH_3COOH.
Molecular formula of a compound can be find out by using the following formula.
Molecular formula = n x empirical formula

Where, $n = \frac{molecular\ mass}{emperical\ formula\ mass}$

Q2. What do you understand by the terms mole and Avogadro's number. Give examples.

Ans: **Mole:**

Atomic mass, molecular mass or formula mass of a substance expressed in grams is called the mole. (OR)
The quantity of a substance containing Avogadro's number of particles (atoms, ions or molecules) is called mole.
It is the basic SI unit of quantity of matter.

Examples:-

1 mole of NaCl = 58.5 gm
1 mole of Na = 23 gm
1 mole of O = 16 gm
1 mole of O_2 = 32 gm

Relation b/w number of moles and amount of substance:

If mass in grams of a substance and molecular mass is given, we can calculate the number of moles by the following formula:

Moles = $\frac{Mass\ in\ gm}{Atom\ /\ Molecular\ mass}$

Avogadro's number (N_A):

The number of particles (atoms, ions or molecules) present in one mole of a substance is called Avogadro's number. (OR)
The number of atoms, ions or molecules which correspond to atomic mass, molecular mass or formula mass of a substance is also known as Avogadro's number. It is a constant value and is equal to 6.023×10^{23}. This number was determined by an Italian Scientist, Amado Avogadro and its called Avogadro's number represented by N_A.

Examples:-

1 mole of H = 1.008 gm = 6.023×10^{23} atoms
1mole of Na = 23 gm = 6.023×10^{23} atoms
1 mole of H_2O = 18gm = 6.023×10^{23} molecules
1 mole of CO_2 = 44gm = 6.023×10^{23} molecules

Mathematically:

$N_A = \frac{No.of\ particles}{No.of\ moles}$

Q3(a) Compare and contrast a mixture and compound. Give examples of each of them.
(b) How will you classify molecules? Support your answer with at least two example of each.

Ans(a)

COMPOUND	MIXTURE
It is formed by chemical combination of atoms.	It is formed by physical combination of atoms.
The constituents lose their original properties.	The constituents retain their original properties.
Compounds always had fixed composition.	Mixture does not have fixed composition by mass.
The components of the compound cannot be separated by physical methods.	The component of the mixture can be separated by physical methods.
Every compound is represented by its chemical formula.	It consist of two or more components and does not have a chemical formula.
Compound have homogenous composition.	Mixture may be homogenous or heterogeneous in composition.
Compound has sharp and fixed melting point.	Mixture does not have sharp and fixed melting point.
Examples Sodium Chloride, Ethyl Alcohol, Hydrochloric acid, Distilled water	Examples Air, Rock, Ice Cream Muddy water, Mineral water, Solution

<u>(b): Molecules:</u>

The smallest particle of matter which can exist free in nature.

A molecule is formed by the chemical combination of atoms. It is the smallest unit of substance. It may be composed of like or unlike atoms. It show all the properties of that particle substances.

Example: H_2O_2, H_2O etc.

<u>Types of molecules:</u>

<u>1. On the basis on number of atoms:</u>

i. Monoatomic Molecule: (Mono = One)

Monoatomic molecule are those molecules which are made of only one atom.

For example, the inert gases He, Ne, Ar etc.

ii. Diatomic Molecule: (di = two)

Diatomic molecules are those molecules which are made of two atoms.

For example: H_2, O_2, HCl etc.

iii. Triatomic Molecule: (tri = three)

Triatomic molecule are those molecules which are made of three atoms.

Fro example: H2o, CO_2, O_3 etc.

iv. Polyatomic Molecule: (poly = many)

Polyatomic molecule are those molecules which are made of three atoms.

For example: H_2SO_4, $C_{12}H_{22}O_{11}$, S_8 etc.

2. On the basis of type of atoms:

i. Homo-atomic Molecule: (Homo means same)

When a molecules which consists of same atoms of the element, it is called homo-atomic molecules. They are also called homo-nuclear molecule. The homo atomic molecules are diatomic and triatomic in nature:

<u>Example:-</u>

H_2, N_2, Fe_2, Cl_2, O_3, etc.

ii. Hetero-atomic Molecules: (Hetero means different)

When a molecules which consists of atoms of the different elements, it is called heteratomic molecules. They are also called hetero-nuclear molecule. The hetero atomic molecules may be triatomic or polyatomic in nature.

<u>Example:-</u> H_2O, CO_2, HNO_3, H_2SO_4, etc.

Q4. (a) What is molecular mass of a compound? How will you differentiate it from formula mass?

Ans. (a) Molecular Mass: The sum of the atomic masses of all the atoms of an element present in one molecule is called molecular mass. (OR)

The mass of a molecule of a compound relative to the mass of lightest isotopes of carbon taken as 12 a.m.u is also called molecular mass.

<u>Example:</u>

Molecular mass of O = 32 + 16 x 2 = 64 a.m.u

Molecular mass of CO_2 = 12 + 1 x 16 = 44 a.m.u

Molecular mass of H_2S = 1 x 2 + 32 = 34 a.m.u

<u>Formula Mass:</u> The sum of the masses of all the ions present in a formula unit of an ionic compound is called formula mass. (OR)

The mass of a formula unit of an ionic compound relative to the mass of lightest isotopes of carbon taken as 12 a.m.u is also called formula mass.

<u>Example</u>

Formula mass of NaCl = 23 + 35.5 = 58.5 a.m.u

Formula mass of $CaCl_2$ = 40 + 35.5 x 2 = 111 a.m.u

(b) Calculate the molecular mass or formula mass, as the case may be the following compounds in a.m.u.

Ans. **i. Benzene (C_6H_6):**

C = 6 x 12 = 72

$H = 6\,x\,1 = 6$

$78\,a.m.u$

ii. Ethane (C_2H_6):

C = 2 x 12 = 24

$H = 6\,x\,1 = 6$

$30\,a.m.u$

iii. Aluminium Chloride ($AlCl_3$):

Al = 1 x 27 = 27

Cl = 3 x 35.5 = 106.5

133.5 a.m.u

iv. Iron Oxide (Fe_2O_3):

Fe = 2 x 56 = 112

O = 3 x 16 = 48

160 a.m.u

Q5. (a) Find out the number of proton, electron and neutrons in the following elements.

Ag, Na, Fe, Ar, Pb, U

Ans: (a) As No. of P = No. of e.

No. of neutrons = A – Z

Element	No of e	No of P	Non of n^0 = A – Z
Ag	47	47	107 – 47 = 60
Na	11	11	22
Fe	26	26	30
Ar	18	18	22
Pb	82	82	125
O	92	92	146

Q5. (b) Complete the following table.

Ans. (b)

	Symbol	Atomic No.	Number of Protons	No. of electrons
a.	K	19	19	19
b.	O	8	8	8
c.	P	15	15	15
d.	Ca	20	20	20
f.	Cl	17	17	17

Chapter # 01

Fundamentals of Chemistry

(SHORT QUESTION ANSWER EXCERCISE)

Q1: How many electron are present in each of the following?

a. HF and Hf b. Co and CO

c. Si and SiO_2 d. $PoCl_2$ and $POCl_3$

Ans: a. HF = two elements (Hydrogen and Flourine). Hf = one element (Hafnium)

b. Co = one element (Cobalt) CO = Two elements (Carbon and Oxygen)

c. Si = One element (Silicon)

SiO_2 = Two element (one atom of Silicon and two atoms of Oxygen).

d. $PoCl_2$ = Two elements (one atom of Polonium and two atoms of Chlorine)

$POCl_3$ = Three elements (one atom of Phosphorus, one atom of oxygen and three atoms of chlorine.

Q2: Cm is the chemical symbol for Curium, named after the famous scientist Madam Curie.

Why Wasn't the symbol C, Cu or Cr used?

Ans: Curium is the radioactive element named after Madam Curie was discovered by T Glen Seaborg

in 1945. Its symbol is Cm. Its atomic number is 96 and present in actinide series in periodic table.

There are two reasons for using Cm symbol for curium istead of C, Cu or Cr.

Reason i:

These symbols were already used i.e. C for Carbon, Cu or Copper and Cr for Chromium.

Reason ii:

These element were discovered before Curium.

Q3: What is atomic number? How of an element does it differ from mass number?

Ans: Differences between atomic number and mass number:

Atomic Number	**Mass Number**
The total number of protons present in the nucleus of an atom is called atomic number. OR The number of electrons present in various shells of an atom is called atomic number.	The sum of protons and neutrons present in nucleus of an atom is called mass number.
It is also known as charge number.	It is also known as nucleon number.
It is represented by "Z".	It is represented by "A"
Atomic number = No. of Protons or number of Electrons.	Mass Number = No. of Protons + No. of Neutrons.
Example	Examples

Hydrogen = Z = 1 Carbon = Z = 6 Oxygen = Z = 8	Hydrogen = A = 1 Carbon = A = 12 Oxygen = A = 16

Q4: Student 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

Ans:
a. Mg = Magnesium Mn = Manganese
b. K = Pottasium P = Phosphorus
c. Na = Sodium S = Sulphur
d. Cu = Copper Co = Cobalt

Q5 .a. Classify the following molecules as monoatomic, diatomic, triatomic and polyatomic molecules. N_2O, N_2, S_8, He, HCl, CO_2, Ar, H2, H_2SO_4, $C_6H_{12}O_6$.

Ans:

Monoatomic Molecule	Diatomic Molecule	Triatomic Molecule	Poly atomic Molecule
He Ar	N_2 HCl	H_2O CO_2	S_8 H_2SO_4 $C_6H_{12}O_6$

Q5 .b. Classify the following as cation, anion, molecular ion, free radical and molecule: CH^+, O^{-2}, CH_3, CO+, CO_2, Cl-, Mg^{+2}, $Co3^{-2}$, O_2, Na+, $C_2H_5O^{-1}$, H_2O, Cl_2.

Ans:

Cation	Anion	Molecular ion	Free Radical	Molecule
Mg^{+2} Na^+	O^{-2} Cl-	$C_2H_5O^{-1}$ CH^+ CO^+ CO_3^{-2}	CH_3	CO_2 O_2 H_2O Cl_2

Q6. Calculate the number of moles of butane, C4H10 in 151g of butane (Atomic masses of
C = 12 amu and H=1 amu).

Ans: <u>Given Data:</u>

Mass in grams = 151g

Molecular mass of butane = C_4H_{10} = (12 x 4) + (1x10) = 48 + 10 = 58 amu

<u>Required Data:</u>

No. of moles = ?

<u>According to formula:</u>

No. of moles = $\frac{mass\ ingrams}{molecular\ mass}$

$= \frac{151}{58}$

No. of moles = 2.63 mol

Q7. What is the mass of 5 moles of ice? (Atomic masses of H = 1 amu and O = 16 amu)

Ans: Given Data:

No. of moles = 5 mol

Molecular mass of ice (H_2O) = (2 x 1) + (1 x 16) = 2 + 16 = 18 amu

Required Data:

Mass in grams = ?

According to formula:

$$\text{No. of Moles} = \frac{mass\ in\ grams}{molecular\ mass}$$

Rearranging the Formula:

Mass in grams = No. of moles x molecular mass

5 mol x 18 amu

Mass in grams = 90g.

Q8: Calculate the number of molecules in 6.50 mol of CH_4 (Methane).

Ans: Given Data:

No. of moles = 6.5 mol

Avogadro's number = N_A = 6.023×10^{23}

According to formula:

$$\text{No. of moles} = \frac{No. of\ molecules}{NA}$$

Rearranging the Formula:

No. of molecules = No. of moles x N_A

$= 6.50 \times 6.023 \times 10^{23}$

$= 39.14 \times 6.023 \times 10^{23}$

$= 3.914 \times 10^{23+1}$

No. of molecules = 3.914×10^{24} molecules

Q9. Calculate the average atomic mass of Lithium for following data.

Isotopes	Natural abundance	Relative atomic masses
^6Li	7.5%	6.0151
^7Li	92.5%	7.0160

Ans: Given Data:

Relative atomic mass of ^5Li = 6.0151

Natural abundance of ^6Li = 7.5%

Relative atomic mass of ^7Li = 7.0160

Natural abundance of ^7Li = 92.5%

Required Data

Average Atomic mass = ?

According to formula:

$$\text{Average atomic mass} = \frac{(R.At\ Mass\ 6Li\ x\ \%age) + (R.At\ Mass\ 7\ Li \times \%age}{100}$$

$= \frac{(6.0151 \times 7.51 + (7.0160 \times 92.5)}{100}$

$= \frac{(45.38) + (648.98)}{100}$

$= \frac{694.36}{100}$

= Average atomic mass of Lithium = 6.94 amu

Q10. Calculate the mass of 6.68 x 10^{23} molecule of PCL_3.

Ans: Given Data:

No. of molecules = 6.68 x 10^{23}

Avogadro's number = N_A = 6.023 x 10^{23}

Molecular mass of PCl_3 = (1x30.97) + (3x35.5)

= 30.97 + 106.5 = 137.47 amu

Required Data:

Mass in grams = ?

First calculate no. of moles by following formular

No. of moles = $\frac{No.of\ molecules}{N_A}$ = 6.68 x 10^{23} / 6.023 x 10^{23} No. of moles = 1.10 mol

Now calculate mass in grams by following formula: No. of moles = $\frac{mass\ in\ grams}{molecular\ mass}$

Rearranging the formula:

Mass in grams = No. of moles x molecular mass

1.109 x 137.47

Mass in grams = 151.62 g.

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Chapter # 02

Structure of Atom
(TOPIC WISE QUESTIONS)

Q1: What are fundamental particles of an atom?

Ans: Fundamental particles of an atom:

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

These particles are called fundamental particles.

a. Electron:

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

b. Proton:

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

c. Neutron:

Neutron is a neutral particle because it has no charge. Its mass is equal to 1.0086654 amu or 1.6749×10^{-27} kg. Neutron is 1842 times heavier than an electron.

Neutrons are present in the nucleus of an atom.

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}

Q2. What is Isotope? Explain by examples.

Ans: **Isotopes:**

Atoms of the same elements having same atomic number but different atomic masses are called isotopes.

Explanation:

The word isotope was first suggested by Soddy scientist since they were occupying the same place in Periodic Table.

In Greek Language "Isos" mean same and "topes" mean place.

In Dutton atomic theory all the atoms of an elements were considered identical but later it was proved that the number of protons in the atoms of an elements remain the same while neutrons number may different therefore, different isotopes will show same chemical properties and their physical properties show variation depends upon the number of Neutrons present in the Nucleus.

Example of Isotopes:

Isotopes of Carbon

Carbon has 3 isotopes, carbon – 12, carob – 13 and carbon – 14

$^{12}C_6$ $^{13}C_6$ C_6

Carbon has atomic number = 6

Carbon – 12 = $^{12}C_6$

Atomic No. = 6

Mass No = 12

No. of electrons = 6

No. of protons = 6

No. of neutrons = 12 – 6 = 6

Carbon – 13 = $^{13}C_6$

Atomic No = 6

Mass No = 13

No. of electrons = 6

No. of protons = 6

No. of neutrons = 13 – 6 = 7

Caron – 14 = $^{14}C_6$

Atomic No = 6

Mass No = 14

No. of electrons = 6

No. of protons = 6

No. of neutrons = 14 – 6 = 8

All the isotopes have same number of electrons and protons

But different number of neutrons.

ii. Isotopes of Chlorine:

Chlorine exist in two isotopes $^{35}Cl_{17}$ and $^{37}Cl_{17}$

The natural abundance of Cl-25 is 75.53% and that of Cl-37 is 24.47%

Chlorine – 35 = $C_{17}^{35}L$

Atomic Number = 17

Mass No = 35

No. of electrons = 17

No. of protons = 17

No. of neutrons = 35 – 17 = 18

Chlorine – 37 = $C_{17}^{37}L$

Atomic Number = 17

Mass No = 37

No. of electrons = 17

No. of protons = 17

No. of neutrons = 37 – 17 = 20

iii. Isotopes of Uranium:

Uranium exist in three isotopes, U_{92}^{234}, U_{92}^{235} , U_{92}^{233}

The percentage composition of U – 234 is 0.005%

The percentage composition of U – 235 is 0.75%

The percentage composition of U – 238 is 99.245%

Uranium – 234 = U_{92}^{234}

Atomic No = 92

Mass No = 234

No. of electrons = 92

No. of protons = 92

No. of neutrons = 234 – 92 = 142

Uranium – 235 = U_{92}^{235}

Atomic no = 92

Mass no = 235

No. of electrons = 92

No. of protons = 92

No. of neutrons = 235 – 92 = 143

Uranium – 238 = U_{92}^{238}

Atomic no = 92

Mass no = 238

No. of electrons = 92

No. of protons = 92

No. of neutrons = 238 – 92 = 146

Note:

Elements of odd atomic number mostly do not more than two stable isotopes. Elements of even atomic number usually contain large number of isotopes.

Q3. What do you mean by the term electronic configuration?

Ans: **Electronic Configuration:**

According to the Bohr's atomic model the arrangement of electrons around the nucleus in various shells and sub-shells is called electronic configuration.

According to the Bohr's atomic model the electrons revolve around the nucleus in different shells orbits. These shells are named as K, L, M, N etc.

Maximum number of electrons in a shell is determined by using $2n^2$ formula where "n" is the number of shell n = 1, 2, 3, 4…….

For example, for the K –shell n = 1 the number of electrons in K-shell is $K=2(1)^2 = 2e$

$L=Shell = (2)^2=8e$

Modern research has shown that the shell is further divided into sub-shells which are s, p, d, f the number of sub-shells in each shell and the number of electrons in each sub-shell are given in the table.

N	Shells	Sub-Shells	No of e	Total no of e
1	K	S	2	2e
2	L	S, P	2 + 6	8e
3	M	S, P, d	2 + 6 +10	18e
4	N	S, P, d, f	2 + 6 + 10 + 14	32e

Q4. Explain the uses of Isotopes?

Ans: **Uses of Isotopes:**

Isotopes are used in chemical, agriculture, 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:

Goiter treatment

i. Iodine – 131 become concentrated in the thyroid gland and is used as cure for goiter.

. **Brain imaging**

ii. Iodine – 123 is used for brain imaging:

Tracer studies:

iii. The heavy hydrogen (deuterium), the heavy carbon (C-13), the heavy nitrogen (N-15) and heavy oxygen (O-18) and Iodine – 131 are used as tracer elements in biochemical and physiochemical research to trace the path of the element to the defective or obstructed part.

Treatment of cancer:

iv. Radio irradiation and cobalt -60 are used in the treatment of cancer and for the diagnosis of tumors

v. Sodium (Na-24) is used for the identification of blood circulatory problems in patients.

vi. Carbon-14 is used to trace the path of carbon in photosynthesis.

. **Smoke detector**:

vii. Americium – 241 is used in smoke detectors. It is also used to determine where oil wells should be drilled.

Finding moisture content

:viii. Californium – 252 is used measure moisture content of soil in road construction. It is also used to inspect airline luggage for hidden explosive.

Use in electrical appliances:

ix. Krypton – 85 is used in clothes washers to measure dust and pollutants levels.

Q5. There are three isotopes of uranium having atomic number 92 and mass number 234, 235 & 238. Calculate the number of neutrons in their nuclei.

i. U-234

A = 234, Z = 92

No. of protons = No. of electron

Number of neutrons = A – Z

234 – 92 = 142

ii. U-235

Number of neutrons A – Z

235 – 92

143 Neutrons

iii. U-238

Number of neutrons = A – Z

283 – 92

146 Neutrons.

Structure of Atom
(LONG QUESTIONS)

Q1: Why Dalton's atomic theory is considered as a base of modern atomic concepts?

Ans: The word 'atom' comes from Atomos which means uncut indivisible or the smallest particle Which are impossible to see with naked eyes. It was an old theory that matter is made of very small particles.

This idea was first proposed by Greek Philosopher Democritus in 400 BC, 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 solid balls and that cannot be further sub-divided. He called them atoms. He presented his theory under the title "A New System of Chemical Philosophy".

The main points of Dalton's atomic theory are as follow:

- Matter is composed of smallest tiny indivisible particles called atoms.
- Atom can neither be created nor destroyed.
- Atoms of the same element are identical in size, shape, mass and their properties.
- Atom of different elements is different in their properties.
- Atom combine together in small whole number and in simple ratio to form compounds.
- All chemical reactions are due to combination or separation of atoms.

Q2: Summarize Rutherford's atomic model of an atom and explain how we developed this. Model based on result of his famous gold-foil experiment

Ans: **<u>Rutherford Atomic Model:</u>**

In 1911 Lord Rutherford performed an experiment a-particle (20,000) which carries positive charge and in fact helium nuclei from a radioactive source (polonium metal). He allowed to fall, a beam of α-particle on a thin gold foil (0.00004 cm). The gold foil was surrounded by photographic plate or zinc sulphide (ZnS) fluorescent screen to detect the particles emitting from the radiation.

<u>Observation:</u>

Rutherford observed that most of the α-particles (19990) passed through the foil undeflected or without changing their path but a few particles (8) were deflected at different angels. Only few rays (2) were bounced back at their original way. From the deflection of α-particles bounced back at the same angle.

<u>Conclusion:</u>

Rutherford concluded that there was a positively charged particles present in the center of atom. So α-particles near this portion were repelled. Because α-particles are also positively charged particles and similar charges repel each other. If α-particles pass very closely to nucleus, they deflected through large angles. Similarly, if do no pass close to nucleus they either deflected through very small angles or do not get deflected at all.

<u>Main Points:</u>

i. The positive charge present in the center of an atom called nucleus. It contains electrons and neutrons.
ii. Majority of α-particle passed without changing their path shows that most of the spaces in atom are empty.
iii. The electrons revolving around the nucleus would require centripetal force. The attractive force of the nucleus on electrons provides centripetal force to the electron.
iv. The size of nucleus is so small as compared to the size of an atom.

v. The whole mass of an atom is present in its center called Nucleon mass due to the presence of protons and neutrons.
vi. The negative charged electrons revolved around the nucleus at a very small distance.
vii. An atom is neutral. As the number of electrons is numerically equal to that of protons.
viii. Nucleus is responsible for mass and energy of the atoms.

Defects in Rutherford's Atomic Model:

The major objections raised against his model were the following.

i. Rutherford's model is based on the laws of motion and gravitation, which are applicable to neutral bodies and not on the charged bodies.
ii. According to Maxwell theory, the revolving electrons being a charged particle, must lose energy continuously and ultimately spiral (fall) into the nucleus. However, it does not happen.
iii. The revolving electron radiates energy continuously and the atomic spectrum should be a continuous one but actually it gives a line spectrum.
iv. It does not provide any explanation about the chemical properties of the elements.

Q3: State the postulates which Bohr suggested to overcome the short comings of Rutherford's atomic Model?

Ans: **Neil Bohr's Atomic Theory:**

To overcome the defects of Rutherford's atomic model. Neil Bohr in 1913, presented an atomic theory. Considering Hydrogen atom as a model, the theory is based on the following assumptions.

i. The negative charged electrons revolve around the positively charged nucleus in certain fixed circular paths called shells, orbits or energy levels.
ii. The energy of the electron in orbit is proportional to its distance from the nucleus. The further the electron from the nucleus, the higher will be the energy and vice versa.
iii. Electron does not radiate energy as long as it is present in an orbit i.e. energy of an orbit is fixed.
iv. The electron absorbs or radiates energy whenever it moves from one orbit to another. The energy change of electron on going from one orbit to another is given by the relationship.

$\Delta E = E_2 - E_z = hv$

Where

hv = plants constant

v = frequency of radiation.

E_2 = the lower energy orbit

E_2= the higher energy orbit

ΔE = the energy difference

v. Electron can reside in the orbit for which its angular momentum (mvr) is integral multiple of n/2π i.e. mvr = $\left[\frac{nh}{2\pi}\right]$. Where n is the number of shells i.e. 1,2,3….., m is the mass, v is the velocity of an electron, r is radius of the orbit and h is plank constant (6.6262 x 10^{-34}js).
vi. Electron can reside in any one of the orbits and cannot stay in between them.

Q4: Complete the following table for neutral atoms of specific isotopes:

Ans:

	Isotopic symbol	Atomic number	Mass number	No. of Electron	No. of protons	No. of Neutron ($N^0 = A - Z$)
A	Xe_{54}^{131}	54	131	54	54	131 -54 = 77
B	Co_{27}^{59}	27	59	27	27	59-27=32
C	Nd_{60}^{144}	60	144	60	60	144-60=84
D	Ti_{22}^{48}	22	48	22	22	48-22=26
E	Hf_{72}^{178}	72	178	72	72	178-72=106
F	Te_{52}^{128}	52	128	52	52	128-52=76
G	Ar_{18}^{40}	18	40	18	18	40-18=22

Q5. (a)Define energy level and sub energy level.

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

Ans. **(a)** Energy Levels:

These are definite circular path at a definite path at the definite distance from the nucleus in which the electrons moves in anti-clock wise direction or any direction. The energy levels are also called shells or obits. The number of electrons in an orbit is constant according to$2n^2$formula presented by Bohr's. These orbits are designed as K, L, M, N etc.

Shells	No. of e – (2n2)
1 = K	$2(1)^2 = 2e^-$
2 = L	$2(2)^2 = 8e^-$
3 = M	$2(3)^2 = 18e^-$
4 = N	$2(4)^2 = 32e^-$

Sub-Energy Levels:

The various regions in the main shells around the nucleus in three dimensional direction where the possibility of finding electrons is maximum is called Sub-Energy Level.

Group of orbitals around the nucleus having same energy is also called sub-energy levels. Sub energy levels are also called orbitals. These orbitals cannot accumulate more than 2e. There are four types of orbitals namely as s, p, d and f, which stand for sharp, principle diffused and fundamentals respectively. The s-orbital is spherical P orbital is dumbbells while d orbital is double dumbbell and f are more complex in shape.

(b)

i. H = 1

K = 1, $1S^1$

Group = 1A, Period = 1

ii. Li = 3

K = 2, L = 1

$1S^2$, $2S^1$

Group = 1A, Period = 2

iii. Na = 11

K = 2, L = 8, M = 1

$1S^2$, $2S^2$, $2p^6$, $3s^1$

Group 1A, Period = 3

iv. K = 19

K = 2, L = 8, M = 8, N = 1

$1s^2, 2s^2, 2p^5, 3s^2, 3p^6, 4s^1$

Group = 1A, Period = 4

Structure of Atom
(SHORT QUESTIONS)

Q1: Aluminum is represented as Al_{13}^{27}. Draw the structure of Aluminum. Write its electronic configuration.

Ans: **Electronic configuration of Aluminum:**

Atomic number of Aluminums is thirteen (13). K=2 es, L=8 es, M=3 es. Its electronic configuration is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^1$.

Q2. The energy of an electron in K and L shells is same or different. Explain.

Ans: The energy of an electron in K and L shells is different. Because according to Neil Bohr Atomic Model the energy of an electron in orbit is directly proportional to its distance from the nucleus. The farther the electrons of L shell are comparatively farther than the electrons of K shell. So L Shell's electron will have higher energy from K shell's electrons.

Q3. Draw the structure of hydrogen isotopes.

Ans: **Isotopes of Hydrogen:**

Hydrogen has three isotopes

1: Protium of ordinary hydrogen $_1^1H$ or $_1^1P$

2: Deuterium or heavy hydrogen $_2^1H$ or $_3^2D$

3: Tritium $_3^1H$ or $_1^3T$

All three having same atomic number but different number of neutrons.

Protium $_1^1H$ =

Ordinary hydrogen or protium have no neutrons.

Atomic No = 1

Mass No = 1

No of Proton = 1

No of Electron = 1

No of Neutron 1 – 1 = 0

Deuterium: $_2^1H$

Similarly, deuterium has same number of electrons, proton & neutron

Atomic No = 1

Mass No = 2

No. of Proton = 1

No. of Electron = 1

Neutron 1 – 2 = 1

1. Tritium: $_3^1H$:

Atomic No = 1

Mass No = 3

No. of Proton = 3

No. of Electron = 1

No of Neutron = 1 – 3 = 2

Q4. How many electrons are present in each of following atoms? Assuming that each is a neutral atom identifies the element.

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

Ans: a. Total number of electrons are eleven (11) and the element is Sodium (Na) metal.

b. Total number of electrons are seventeen (17) and the element is Chlorine (Cl) nonmetal.

c. Total number of electrons are twenty (20) and the element is calcium (Ca) metal.

Q5. Why atom is considered as neutral particle? Give reason.

Ans: In atom the number of negatively charge electrons are equal to the number of positively charged proton. They are equal in number and cancelled the effect of each other. Therefore, atom as a whole in neutral particle.

Q6. The mass of an atom is present in its nucleus. Can you explain it?

Ans: Nucleus of the atom is composed of protons and neutron. Protons and neutrons are heavy sub-atomic particles of the atom and they occupied central position in the atom. Therefore, most of the mass is present in the center.

Q7. What is the reason that physical properties of the isotopes are different but their chemical properties are the same?

Ans: Isotopes have different number of neutrons or atomic masses, which shows physical chemical properties. There, isotopes have different physical properties but have same chemical properties.

Q8. Draw the structure of carbon isotopes. Then write down the number of proton, neutron and electron.

Ans: Carbon has 3 isotopes, carbon – 12, carbon – 13 and carbon – 14

$C_6^{12} C_6^{13} C_6^{10}$

Carbon has atomic number = 6

Carbon – 12 = C_6^{12}

Atomic No = 6

Mass No = 12

No. of electrons = 6

No. of protons = 6

No. of neutrons = 12 – 6 = 6

Carbon – 13 = C_6^{13}

Atomic No = 6

Mass no = 13

No. of electrons = 6

No. of protons = 6

No. of neutrons = 6

No. of neutrons = 13 – 6 = 7

Carbon- 14 = C_6^{14}

Atomic No = 6

Mass No = 14

No. of electrons = 6

No. of protons = 6

Q9. How many electrons could contained in K, L, M and N energy levels.

Ans:

Number of shell (n)	Name of shell	Formula	No. of electrons
1	K	$2n^2$	$2(1)^2 = 2(1) = 2es$
2	L	$2n^2$	$2(2)^2 = 2(4) = 8es$
3	M	$2n^2$	$2(3)^2 = 2(9) = 18es$
4	N	$2n^2$	$2(4)^2 = 2(16) = 32es$

Q10. Write detailed electronic configurations for Li_3^7, C_6^{12} and Mg_{12}^{24}.

Ans. **Electronic Configurations:**

a. Li_3^7:

Atomic number of Li = 3

K = 2es, L = 1e

$1s^2, 2s^1$

Period = 2nd , Group = 1A

b. C_6^{12}

Atomic number of C = 6

K = 2es, L = 4es

$1s^2, 2s^2, 2p^2$

Period = 2nd , Group = IVA

c. Mg_{12}^{24}

Atomic number of Mg = 12

K = 2es, L = 8es, M = 2es

$1s^2, 2s^2, 2p^6, 3s^2$

Period = 3rd , Group = IIA

Q11. 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.

Ans:

	ISOTOPE NAME	SYMBOL
a	Tritium (T)	$_3H^1$ or T
b	Deuterium (D)	$_2H^1$or D
c	Protium (H)	$_1H^1$

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Periodic Table and Periodicity of Properties

(TOPIC WISE QUESTIONS)

Q1: Define the modern periodic law. Discuss the periods and groups in the modern form of the periodic table.

Ans: **Modern Periodic Law:**

In 1911, Moseley presented a new idea for the classification of elements on the bases of increasing their atomic number instead of atomic masses, It is stated that;

The physical and chemical properties of the elements are the periodic function of their atomic number. This means that the elements are arranged in ascending order. (increasing order) of their atomic number. The elements possessing similar properties and valence shells electronic configuration were repeated at a regular interval.

Periods:

The horizontal rows in the periodic table are called period. There are seven periods in the modern periodic table.

First Period:

1^{st} period is the shortest period containing two elements Hydrogen (H) S-Block and Helium (He) P-Block.

2^{nd} and 3^{rd} Period:

2^{nd} and 5^{th} period are called long period. Each containing 18 elements out of 18, 2 elements are S-block (representative) 10 are outer transition (d-block) and 6 are P-block, elements (non-metal)

6^{th} Period:

The 6^{th} period is a long period containing 32 elements. Out of 32 elements, 2 are S-Block, 10 are outer transition or d-block and 6 are p-block. The remaining 14 are inner transition or lanthanide series of f-block.

7^{th} Period:

The 7^{th} period is also called long period. Still incomplete and continue it contains 2 S-block elements 6 P-block and 10 outer transition (d-block) and 14 inner transition elements (f-block) called actinides series. They are present at the bottom of the periodic table.

Groups:

The vertical columns in the periodic table are called groups. There are eight groups in the periodic table. These groups are further sub-divided into two subgroups. Sub group A called normal elements or representative elements and sub-group B are called transition elements.

The groups are given special number.

Group – I A = Alkali metals
Group – II A = Alkaline metals
Group – III A = Boron family
Group – IV A = Carbon family
Group – V A = Nitrogen family
Group – VI A = Chalcogens or oxygen family
Group – VII A = Halogens
Group – VIII A = Noble gases or inert gases or zero groups.

The groups IA and IIA are called S-block element. The group IIIA – VIII A are called P-block elements. All the S and P-block are called normal elements or representative elements. The transition elements are called the d-block elements while the lanthanide and actinide series are called f-block elements.

Q2. Define and explain the ionization energy of an elements. Discuss the periodic variation energies of the elements in the periodic table.

Ans: **Ionization Energy (I.E):**

The minimum energy required to remove an electron from its gaseous atom to form an ion (cation) is called ionization energy. It is also called ionization potential.

The energy which is required for the removal of electron is given in the form of heat, light or electrical conductance. The process by which electron is removed from the valence shell of an atom to form a gaseous positive ion (cation) is called ionization. The process of ionization can be represented by the following general equation.

$A_{(g)}$ + energy ➔ $A^+_{(g)}$ 1 e^-

Where "A" is an any gaseous atom another chemical examples are as follow:

$Na_{(g)}$ + energy ➔ Na^+ + 1e = 1.E = 496 Kj/mol

Mg $_{(g)}$ + energy ➔Mg^+ + 1e = 1.E = 738 Kj/mol

The above two values for $Na_{(g)}$ and $Mg_{(g)}$ atoms are called first ionization potential. The energy required to remove second electron in the 2nd attempt is called second ionization potential.

e.g. Mg^+ energy ➔ Mg^{++} + 1e- I.E = 1450Kj/mol

The energy required to remove the 3rd electron from an atom is very difficult in the 2nd and 3rd attempt due to higher charge density and smaller size of the atom. The removal of 1st electron causes the shrinkage in size of the atom as well as increase in nuclear charge. This make the valence electron tightly bounded and increases the value of ionization energy for 2nd and 3rd electron.

It is observed that ionization energies of atoms depend upon several factors. They include

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

Variation of I.E values in the periodic table

Period Wise:

In a period from left to right the I.E increase with the increase in atomic number. This is because with the increase in atomic number the charge on the nucleus also increase which leads to a stronger force of attraction between the nucleus and electrons. This ultimately causes a decrease in the atomic size and hence the valence electrons need more energy for their removal.

Group Wise:

In a group from top to bottom the I.E decrease. This is because of successive addition of electronic shells due to which valence electronic are placed at larger distance from the nucleus. As the force of attraction between the nucleus and valence electron decrease with the increase in distance. The valence electrons can be easily removed.

Q3. What do you mean by Electron affinity of an elements? Discuss its periodic variation in the table.

Ans: **Electron Affinity (E.A):**

Electron affinity means love for an accepting electron. The electron affinity of an atom is measured in term of energy. It is therefore defined as

"The energy released when an electron is absorbed by (added to) the gaseous atom in its outer most shell to form an anion (-ve charges)".

The energy released is measure in joules or kilo joules per mole of an element.

For example

Cl + 1e ➔ Cl = 349kj/mol

The new incoming electron when absorbed by the atom is tightly bound by the nucleus through attractive force. The case an evolution of heat energy. The heat energy is released outside, therefore the sign of E.A will be negative.

Atom with smaller atomic radii greater nuclear charge and poor shielding effect have usually high E.A values.

Variation of E.A values in the periodic table:

Period Wise:

In a periodic wise from left to right, the value of E.A increase with the increase in atomic number. As the atomic number increase the charge on the nucleus is also increase which lead to a stronger force of attraction between the nucleus and in coming electron. As a result, atomic radii greater nuclear charge and poor shielding effect have usually high E.A values.

Group Wise:

In a group from top to bottom the value of electron affinity (E.A) decreases. This is because of successive addition of electronic shells due to which attraction between the nucleus and incoming electrons are decreases.

When the atomic size increases group wise increases, the incoming electrons are less tightly bounded to the nucleus. Thus E.A decreases. The shielding effect of low-lying electron also causes to decrease the force of attraction between the nucleus and incoming electrons, Thus electron affinity is decreases from top to bottom.

Q4. Define the electro negativity of an element. Discuss its periodic variation in a period and in a group in the periodic table?

Ans: **Electro Negativity:**

The ability of an atom of an element to attract the shared pair of electrons towards its self in a covalent bond is called electro negativity.

E.N is a property associated with the atoms. When they are chemically bonded to the each other in a covalent bond the two atoms involve in bond formation mutually share electrons. This shared pair of electrons is then attached by the nuclei of both the atoms. But different atom has different abilities to attract the shared pair of electrons to form covalent bond. If the two atoms have the same ability to attract the shared pair of electrons equally both them is said to be Non-polar covalent bond. e.g.

1: $H_2 \rightarrow H - H$

2: $Cl_2 \rightarrow Cl - Cl$

On the other hand, if the bond is formed between atoms of different E.N it is said to be polar covalent bond e.g.

H_2O H^{-8} O^{-8} H^{-8}

HCl H^{-8} cl^{-8}

The E.N of an element mainly depend upon on

The Atomic size:

The larger the atomic size of an element the smaller will be E.N of that element.

Nuclear Charge:

The smaller the nuclear charge of an atom, the smaller will be E.N of that element.

The nature of bond formation.

Pauling calculated and developed the orbitary E.N values and assigned fluorine (F) a value variation of E.N values in the periodic table.

Period Wise:

In a period from left to right the values of E.N increase due to the decrease in atomic size and increase in nucleus charge.

Group Wise:

In a group from top to bottom, the E.N value of the element generally decreases. This is due to the increase in the atomic size of the elements.

Properties	Left to right (Period Wise)	Top to bottom (Group Wise)
I.E	Increases	Decreases
E. A	Increases	Decreases

Shielding Effect	Constant	Increases
F. N	Increases	Decreases
Atomic Size	Decreases	Increases

Periodic Table and Periodicity of Properties

(LONG QUESTIONS)

Q1: How modern periodic table is different from Mendeleev periodic table?

Ans: **Differences between Modern periodic table and Mendeleev periodic table:**

MODERN PERIODIC TABLE	MENDELEEV PERIODIC TABLE
a. This periodic table was put forward by Henry Moseley, an English physicist in 1914.	a. This periodic table was put forward by Dmitri Mendeleev a Russian chemist in 1869.
b. This periodic table is based on increasing order of atomic number of elements.	b. This periodic table is based on increasing order of relative atomic masses of elements.
c. It is based on modern periodic law which states that "The physical and chemical properties of the elements are the periodic function of their atomic number.	c. It is based on Mendeleev periodic law which states that "The physical and chemical properties of the elements are the periodic function of their relative atomic masses.
d. It contains 118 elements.	d. It contains 64 elements.
e. In this table elements are arranged in groups in such manner that elements in same group showing same physical and chemical properties.	e. In this table elements were arranged in groups in such manner that elements in same group showing different physical and chemical properties. Such as Ar with K, Co and Ni.
f. In this table the position of f-block elements (lanthanides and actinides) is clear.	f. In this table the position of f-block elements (lanthanides and actinides) is not clear.
g. There are eighteen groups which are divided into eight sub-groups (sub-group A and sub-group B) and seven periods in this table.	g. There are eight groups and twelve periods in this periodic table.
h. This table is known as Long form of periodic table.	h. This table is known as Short form of periodic table.

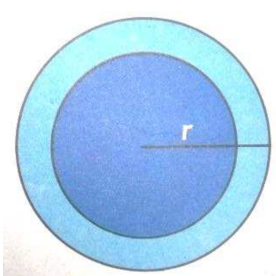
Q2. Differentiate between Atomic radii and covalent radii. Explain the trends of atomic radius in the periodic table.

Ans: **Atomic Size:**

The size of the atom is not rigidly fired but it varies when atom combined with other atoms. The same atom may have different sizes in different combination. The atomic sizes are usually expressed in term of atomic radii or radius, covalent radii or ionic radii.

Atomic Radii or Radius:

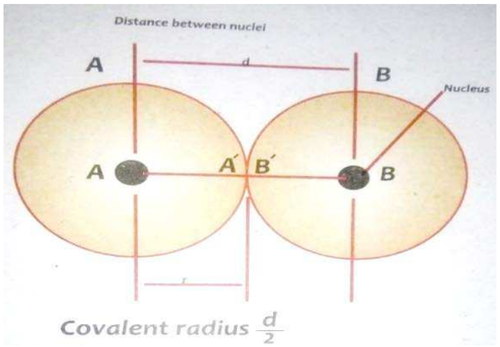
The distance between the nucleus and the valence shell of an atom is called atomic radii or radius.



One way for the prediction of the size of the single atom is the covalent radius may be defined as:

Covalent Radius:

Half of the distance between the centers of the two adjacent bonded atoms are called covalent radius.



The unit of atomic radii are Nanometer (10^{-9}), Picometer (10^{-12}m) and Angstrom (10^{-10}cm)

Trends in the Periodic Tables

Period Wise:

i. In a periodic the atomic size decrease from left to right. This is because one proton is added to the nucleus period wise. Therefore, the attraction of the nucleus is increase for the valence shell electron which pull them nearer to the nucleus.

Group Wise:

ii. In a group the atomic size increase from top to bottom because the shell after shell are added due to which the atomic size increases.

Q3. What is electronegativity? Identify the most and least electronegative groups of elements in the periodic table. Why fluorine is special in terms of electronegativity?

Ans: **Electro Negativity:**

The ability of an atom of an element to attract the shared pair of electrons towards its self in a covalent bond is called electro negativity.

E.N is property associated with the atoms. When they are chemically bonded to the each other in covalent bond the two atoms involve in bond formation mutually share electrons. This shared pair of electrons is then attached by the nuclei of both the atoms. But different atom has different abilities to attract the shared pair of electrons to form covalent bond.

The E.N of an element mainly depend upon on

The atomic size.

Nuclear Charge.

The nature of bond information.

Besides this three points electronegativity also depends upon atomic volume, the value of electron affinity and the value of ionization energy.

The most electronegative group in periodic table group VIIA (Halogen family) in which fluorine has highest electronegativity value (4.0). The least electronegative group of periodic tables is IA (Alkali metals) in which cesium has the lowest electronegativity value (0.7).

Pauling calculated the electro-negativities values of element and made an arbitrary scale. On this Fluorine has assigned E.N value of 4.0 which is highest among all elements. Electronegativity E.N depends upon the atomic size. Greater the atomic size smaller will be its E.N. value and smaller atomic size show higher E.N value. Fluorine is the top element group VIIA (halogen) which are present in the right side of Periodic-table. Due to smaller size it’s considered the most E.N element the whole periodic table and that’s why it is special in term of electronegativity.

Q4. Define shielding effect and its affects the ionization energy, electron affinity and electronegativity?

Ans: **Shielding Effect:**

The reduction in force of attraction b/w the nucleus and electron by the electrons present in the inner sub-shell is called shielding effect is also called screening effect.

Electrons present in the inner shells cut off attraction between the nucleus and valence shells electrons due to which shielding effect is increases. Shielding effect is responsible for decrease in the force of attraction between the nucleus and electrons present in the valence shell. It has therefore, a direct impact on the atomic radii, ionization potential and electron affinities of the elements. As the shielding effect increases the atomic size will be also increases but the ionization potential and electron affinities will be decreases.

Variation of shielding effect in the periodic table:

Period Wise:

In a period from left to right, the number of electrons in the inner shells remain constant. Therefore, the shielding effect remain constant in a period the positive charge or the nucleus increases, with the increase in atomic number. As the atomic number increases the shielding effect decrease due to the less number of shells and the attraction between valence electron and nucleus increases. This result in the contraction of atomic size and increases in an ionization energy, electron affinities and electronegativities of the element.

Group Wise:

In a group from top to bottom the number of electronic shells increases. So, the number of electrons in the inner shells also increases. As a result, shielding effect increase. This is because a new shell is added each time down the group, which screen out the outer electrons form the nucleus and decrease the force of attraction between the nucleus and outer electrons. This cause a decrease in the I.E, E.A and E.N of the element down the group.

Q5. Explain the following terms:

- **a. Periodicity of Properties.**
- **b. Electron affinity.**
- **c. Modern periodic law.**

Ans: **a. Periodicity of properties:**

The properties of elements are gradually and repeated after some interval from left to right period wise and top to bottom group wise. Their repetition of properties of elements after a certain interval in group wise and period wise are called periodically of properties and the phenomena is called periodicity.

b. Electron Affinity:

Electron affinity means love for an accepting electron. The electron affinity of an atom is a measured in term of energy. It is therefore defined as:

"The energy released when an electron is absorbed by (added to) the gaseous atom in its outer most shell to form an anion (-ve charges)".

The energy released is measure in joules or kilo joules per mole of an element.

For example

$Cl + 1e^{-} \rightarrow Cl^{-} = 349$ km/mol

The new incoming electron when absorbed by the atom is tightly bound by the nucleus through attractive force. The case an evolution of heat energy. The heat energy is released outside, therefore the sign of E.A will be negative.

Atom with smaller atomic radii greater nuclear charge and poor shielding effect has usually high E.A values.

C. Modern Periodic Law:

In 1911, Moseley presented a new idea for the classification of elements on the bases of increasing their number, instead of atomic masses. It is stated that "The physical and chemical properties of the elements are the periodic function of their atomic number".

This mean that the elements are arranged in ascending order (Increasing order) of their atomic number. The elements possessing similar properties and valence shell electronic configuration were repeated at a regular interval.

Periodic Table and Periodicity of Properties
(SHORT ANSWERS)

Q1: Which element to group IA is not an Alkali metal and why?

Ans: Hydrogen of group IA is not an Alkali metal because.

a. Hydrogen is non-metal while group IA elements are metals.

b. Hydrogen is gas at room temperature while Alkali metals are solids at room temperature.

c. Hydrogen can make both covalent compounds and ionic compounds while Alkali metals only make ionic compounds.

d. Ionization energy of Hydrogen is very high as compare to alkali metals i.e. ionization of Hydrogen is 1312 kj/mol and the highest ionization energy among alkali metals is 520kj/mol.

Q2. Place the following elements in order of increasing ionization energy: Na, S, Mg and Ar.

Ans: Sodium – Magnesium – Sulphur – Argon (OR)

Na, Mg, S, Ar

Q3. Name the group and state the group number of each of the following elements.

a. K b. Ne c. Be d. Cl e. C

Ans:

	Elements	Symbol	Group No.
A	Potassium	K	Group IA
B	Neon	Ne	Group VIII A
C	Beryllium	Be	Group IIA
D	Chlorine	Cl	Group VIIA
E	Carbon	C	Group IVA

Q4. Which element is the most electronegative among C, N, O, Br and S? Which group does it belongs to?

Ans: The most electronegative element among C, N, O, Br and S is O (Oxygen). It is present in period 2nd and group VIA which is known as Chalcogen family or Oxygen family.

Q5. How do first ionization energies of representative elements vary across a period and down a group?

Ans: First ionization Energies of Representative Elements.

<u>Period Wise:</u>

In a period from left to right the I.E increase with the increase in atomic number. This is because with the increase in atomic number the charge on the nucleus also increase which leads to a stronger force of attraction between the nucleus and electrons. This ultimately causes a decrease in the atomic size and hence the valence electrons need more energy for their removal.

<u>Group Wise:</u>

In a group from top to bottom the I.E decrease. This is because of successive addition of electronic shells due to which the valence electronic are placed at larger distance from the nucleus as the force of attraction between the nucleus and valence electron decrease with the increase in distance. The valence electrons can be easily removed.

Q6. Which element is found in,

a. Period 2, Group VIIA   **b. Period 4, Group IIIA**

c. Period 5, Group VIA   **d. Period 1, Group VIIIA**

Ans: a. Fluorine (F) is present in Period 2, Group VII.

b. Gallium (Ga) is present in Period 4, Group IIIA

c. Tellurium (Te) is present in Period 5, Group VIA

d. Helium (He) is present in Period 1, Group VIIIA

Q7. How will you differentiate between representative and transition elements?

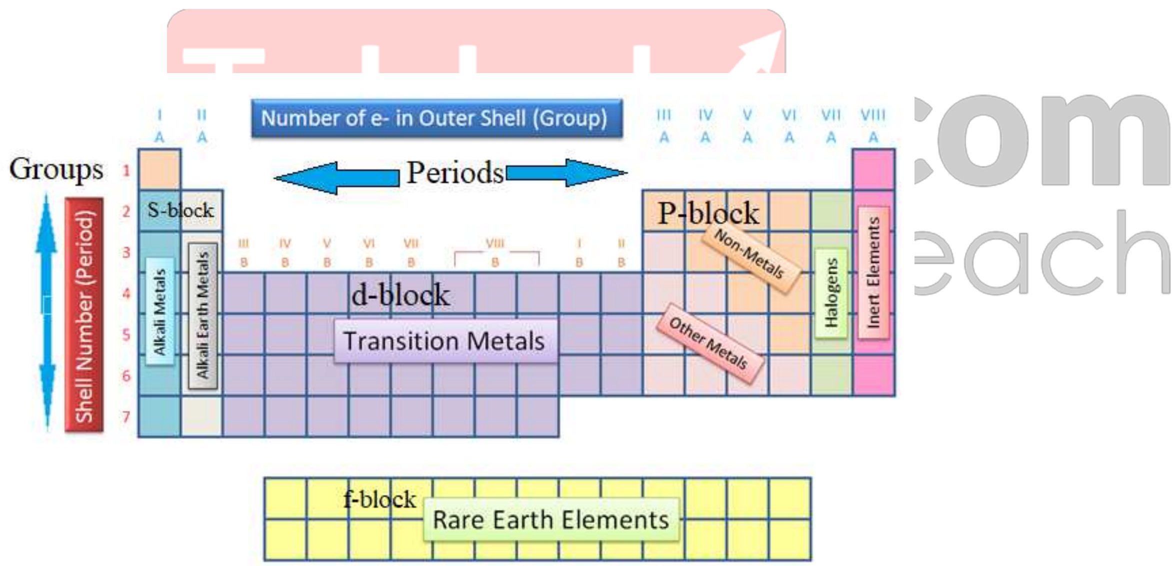
Ans: Differentiate between representative and transition elements.

REPRESENTATIVE ELEMENTS	TRANSITION ELEMENTS
a. All the s and p-block elements are called representative elements.	a. All the d and f-block elements are called transition elements.
b. They have "s" and "p" subshells in the process of completion.	b. They have "d" and "f" subshells in the process of completion.
c. They show constant valencies.	c. They show variable valencies.
d. Mostly they form colorless compounds.	d. Mostly they form colored compounds.
e. Representative elements contain metals, non-metals are also metalloids.	e. Transition elements are all metals.
f. All elements of sub group A (group IA to group VIIIA) are the example of representative elements.	f. All elements of sub group B (group IB to VIIIB) as well as lanthanides and actinides are examples of transition elements.

Q8. Make a general sketch of the periodic table showing s, p, d and f-block elements (without showing the symbols of elements).

Ans: **General sketch of the periodic table.**

- Sketch the general shape of the periodic table and label the s-, p-, d-, and f-blocks.



Q9. Why the s-bock elements have two groups only?

Ans: All elements of S-block will complete its sub shell by entering its last electron in their s-sub shell are called s-block element. The sub shell's can accommodate only two electrons with opposite spin. So, two groups come in this category which is group IA (Alkali metal) and group IIA (Alkaline earth-metal).

Q10. What type of element in Sulphur (S), a representative element, a transition element or lanthanide element?

Ans: Sulphur (S) is a representative element. As by definition that all the "S" and "P" blocks elements are called representative elements whose "S" and "P" sub shells are in the process of completion. Atomic number of Sulphur is 16 i.e. K = 2es, L= 8es and M = 6es. Its electronic configuration is $1s^2$, $2s^2$, $2p^5$, $3s^2$, $3p^4$. From electronic configuration it is clear that it has p-orbital in the process of completion so it is a representative element.

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Structure of Molecules
(TOPIC WISE QUESTIONS)

Q1. What is chemical bond?

Ans: **Chemical Bond:**

The attractive force which holds the atoms together in a molecular or compound during chemical reaction is called chemical bond.

During reaction old bonds are broken down on reactant hand side and new bonds are formed on product hand side to form a product.

Example:

H_2, O_2, Cl_2, N_2 etc.

In given examples each atom of every element combine together to form a molecule.

Q2. Write the two concepts which explain the chemical bonding.

Ans: **The two concept are:**

i. 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 to attain the inert or noble gas electronic configuration.

When atoms have two or eight electrons in their outermost shell, they are stable. The electron theory of valence can be named as octet or duplet theory of valence.

a. Octet Theory of valence or Rule of Eight:

The tendency of atoms to attain eight electrons in the outermost shell in order to attain stability.

For example oxygen (O) atom has six electrons in their valence shell. It shares or gains two electron in its outermost shell to attain the stability by completing its outermost shell with eight electron.

b. Duplet Rule or Rule of two:

The tendency of atoms to attain two electrons in the valence shell in order to attain sability.

For example, Helium (He) has two electrons in its valence shell and is stable.

The elements in Group VIIIA of periodic table are called noble gases. They are very stable and rarely take part in chemical reactions to form compounds. Their stability comes from their completely filled outermost shells.

Except Helium that has two electrons in outermost shell, all other noble gases have eight electrons in their outermost shells. A shell with eight electron is called an octed shell and is very stable. Thus, when atoms take part in a chemical reactions, they tend to combine in ways to complete eight electrons in their outermost shell, to attain the electronic configuration of the noble gases.

ii. 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 other, they overlap each other. This overlapping either endwise or sidewise, Endwise, Overlapping produce sigma bond and sidewise overlapping produce pi-bond.

Q3. Define ionic bond. Explain ionic bond formation in NaCl and $CaCl_2$.

Ans: **Ionic bond:**

A type of chemical bond which is formed by complete transfer of electron from one atom to another atom is called ionic bond.

The transfer of electrons between atoms completes the octets and duplets. This type of bond is always between metal and non-metal. Metals always lose electrons to form cations and non-metals always gain electrons to form anions.

In ionic bond formation, one atom loses electron(s) and other gains it. The atom that loses, the electron acquires positive charge and the other atom, which gains the electron, becomes 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.

i. Bond formation in Sodium Chlroide:

It is explained as follow:

Formation of Na^+ ion:

Atomic number of sodium = 11,

K = 2e

L = 8e

M = 1e

Sodium atom has one electron in its valence shell. It has larger atomic size and low ionization potential (I.P). Therefore, it gives one electron to Cl atom to form Na^+ ion and thus it attain the electronic configuration of the nearest noble gas Neon (K=2, and L=8)

Na energy $Na^+ + 1e^-$

(2,8,1) ➔(2, 8)

Formation of Cl ion:

Atmic number of chlorine = 17

K = $2e^-$

L = $8e^-$

M = $7e^-$

Chlorine atom has seven electrons in its valence shell. It has smaller atomic size value, therefore, it gains one electron from Na to form chloride in Cl and to attain the electronic configuration of the nearest noble gas Argon (K=2, L=8 and M=8)

$Cl + 1e$ ➔ Cl^-

(2,8,7) ➔ (2, 8, 8)

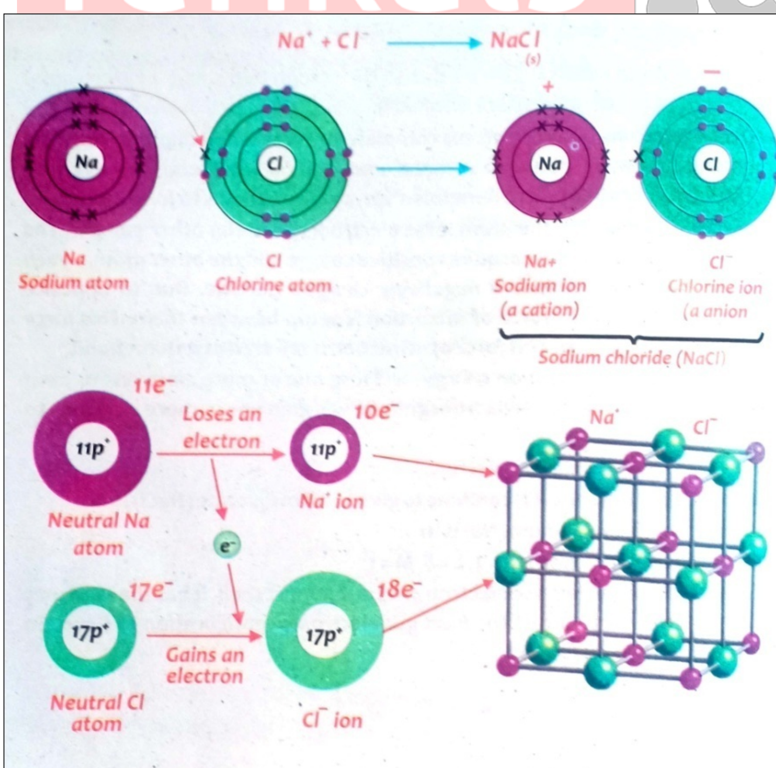


Fig. 4.2 Formation of NaCl compound

Establishment of Electrostatic force:

Now both these ions (Na+ and Cl) attract each other due to electrostatic force of attraction forming ionic bond. Thus sodium chloride is formed

Na^{+} Cl^{-} ➔ NaCl

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

It is explained as follow:

Formation of Ca^{+} ion:

Calcium atom has twenty (20) atomic number. It has two electrons in its valence shell. It has the tendency to lose two electrons to attain the electronic configuration of Argon. So the calcium (Ca^{+2}) is formed $Ca_{(2,8,8)} + 2e^{-}$

Formation of Cl^{-} ion:

Chlorine (Cl) has seventeen (17) atomic number. It has one electron in its valence shell. It needs one electron to complete its outermost shell to attain the electronic configuration of Argon. So it form chloride ion (Cl^{-}).

$Cl_{(2, 3, 7)} + e$ ➔ $Cl_{(2, 8, 8)}$

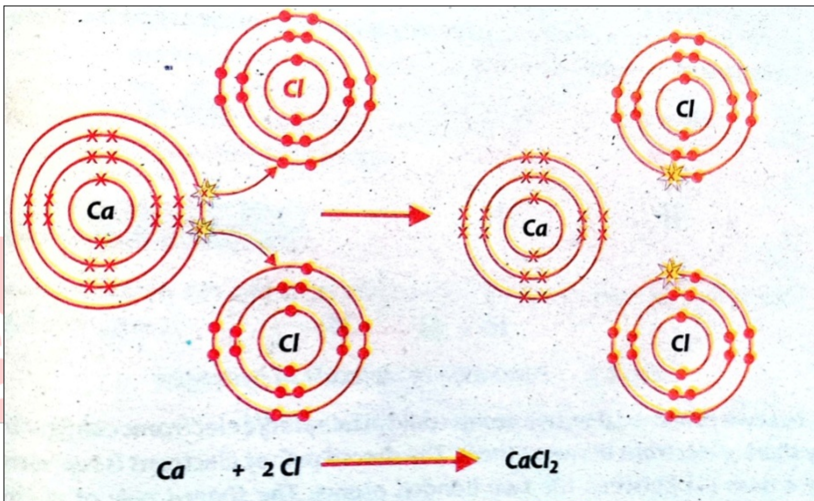


Fig. 4.3 Formation of $CaCl_2$

Establishment of Electrostatic force:

Now both these ions on Ca^{++} and two Cl^{-} ion attract each other due to electrostatic force of attraction forming ionic bond. Thus calcium chloride is formed.

Ca + 2Cl ➔ $CaCl_2$

This explains the divalent nature of Calcium (Ca).

Q4: Define covalent bond and explain its types in detail.

Ans: **Covalent Bond:**

A type of chemical bond which is formed by the mutual sharing of electrons b/w two atoms is called a covalent bond. (OR)

It is that force of attraction that arises b/w two atoms due to the mutual sharing of an electron or electrons.

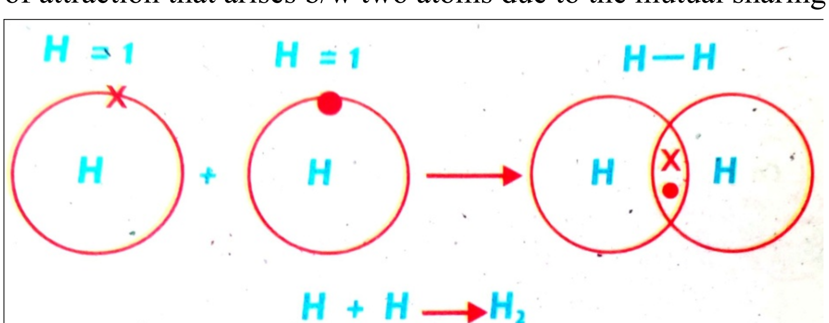


Fig. 4.4 Formation of molecule of hydrogen

Explanation:

Every atom wants to complete its valence shell to become stable and lower their energy by transferring of an electron is not only necessary. However some atoms decrease their energies by sharing their electrons.

The two bonding atoms contribute the same number of electrons to bond formation. The atoms will share their valence shell electrons are not localized by only one atom. The share electrons b/w the atoms are responsible for lowering the energies and shifting towards stability.

One the basis of electron/electrons sharing, covalent bond may be classified as single, double and triple covalent bonds.

Single Covalent Bond:

The bonds in which two atoms share one electron each they 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 (Cl_2). Therefore, both chlorine atoms attain inert gas (Ar) electronic configuration and complete their octet.

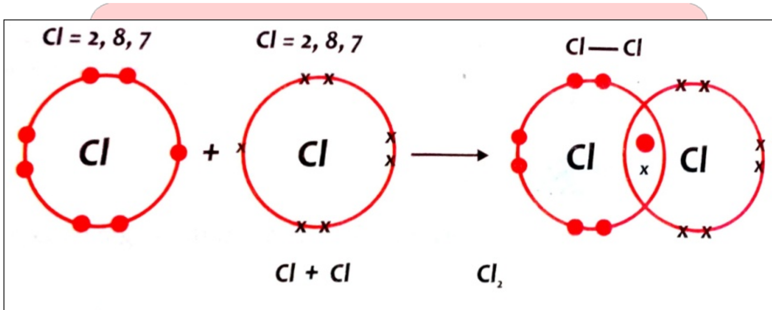


Fig. 4.6 Molecule of Chlorine

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

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

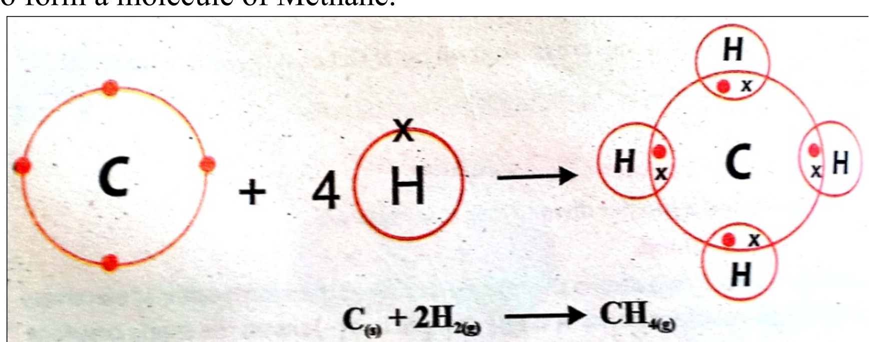


Fig. 4.7 Molecule of Methane

ii. Double Covalent Bond:

The bonds in which two atoms share two electrons each they 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 to form oxygen (O_2). In this way, both oxygen atoms attain inert gas (Ne) electronic configuration and complete their octet.

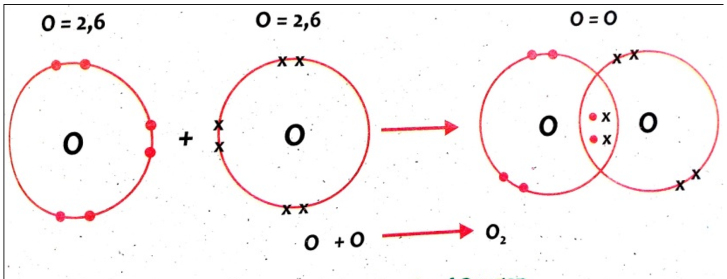


Fig. 4.8 Molecule of Oxygen

b. Carbon dioxide (CO_2)

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

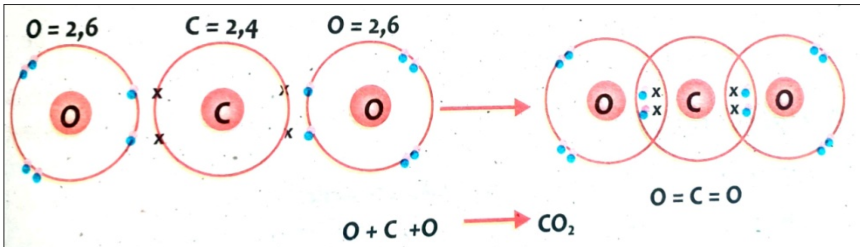


Fig. 4.9 Molecule of Carbondioxide

iii. Triple Covalent Bond:

The bonds in which two atoms share three electrons each they form three pairs of electrons is called triple covalent bond.
A triple straight line shows such 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 its shares three electrons with another nitrogen to form nitrogen (N_2). In this way, both nitrogen atoms attain inert gas (Ne) electronic configuration and complete their octet.

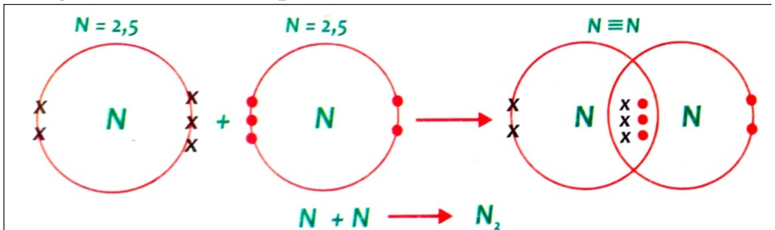


Fig. 4.10 Molecule of Nitrogen

Q5. Define co-ordinate covalent bond.

Ans: **Co-ordinate covalent bond or dative bond:**

The type of covalent bond in which the shared pair electrons is denoted by only one of the two bonded atoms is known as co-ordinate covalent bond or dative bond.

Explanation:

In the formation of co-ordinate covalent bond, the atom which denotes the electron pair is called donor atom which accepts it is known as acceptor atom.

Generally, those atoms which have lone pair of electrons act as a donor atoms while electrons deficient species act as acceptor atom. Co-ordinate covalent bond is just different from covalent bond in mode of formation and after formation in adopts the covalent characters.

Representation:

Co-ordinate covalent bond is represented by an arrow, the head of arrow which is directed towards the acceptor atom.

Examples:

i. Formation of Ammonium Ion (NH_4^+):

In ammonia molecule, the nitrogen atoms is bonded to three Hydrogen atoms. There is still one unshared pair of electrons with the Nitrogen, an electron rich species. The Hydrogen (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^+).

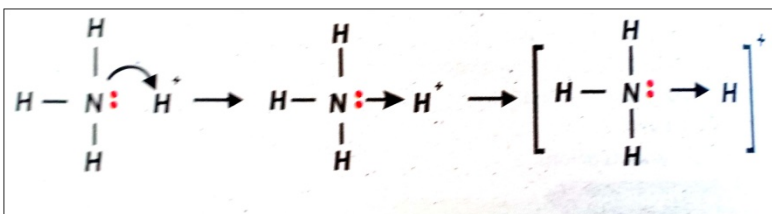


Fig. 4.11 Structure of Ammonium

ii. Formation of Hydronium ion (H_3O^+):

In the above example the hydrogen of water molecule give a lone pair of electrons to hydrogen ion forming a coordinate covalent bond to form hydronium ion.

iii. 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 BCl_3 is short of two electrons to complete its octet. An addition compound is formed when nitrogen of NH_3 donates this lone pair of electron to the Boron of BCl_3 which accepts it and forms the addition compound.

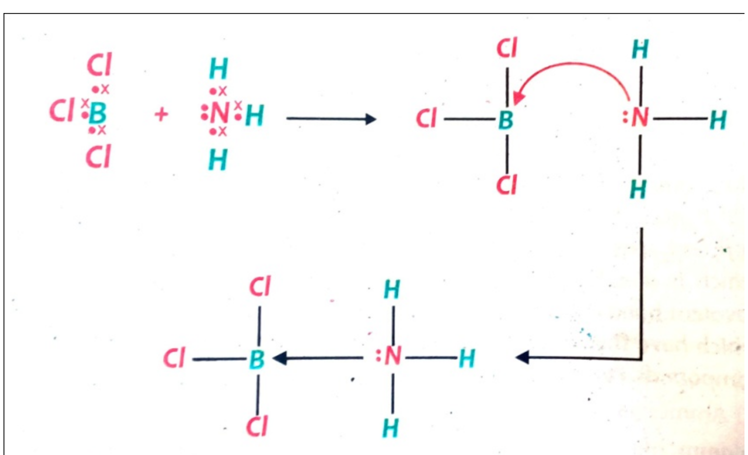


Fig. 4.12 Structure of NH_3 and BCl_3

Q6. What is meant by Lewis structure?

Ans: **Lewis structure:**

"The structure which represents the valence electrons of atoms or molecules by dots and crosses is called Lewis structure or dot and cross structure".

This structure was investigated by G.N Lewis for covalent compounds.

In this structure, the valence electrons of atoms or molecules are written around the symbols for shared pair of electrons by a dot or cross.

Examples:

The Lewis structure of some atoms and molecules are given below:

Hydrogen gas = H_2 Lewis structure: H•——×H

Oxygen gas = O_2 Lewis Structure: O×× ═══ ••O

Q7. What is metallic bond?

Ans: **Metallic Bond:**

The force of attraction b/w the positive metal ions and the mobiles sea electrons is called metallic bond. It is a type of bond which is formed b/w the atoms of a metal through free electrons. Metallic bond exists only in metal such as Aluminum, Gold, Copper and Iron etc. In metals, each atom is bonded to several other metal atoms and their electrons are free to more throughout the metal structure.

Explanation:

Metallic bond can be explained by electrons sea theory or model.

Electron Sea Theory:

Drude and Loven proposed a theory of metallic bonding called electron gas or electron sea theory. According to this theory "the valence electrons of metal atoms are not firmly held by the nucleus due to the large size and low I.P of metal atoms. The outer valence electrons are lost by the atoms and thus from a sea of free electrons the positively charged metal ions are held together by this sea of electrons and are at a measureable distance from each other in crystals.

The valence electrons do not belong to an individual ion. They move freely but do not escape from the bulk. When an electric current is passed, electrons jump from one atom to the other and in this way metal conduct electricity.

Q8. Draw the shapes of molecules:

Ans: **Shapes of molecules:**

Molecules are very small and cannot be seen with naked eyes. However, their shapes are determine by various experiments. The experiments have shown that the molecules may be linear bend, tetrahedral, pyramidal and pyramidal etc.

Example:

Co_2 molecules have linear shape BCl_3 has triangular shape, CH_4 has tetrahedral shape, NH_3 has pyramidal shape and PCL3 has pyramidal shape.

i. Carbon-dioxide (Co_2):

It is linear in shape.

ii. Water (H_2O):

It is bent or angular in shape

iii. Boron tri-chloride (BL_3)

It is tetrahedral in shape.

iv. Ammonia (NH3):

It is pyramidal in shape

v. Methane (CH_4):

It is tetrahedral in shape

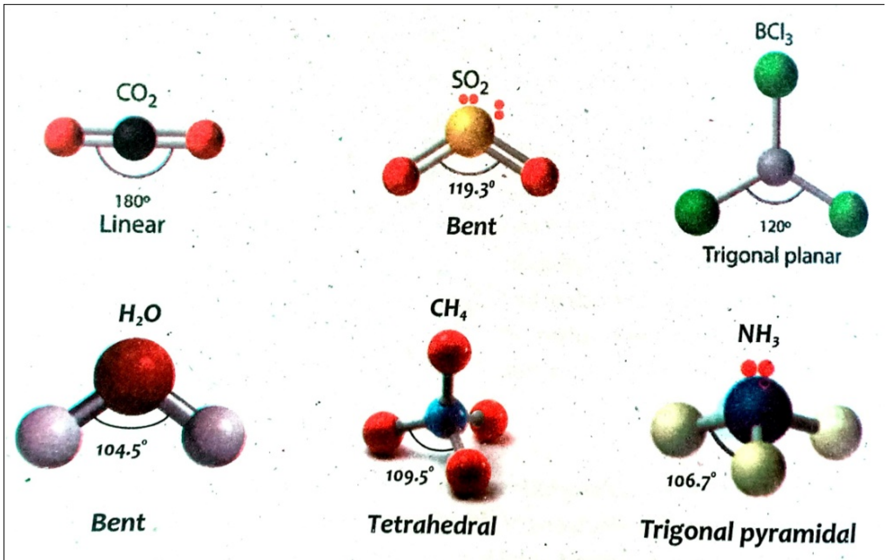


Fig. 4.18 Shapes of CO_2, SO_2, BCl_3, H_2O, CH_4 and NH_3

Q9. What is intermolecular forces? Explain with reference to example.

Ans: **Intermolecular forces:**

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

Explanation:

The intermolecular forces are of three types and collectivity called Van Der Walls forces. Intermolecular forces are the weaker forces of attraction. It is 25 time weaker than covalent bond.

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 and boiling point of liquids depends on the strength of these forces.

Example:

In H_2O molecules, it require 464 kj.mol to break the H-O bonds within a water molecule and needs only 19 kj.mol to break the intermolecular forces between water molecules.

Q10. Explain Dipole-Dipole interaction in detail.

Ans: **Definition:**

The type of interaction forces in which the positive or negative end of one polar molecule attract the negative or positive end of the other molecule is known as Dipole-Dipole interaction.

Explanation:

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

A polar molecule is formed between different atoms which has two ends-negative and positive. Therefore, they are called Dipole-Dipole when these molecules are brought closes to each other the positive end of one molecule attract the negative end of the other molecule is called Dipole-Dipole interaction.

Examples:

H^{+8}……. Cl^{-8}

In above example, the dipole-dipole interaction b/w hydrogen (O+) and chlorine (O-) is shown by (…….). A network of partial positive (O+) and negative (O-) charges attract molecules to each other. Because of force of attraction between oppositely charge ends, there is a small dipole-dipole force of attraction between the adjacent HCl molecules. The dipole-dipole interaction is relatively weak; only 3.3 kj.mol energy is required to break that interaction. The force of attraction is so small that Hydrogen chloride (HCl) boils at -85.0^{0}c

Factors affecting Dipole-Dipole forces:

The following two factors affect the dipole-dipole forces i.e. intermolecular distance and electro negativity.

i. Intermolecular distance:

Greater the intermolecular distance weaker will be the dipole-dipole forces and vice versa. The dipole-dipole forces are very weak in gases because the molecules of gases are far apart from each other.

ii. Electro negativity difference:

Greater the electro negativity of bonded atoms, stronger will be dipole-dipole forces and vice versa. Those molecules have strong dipole-dipole forces which have high melting points and boiling points etc.

Q11. What do you understand from Hydrogen bonding?

Ans: **Hydrogen Bonding:**

The force attraction b/w the partial positively charged hydrogen atom to one molecule and the partial negatively charged atom of other molecule is known as Hydrogen bonding. (OR)

"When Hydrogen is covalently bonded to E.N atoms (O_2, F_2, Cl_2) then the strong inter molecular force exists is called Hydrogen bonding.

Explanation:

The hydrogen bonding is shown by dotted lines. Hydrogen bonding is only formed when hydrogen is bonded with F, N or O. This is a type of bond containing intermolecular forces which are roughly ten times stronger than dipole-dipole forces. However, hydrogen bonding is roughly ten times weaker than ionic, covalent or coordinate covalent bond.

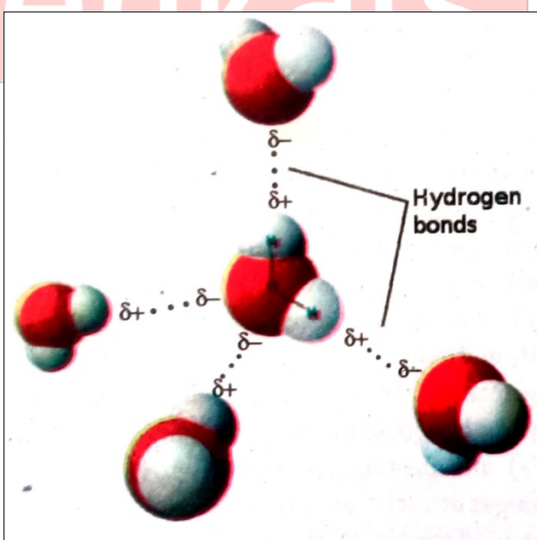


Fig. 4.20 Hydrogen Bonding in water

Example:

Consider water (H_2o) in which oxygen is more electronegative therefore the bonded electron pair b/w hydrogen and oxygen will be attracted more by the oxygen atoms which will produce a partial positive charge on hydrogen atom. As a result, the partial negatively charged oxygen atoms of another water molecule from hydrogen bonding.

Similarly, hydrogen bonding can be seen in HF, NH_3 etc.

Q12. Write a note on Application and properties of Hydrogen Bonding.

Ans: **Application of Hydrogen Bonding:**

i. These forces are extremely important in determining properties of water, biological molecules such as protein and synthetic material such as glue, paints, resins etc.
ii. The adhesive action of paints and dyes is developed due to Hydrogen bonding.
iii. Synthetic resins also binds two surfaces together by hydrogen bonding.

Properties of Hydrogen Bonding:

i. Hydrogen bonding is stronger than dipole-dipole interaction but weaker than covalent bonding. It is about twenty times weaker than covalent bonding and ten times stronger than the dipole-dipole interaction.
ii. Hydrogen bond is directional.
iii. Hydrogen bond forms long chains and help in the formation of networks of molecules.

Q13. Write the properties of ionic compounds?

Ans: **Properties of ionic compounds:**

i. All the ionic compounds are solids at room temperature.
ii. Ionic compounds have sharp melting and boiling points.
iii. Ionic compounds are soluble in polar solvents like water.
iv. Ionic compounds are good electrolytes in molten state or solution form.
v. Ionic compounds have non-directional bonds.
vi. Ionic compounds have reactions in molten form or solution form.
vii. Ionic compounds are composed of cations and anions in crystalline form.

Q14. Write the properties of covalent compounds.

Ans: **Properties of covalent compounds:**

Covalent compounds are also called molecular compounds. The properties of covalent compounds depend upon.

i. Geometrical shape of molecular.
ii. Polarity and intermolecular forces among molecules.
iii. Bond type whether single, double or triple. Some of the properties of covalent compounds are given as follows:

a. Covalent compounds have low melting and boiling points.
b. Covalent compounds are non-electrolytes in their solution form.
c. The bonds in covalent compounds are directional.
d. The crystals of covalent compounds are composed of molecular crystals.
e. Reaction of covalent compounds are slower than the ionic compounds.
f. Polar covalent compounds are soluble in polar solvent like water, Alcohol etc. While non-polar covalent compounds are soluble in non-polar solvents like carbon tetrachloride (CCl_4), Benzene (C_6H_6), Acetone (C_3H_6O) etc.

Q15. Write a note on properties of metals.

Ans: **Properties of metals:**

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

i. All metals are solid at room temperature and on atmosphere pressure except mercury.
ii. Metals are malleable i.e. they can be beaten into sheets and foils.
iii. Metals are ductile i.e. they can be drawn into wires.
iv. All the metals are good conductors of heat and electricity.
v. Metals are lustrous i.e. they have shiny surfaces.
vi. Metals are sonorous i.e. they produce ringing sound when struck.

Structure of Molecules
(LONG QUESTIONS)

Q1. Describe the octet rule in terms of noble-gas configuration and stability.

Ans: **The valence concept (Elecronic 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, atom take part by losing, gaining or sharing of electrons, so as to attain the inert or noble gas electronic configuration.
When atoms have two or eight electrons in their outermost shell, they are stable. The electron theory of valence can be named as Octet or Duplet theory of valence.

Octet Theory of Valence or Rule of Eight:

This rule states that "The tendency of atoms to attain eight electrons in the outermost shell in order to attain stability".

Explanation:

The elements in group VIIIA 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 stable and rarely take part in a chemical reactions to form compounds. Their stability comes from their completely filled outermost shells.
Except to helium it has two electrons in the outermost shell, all the other noble gases have their outermost shells filled with eight electrons. A shell with eight electrons is called an octet shell and is very stable. Thus, when atoms take part in chemical reactions, they tend to complete eight electrons in their outermost shell, to attain electronic configurations of the noble gases (except Helium).

Example:

Oxygen (O) atom has six electrons in their valence shell. It shares or gains two electrons in its outermost shell to attain the stability by completing its outermost shell with eight electron.

Q2 (a). What is the main distinction between ionic and covalent bonding? Explain your answer with suitable examples.

Ans:

IONIC BOND	COVALENT BOND
A type of chemical bond which is formed by complete transfer of electron from one atom to another atom is called ionic bond.	A type of chemical bond which is formed by the mutual sharing of electrons b/w two atoms is called a covalent bond.
It contains ions.	It contains atoms (Atoms may be neutral or partially charged.
It is usually exists between metal and non-metal.	It usually exists between non-metals.
It is non-directional.	It is directional.
It is formed between those atoms, having electro negativity difference more than 1.7.	It is formed between those atoms, having electro negativity difference less than 1.7.
It is not classified i.e. it has no types	It is classified further i.e. it may be single, double or triple.
Examples: $NaCl$, KCl, CaO, MgO and Al_2O_3 have ionic bond.	**Examples:** H_2, NH_3, H_2O, CH_4 have covalent bond.

Q2 (b). How is electronegativity used in determining the ionic or covalent character of the bonding Between two elements?

Ans: **<u>(b) The difference in electronegativity of two bounded elements:</u>**

1. If the E.N difference between two bonded atoms is greater than 1.7 then the bond will be ionic.
For example: NaCl
E.N of Na = 0.9
E.N of Cl = 3.0
Difference between the E.N = 3.0 – 0.9 = 2.1
Hence the bond NaCl is ionic in nature.

2. If the E.N difference between two bonded atoms is from 0 to 0.4 then the bond will be non-polar covalent.
For Example: Cl_2
E.N of Cl = 3.0
E.N of Cl = 3.0
Difference between the E.N = 3.0 – 3.0 = 0
Hence the Cl_2 bond is non-polar covalent in nature.

b. if the E.N difference between two bonded atoms is from 0.5 to 1.6 then the bond will be polar covalent.
For example: HCL
E.N of H = 2.1
E.N of Cl = 3.0
Difference between the E.N = 3.0 – 2.1 = 0.9
Hence the HCl bond is polar covalent in nature.

3. If the E.N difference between two bonded atoms is 1.7 then the bond will be 50% ionic and 50% covalent.

Q3. Draw the Lewis structure for each of the following compounds.

a. CO b. HCl c. SO2 d. CCl4 e. BF3 f. NH3

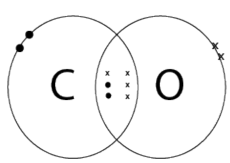
Ans:

a. <u>CO:</u>

Valance electrons of C = 4.
Valance electrons of O = 6

<u>Lewis structure:</u>

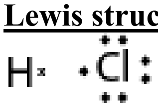
:C≡O: OR



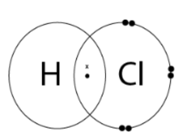
b. <u>HCL:</u>

Valance electrons of H = 1.
Valance electrons of Cl = 7

<u>Lewis structure:</u>



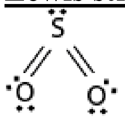
OR



c. <u>SO_2:</u>

Valance electrons of S = 6.
Valance electrons of O = 6

<u>Lewis structure:</u>



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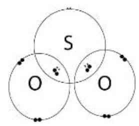
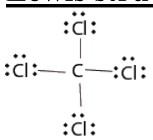


d. <u>CCl_4:</u>

Valance electrons of C = 4.
Valance electrons of Cl = 7

<u>Lewis structure:</u>



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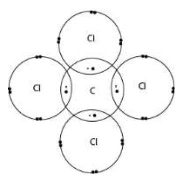
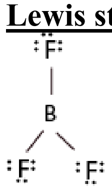


e. <u>BF_3:</u>

Valance electrons of B = 3
Valance electrons of F = 7

<u>Lewis structure:</u>



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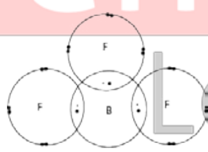
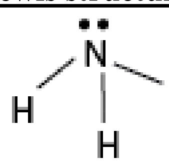


f. <u>NH_3:</u>

Valance electrons of N = 5
Valance electrons of H = 1

<u>Lewis structure:</u>



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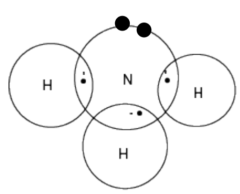


Q4. Explain why most metals are malleable and ductile but ionic crystals are not.

Ans: **<u>Metals are malleable and ductile but ionic crystals are not, because:</u>**

When external force is applied on metal, its layers slide over each other, thus change the position of the metal cations however, the attractive force between the metal cations and Electron Sea around it, remains same. Due to this reason, metals, do not break, so, they can be drawn into sheets and wires, without breaking.

While, the ionic solids, contains parallel layers of cation and anions in alternate position. So, when an external force is applied on it, then one layer of ions slide a bit over the other layer, so, the similar ions come in front of each other and hence begin to repel each other, due to this repulsion, ionic crystals cannot be drawn into sheets and wire i.e. break down.

Q5. (a) What is meaning of the term polar, as applied to chamial bonding?

Ans: **a. Polar:**

Polar means opposite charges i.e. positive and negative charges.

"Molecules with partial positive (+ive) and partial negative (-ive) charges on atom are called polar molecules and phenomena is called polarity". Polarity is shown by the covalent compounds containing different atoms and as well as by the ionic compounds.

In polar covalent compound, the electrons are unequally shared b/w the two atoms. The most E.N atoms attract the shared pair of electrons toward itself possess partial negative charge and the opposite atom possess partial positive charge. E.g

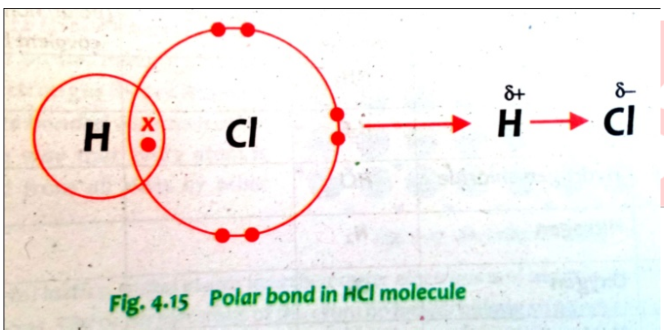


Fig. 4.15 Polar bond in HCl molecule

Q5. (b) Distinguish between polar-covalent and non-polar covalent bonds.

Ans: **b. Polar covalent bond:**

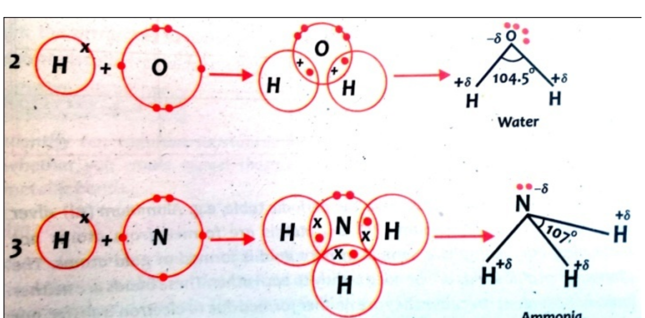
A type of covalent bond which is formed b/w atoms of different elements is called polar covalent bond.

Explanation:

As two different bonded atoms have different values of electro negativity so the shared pair of electrons will be attracted more towards the more electronegative atom. This will produce a partial negative charge on the more electronegative atom and a partial positive charge on the other atom. A polar covalent bond is also called Dipole (meaning two poles i.e. positive and negative).

Example:

The bond formed b/w hydrogen and chlorine atoms is polar because the shared pair of electron will be attracted more towards the more electronegative chlorine atom. This will produce a partial negative charge on chlorine atom and a partial positive charge on hydrogen atom. Similarly, the bond b/w hydrogen and oxygen in water is also polar covalent.



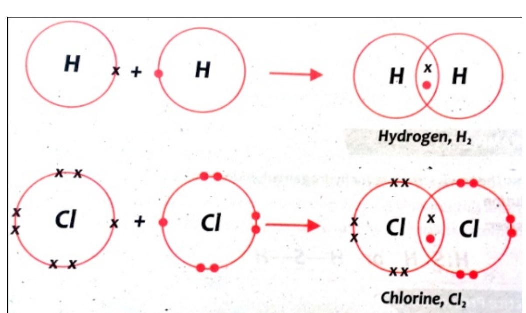


Fig. 4.14 Non-Polar bond in H_2 and Cl_2 molecules

Non-Polar Covalent bond:
A type of covalent bond which is formed b/w atoms of the same element is called non-polar covalent bond.

Explanation:
As two same bounded atoms have the same value of electro negativity. Therefore the shared pair of electrons will be equally attracted by both the atoms. The bond pair will remain in b/w the two atoms and no partial positive or negative charge will be produced as either end of the molecule.

Example:
$H + H = H - H$ or H_2
$Cl + Cl + = Cl$ or Cl_2
$N + N = N = N$ or N_2

Structure of Molecules
(SHORT QUESTIONS)

Q1. What is electron sea model of metallic bonding?

Ans: **Electron Sea Theory:**

The force of attraction b/w the positive metal ions and the mobile sea of electrons is called metallic bond.

Drude and Loven proposed a theory of metallic bonding called electron gas or electron sea theory. According to this theory the valence electrons of metal atoms are not firmly held by the nucleus due to the large size and low I.P of metal atoms. The outer valence electrons are lost by the atoms and thus from a sea of fee electrons the positively charged metal ions are held together by this sea of electrons and are at a measurable distance from each other in crystals.

The valence electrons do not belong to an individual ion. They move freely but do not escape from the bulk. When an electric current is passed, electrons jump from one atom to the other and in this way metal conduct electricity.

Q2. Why most atoms are chemically bonded to other atoms in nature?

Ans: Every system in the universe tends to lower its energy to attain stability like water which lower Level. Electricity which flows from higher potential to lower potential and like heat which move from hot bodies to cool bodies.

In the same way, the energy of two isolated atoms is more than two bounded atoms. To lower its energy and from a molecule and to attain inert gases configuration.

Example:

H_2 and Cl_2

i. H2 Molecule:

$H^{O} + H^{x} \longrightarrow$ H – H or H_2

Higher energy	lower energy
Unstable	stable
Reactive	un reactive

ii. Cl_2 molecules:

$H^{\cdot} + Cl^{x} \longrightarrow$ H – Cl or Cl_2

Higher energy	Lower energy
Unstable	Stable
Reactive	Unreactive

Q3. Identify and define the four major types of chemical bonding.

Ans: **The four major types of chemical bonding are:**

a. Ionic bond **b. Covalent bond** **c. Co-ordinate covalent bod**
d. Metallic bond

a. Ionic Bond:

Ionic bond is define as "The type of bond which is formed by complete transfer of electron(s) from one atom to other atom is called ionic bond.

It is also called electrovalent bond. This type of bonding is always formed between a metal and non-metal.

b. Covalent Bond:

A type of chemical bond which is formed by the mutual sharing of electrons b/w two atoms is called a covalent bond. (OR)

It is that force of attraction that arises b/w atoms due to the mutual sharing of an electron or electrons.

c. Co-ordinate covalent bond or dative bond:
The type of covalent bond in which the shared pair electrons is denoted by only one of the two bonded atoms is known as co-ordinate covalent bond or dative bond.
In the formation of co-ordinate covalent bond, the atom which denotes the electron pair is called donor atom while the atom which accepts it is known as acceptor atom.

d. Metallic Bond:
The force of attraction b/w the positive metal ions and the mobile sea electrons is called metallic bond. It is type of bond which is formed b/w the atoms of a metal through free electrons. Metallic bond exists only in metals such as Aluminum, Gold, Copper and Iron etc. In metals each atom is bounded to several other metal atoms and their electrons are free to more throughout the metal structure.

Q4. Arrange the following from strongest to weakest attraction:
a. Covalent bond
b. Dipole-Dipole interaction
c. Hydrogen bond d. Ionic bond

Ans: Ionic bond > Covalent bond > Hydrogen bond > Dipole-dipole interaction.

Q5. Why ionic compounds are good electrolyte is molten and solution form not in solid form.
Ans: Ionic compounds are good electrolyte is molten and solution form because in molten or solution ionic compounds contain free ions and conduct electricity while in solid form, (+ive) and (-ive) ions are tightly picked together and do not move from its position due to strong attractive forces and there is no free movement of ions.

Q6. What type of element/atoms tends to form the following types of bonding?
a. Ionic b. Covalent c. Metallic

Ans: **a. Ionic Bond:**
Ionic bond is formed metal and a non-metal i.e. ionic bond is formed between group IA elements and group VIIA elements or group IIA elements with group VIA elements. The electro negativity difference between elements for ionic bond is greater than 1.7. Example of ionic bond is NaCl.

b. Covalent Bond:
Covalent bond is mostly formed between non-metal and non-metal i.e. covalent bond is formed between group IVA, group VA, Group VIA and group VIIA elements. The electro negativity difference between elements for covalent bond is less than 1.7. Example of covalent bond is H_2O.

c. Metallic Bond:
Metallic bond is formed between metals. It is a type of bond which is formed between the atoms of a metal through free electrons among a lattice of cations. Example Au, Ag

Q7. Give an example of non-polar molecule with polar bonds. Give reasons.
Ans: **Co$_2$ is an example of non-polar molecule with polar bonds because:**
As we know that Co$_2$ is formed between one atom of carbon and two atoms of oxygen. Carbon and Oxygen mutual shares pair of electron. These shared paired of electron are equally attracted by both of atoms but due to high electro negativity of Oxygen the shared paired of electron will be more attracted towards oxygen creating partial positive charge on carbon and partial negative charge on oxygen. Hence there are two oxygen atoms on both sides of carbon and the structure of molecule is linear, therefore the negative pole of oxygen cancels the effect of other oxygen of other pole and molecule becomes non-polar as a whole.

Q8. Predict the bond type (ionic, polar covalent, non-polar covalent) in each of the following:
a. $CaCl_2$ b. H_2Oc. CO_2 d. C_2H_4

Ans: a. The bond in $CaCl_2$ is an ionic.
b. The bond in H_2O is polar covalent.
c. The bond in CO_2 in non-polar covalent.
d. The bond in C_2H_4 is non-polar covalent.

Q9. Why ionic compounds are good conductors compared to covalent compounds.

Ans: **Reason:**

Generally, ionic compounds completely ionize in aqueous solutions or in molten state, they are good conductors of electricity due to the movement of their ions.

While, covalent compounds generally do not ionize in aqueous solution or in molten state. They do not partially ionize in aqueous solution as compared to ionic compounds.

For Example:

Ionic compounds like NaCl, KCl, $MgCl_2$ are good conductors of electricity in aqueous solution while sugar, kerosene oil etc are bad conductors of electricity due to the lack of ions.

That is why, ionic compounds are good conductors as compared to covalent compounds.

Q10. Give the reason that why bond is always polar?

Ans: In co-ordinate covalent bond or dative covalent bond, the shared pair of electrons is donated by only one atom which is called donor atom. The other atom is acceptor atom. The donor atom, after giving electrons, develop a partial positive charge and due to the storage of electron the acceptor atom develop parial negative charge.

That is why dative or co-ordinate covalent bond is always polar

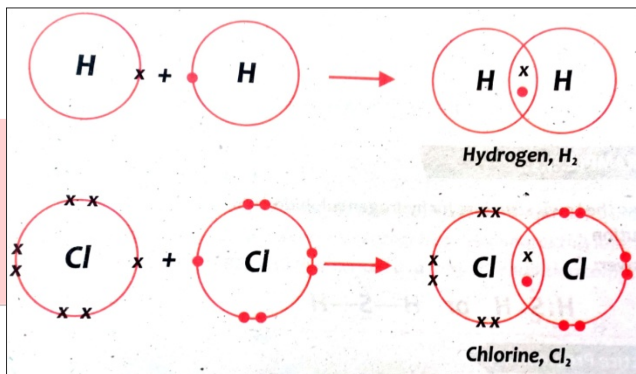


Fig. 4.14 Non- Polar bond in H_2 and Cl_2 molecules

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PHYSICAL STATES OF MATTER

(Topic Wise Questions)

Q1. Write the typical properties of gases.

Ans: **Some of the typical properties of gases are as follow:**

i. Indefinite Volume:

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

ii. Indefinite shape:

Gases do not have any shape but take the shape of the container.

iii. Pressure:

The molecules of a gas, while moving randomly, hit each other molecules and the walls of the container. The hitting of the molecules to the walls of the container produces pressure. In this way, gases exert pressure on the container, which increase with increase in temperature. Normal atmosphere pressure of the air is 760mm of Hg or 1 atmosphere.

iv. Compressibility and expansion:

Gases can be compressed by applying pressure because they have large empty spaces between their molecules. Similarly, gases expand by decreasing pressure.

v. Mobility:

Molecules of gases are in constant random motion because of the weak intermolecular force present between them.

vi. Diffusion:

We know that gas molecules are constantly moving so, they have the property of mixing with one another. The spontaneous mixing of the molecules of one gas with another at a given temperature and pressure is called Diffusion is inversely related to the mass of the gas. Greater the mass of the gas lower will be the diffusion.

vii. Effusion:

The escape of the gaseous molecules one by one from the container through a small hole of the molecular size is called effusion.

viii. Density:

Density is defined a mass per unit volume (m/v). It is evident from the formula that density of the gas is inversely proportional to its volume. Since, the volume occupied by a gas possesses mostly empty spaces, which increases the volume so density of a gas is very low as compared to the same amount of a liquid or a solid.

Q2. State and explain diffusion in gases.

Ans: **Diffusion:**

“The mixing of molecules of different gases by random motion and collision to form a homogenous mixture is called diffusion”. (OR)

“The flow of molecules from a region of higher concentration to a region of lower concentration is called diffusion”.

Explanation:

Diffusion is the property of a gas molecule. It occurs due to two reasons, random motion of gas molecules and lack of intermolecular force, for example if air fresher or perfume is sprayed in one part of a room it diffuses and its smell can be felt throughout the room.

Diffusion is inversely proportional to the molecular mass of the gas molecule. Greater the molecular mass of the gas molecules lesser will be its diffusion less the molecular mass of the gas molecules greater will be its diffusion.
e.g H_2 is lighter gas as compared to Helium and oxygen gas. Its molar mass is 2 a.m.u and Helium is 4 a.m.u and O_2 is 32 a.m.u. The rate of diffusion of H_2 is four times faster than O_2 gas.
Similarly the rate of diffusion of NH_3 gas is greater than that of HCl gas. Molar mass of NH_3 is 17 a.m.u and HCl is 36.5 a.m.u. NH_3 gas is lighter in mass, therefore, NH_3 molecules will be move at a faster rate and will diffuse rapidly as compared to HCl gas.

Q3. Write the typical properties of liquids.

Ans: The characteristics of a liquid state are in between the solid and the gases states.

i. Volume & Shape:

Liquid have a definite volume but no definite shape that is why it takes the shape of the container.

ii. Mobility:

The kinetic theory of matter (particle theory) says that the liquids consist of molecules which are constantly moving or in constant state of motion. This mobility of molecules depends upon 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.

iii. Diffusion:

Diffusion can also happen in liquids. This is because the particles in liquids are in continuous state of motion. The molecules of liquids 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 a little bit ink into glass of water, the ink will spread slowly throughout the water.
Diffusion in liquids is slower than the diffusion in gases because the molecules in liquids move slowly than the 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.

iv. Evaporation:

The spontaneous change of a liquid into the gaseous state is called evaporation. Liquids evaporate at all temperature depend upon on the strength of intermolecular forces, temperature and area of the liquid.

v. Vapor Pressure:

Liquids also exert pressure at different temperature. Greater the temperature more is the evaporation and hence, higher is its vapor pressure. Vapor pressure also depend upon on the strength of intermolecular forces. Stronger the forces of attraction less would be the vapour pressure and nice versa.

vi. Bioling point:

Different liquids have different boiling point depend upon an atmospheric pressure temperature and intermolecular forces. Liquids generally have high boiling point.

vii. Freezing point:

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

viii. Density:

Density is mass per unit volume. Liquid 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 thy cannot expand and have fixed volume.

Q4. State and explain Evaporation. What are factors affecting rate of evaporation?

Ans: **Evaporation:**

"The conversion of liquid state into gaseous state is called evaporation" (OR)

"The phenomena is which a liquid is converted into its vapors without externat heating is called evaporation".

Explanation:

The molecules of liquid move with different kinetic energies. The molecules of the liquid whose kinetic energies are higher, move faster and overcome the intermolecular forces. These come out of the liquid and convert to gaseous state.

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

Factors Affecting Evaporation:

a. Surface Area:

The process of evaporation takes place from the surface. Greater the surfaces 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 forces than alcohol, therefore alcohol evaporate quickly than water.

Q5. What is vapor pressure? What are the factors which effect the vapor pressure?

Ans: **Vapour Pressure:**

The pressure exerted by vapours above the liquid when rate of evaporation of the liquid become equal to the rate of condensation at a given temperature is called vapor pressure.

Explanation:

Consider a liquid in a close container. Temperature of the liquids is constant. The molecules with higher kinetic energy come to the liquid surface and start gathering above its surfaces.

This process is called evaporation.

Some the vapours lose energy and come back into the liquid. This process is called condensation, initially, the rate of condensation is less than that of evaporation. But after certain time rate of condensation become equal to the rate of evaporation and a state of dynamic equilibrium is reached. At this stage pressure exerted by the vapors of the liquid on walls of the container and surface of the liquid is called vapor pressure of that liquid.

Factor affecting vapour pressure:

The vapor pressure is independent of the amount of liquid and surface area of a liquid. It depends upon the nature of liquid, intermolecular forces, size of the molecules and temperature.

a. Nature of the liquid:

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

b. Intermolecular Forces:

Stronger the intermolecular forces in a liquid lesser will be its tendency to evaporate, so lesser will be its vapor pressure at the given temperature. For example, the vapors 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 form more vapor that the big sized liquid molecules 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 vapor pressure than decane at the same temperature.

Some compounds and their Vapor pressure at 289k

Name of the compound	Formula	Vapor pressure at 298k (25^{0}c) (mm Hg)
Chloroform	$CHCl_3$	170
Carbon tetrachloride	CCl_4	87
Water	H_2O	18
Glycerol	$C_3H_8O_3$	0.00016

ii. Temperature:

Increase in temperature of liquid, increase the average kinetic energy (K.E) of the molecules. The intermolecular distance among the molecules of the liquid increases, which cause the decrease in the intermolecular attractive forces between the liquid molecules. Hence, the rate of evaporation of the liquid increases and thus vapors pressure increases.

Q6. Define boiling point. How it depends on the nature of the liquid?

Ans: **Boiling Point:**

Boiling point is the temperature at which the vapour pressure of a liquid become equal to the atmospheric pressure and liquid start boiling.

Normal boiling point of a liquid is that temperature at which its vapour pressure is 760mm of Hg, which is equal to the 1 standard atmosphere pressure.

When a liquid is heated, the Kinetic energy of the molecule gradually increases which increases the vapour pressure. A time reaches when the vapour pressure of the liquid become equal to the atmospheric pressure and the liquid starts boiling.

Boiling Point of the liquid varies with the change in atmospheric pressure e.g in Hilly areas, water boil at temperature lower than 100^{0}C because atmospheric pressure is below 760m of Hg, at high altitude, food takes more time in cooking because of low boiling temperature.

The boiling point of the liquid depends upon the following factors

a. Nature of liquid (intermolecular force):

Boiling point greatly affect the nature of the liquids. It effect the intermolecular force of attraction present in the liquid. Liquids having strong intermolecular forces have high boiling points. More energy will be required to break the intermolecular forces and change the liquid into vapors form. Liquid having weak intermolecular forces have low boiling points. Low energy is required to convert the liquid into vapors form.

E.g. Boiling point of water is greater than that of Hydrogen Sulphide. Hydrogen sulphide have weak intermolecular force. Similarly, boiling point of Ethyl Alcohol is greater thatn that of diethyl ether.

b. External Pressure:

Boiling point of liquids depend on the external pressure. With increase in atmospheric, the boiling point also increases. Similarly, the decrease in atmospheric pressure causes decrease in boiling point.

At high altitude, the boiling point of a liquid will be less because the atmospheric pressure is lower. Hence, water boils lower than 100^0C because the atmospheric pressure is less than 760mm of Hg. The vacuum distillation is based on the decrease in external pressure. While with the increase in external pressure, the boiling point of liquid also increases. The pressure cooker works on this principle.

Q7. Write the typical properties of solids.

Ans: The characteristics of solid are as follows:

ii. Rigidity:

The solids possess the property of rigidity i.e. they resist the deforming forces due to hard structure and strong intermolecular forces.

iii. Density:

Density is defined d = m/v as intermolecular forces in solids are strong hence their molecules / particles are close to each other and their mass per unit volume is greater, so solids have greater density as compared to liquids and gases.

iv. Melting Point:

The temperature at which the solids starts changing into liquids is known as melting point. Melting point does not change and remains constant constant until the whole solid changes into the liquid state. Melting point is the characteristic property of a crystalline solid by which its purity can be checked.

Q8. What is Allotropy? Why elements show allotropy? Give example.

Ans: **Allotropy:**

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

The difference forms are called allotropic forms.

The allotropy of an element is due to following reasons.

i. The existence of two or 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) due to different arrangement of molecules (S_8) in the crystals.

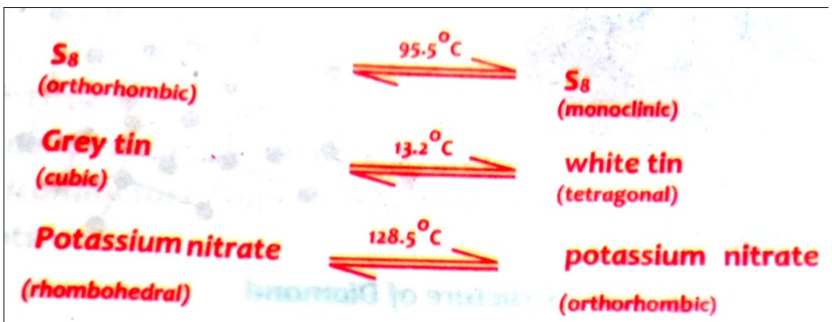
Explanation:

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

The temperature at which one allotropes 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 95.5^0C with monoclinic form.



PHYSICAL STATES OF MATTER

(Short Question Answers)

Q1. Can you reason why it take takes longer time to cook at high altitude?

Ans: Cooking is usually done in water which has a boiling of 100^0C. The boiling point is directly Proportional to external pressure. At high altitude the external pressure becomes lower. Therefore, at high altitudes the water boils below 100^0C due to which the heat is lost in the form of vapors because in boiling the heat supplied is used to convert the molecules of liquid to gaseous states that is why it takes a long time to cook at high altitude.

Q2. Glass softens over at wide range of temperature. Ice melts at a specific temperature. Explain the reason for this difference.

Ans: As we know that ice is crystalline solid. Its molecules (particles) are arranged in a regular form (three dimensional form) and have fixed position. So it has a sharp melting point and melts over a specific temperature (melting point is 0^0C or 273k).

While, glass is an amorphous substance. Its particles are not arranged in regular form and have no fixed position. So it does not has a sharp melting point and melts over a wide range of temperature (melting point is 1400^0C to 1600^0C).

Q3. Explain why it happens that on a hot summer day when there is sweet on the body of a person, one feels cool under fast moving fan?

Ans: **Reason:**

In hot summer when sweat comes on the body then it starts changing into vapors. During this process, it takes the heat of the body. The fast moving fan increase the rate of evaporation which evaporate heat of the body more quickly and thus the body becomes cool. That is why in hot summer days, when there is sweat on the body of a person, he felt cool under fast moving fan, in hot summer day.

Q4. Why are the densities of gases lower than that of liquids?

Ans: In gases, the molecules are far apart from each other having large empty spaces and occpy large space and thus showing large volume. It is evident from formula the formula that $d\alpha\frac{1}{x}$ shows that volume is inversely proportional to density. Therefore, gases have low densities. While in liquids, the molecules are close having less empty spaces than gases and have less volume. Therefore, liquids have greater densities than gases.

Example: The density of gaseous oxygen has the density of $0.00142g/cm^3$ at -0^0C and 1 atmospheric pressure.

Q5. What is the relationship between the atmospheric pressure and boiling point of a liquid?

Ans: There is directly relationship between atmospheric pressure and boiling point, With increase in Atmospheric pressure, the boiling point also increases. Similarly, the decrease in atmospheric pressure causes decrease in boiling point.

At high altitude, the boiling point of a liquid will be less because the atmospheric pressure is lower. Hence, water boils lower than 100^0C because the atmospheric pressure is less than 760mm of Hg. While with the increase in external pressure. The boiling point of liquid also increases. The pressure cooker works on this principle.

Q6. Why a gas is compressible but a solid is not compressible? Give reason.

Ans: In gases, the molecules are far apart from each other having large empty spaces. Therefore, a gas can easily be compressed. On the other hand, in solids the molecules are tightly packed together having no space between them. Therefore, a solid cannot be compressed.

PHYSICAL STATES OF MATTER
(Long Questions Answers)

Q1. Define Boyle's law and verify it experimentally.

Ans: **Boyle's Law:**

Robert Boyle, in 1662 give a relationship between pressure and volume of a given mass of a gas at a constant temperature.

Statement:

According to Boyle's Law:

"The volume of a given mass of a gass is inversely proportional to its pressure if the temperature of the gas is kept constant".

Explanation:

According to this law, Volume (V) of a given mass of a gas decrease with the increase of pressure (P). Mathematically, it is represented as

$v\ a \frac{1}{p}$ (at constant temperature)

$V = K_b \frac{1}{p}$ (where k is proportionality constant) (OR)

$PV = K_b$

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 is written as,

$P_1V_1 = P_2V_2 = K_b$

Where P_1 = initial pressure P_2 = final pressure

V_1 = initial volume V_2 = final volume

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

Experimental Verification:

The apparatus used for the experimental verification of Boyle's law as shown in the figure. A certain mass 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 temperature and it is found according to the Boyle's law.

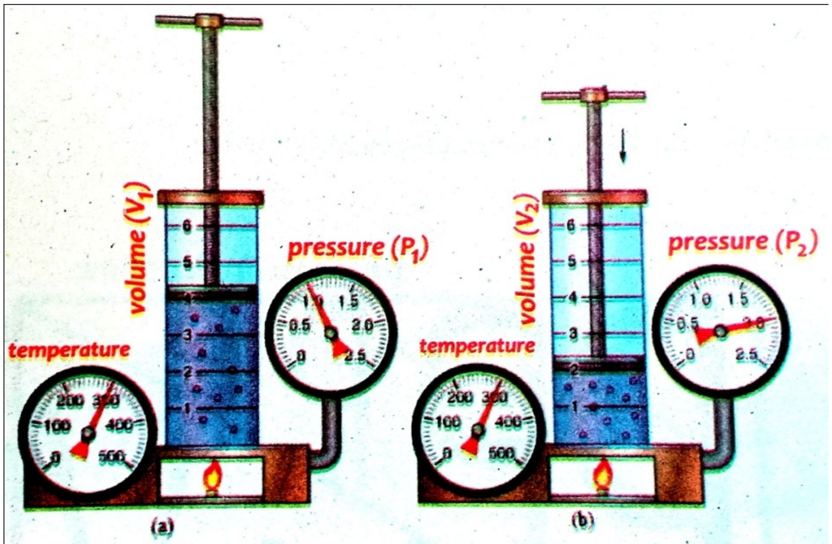


Fig. 5.4 Experimental Verification of Boyle's Law

Graphical respresentation of Boyle's law:

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

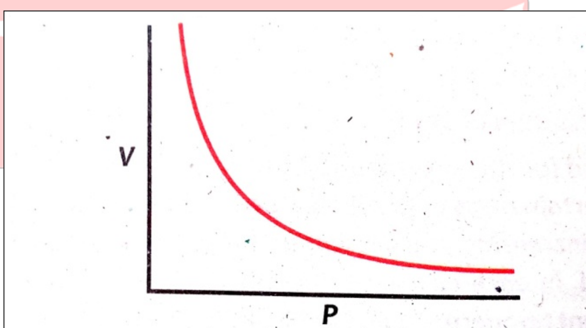


Fig. 5.5 Graphic Representation of Boyle's Law

Diagrammatic Representation:

The Boyle's law can be digrammatically explained as,

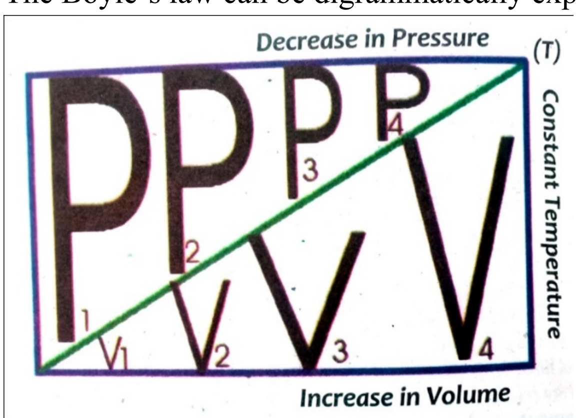


Fig. 5.6 Diagrammatic Representation of Boyle's Law

Q2. Differentiate between.

a. Evaporation and Boiling point.
b. Effusion and Diffusion of gases.
c. Condensation and Evaporation.

Ans: **a. Difference between Evaporation and Boiling point.**

EVAPORATION	BOILING
1. The spontaneous conversation of liquids into vapors is known as evaporation.	1. The conversion of liquids into vapors, when vapor pressure of the liquid become equal to the atmospheric pressure is called boiling point.
2. Evaporation occurs at all temperature evenly at 0^0C.	2. Boiling cannot occur at all temperature except 100^0C under 1 atm pressure.
3. In evaporation the energy of the molecule in not equally distributed.	3. Boiling point depend upon the atmospheric pressure and intermolecular forces.
4. Evaporation depend upon the surface area, temperature and intermolecular forces of attraction.	4. Boiling does not depend upon the surface area.
5. Evaporation occurs very slowly.	5. Boiling occurs rapidly at high rate.
6. Evaporation occurs without bubbling.	6. Boiling occurs with bubbling
7. Evaporation occurs from the surface of a liquid.	7. Boiling occurs inside the whole liquid.

b. Difference between diffusion and effusion of gases.

DIFFUSION	EFFUSION
1. In diffusion, different gas molecules are mixed with each other.	1. In effusion, gas molecules are escape or separated from each other.
2. In diffusion, molecules randomly collide with each other	2. In effusion, gas molecules move in one direction to the hole side.
3. In diffusion, more than one gas molecules move freely without any membrane.	3. In effusion, one kind of gas molecules move through a membrane.
4. In diffusion, molecules move from higher concentration region to the lower conce	4. In effusion, molecules move from a higher pressure region to a lower pressure region.

c. Differences between Condensation and Evaporation.

EVAPORATION	CONDENSATION
The spontaneous change of liquid state into gaseous state is called evaporation.	The spontaneous change of vapors back into liquid is called condensation.
Evaporation is an endothermic process.	Condensation is an exothermic process.
Evaporation involves breaking of intermolecular process.	Condensation involves formation intermolecular forces.

Evaporation is speed up by heating.	Condensation is speed up by cooling.
Evaporation takes place at all temperature	Condensation takes place only at due point.

Q3. Define the term allotropy with examples. Explain the three allotropes forms of carbon in detail.

Ans: **Allotrophy:**

The existence of an element in more than one crystalline or physical forms are called allotropes and the phenomena is called allotropy. All the allotropes have same chemical properties due to the same kind of atoms but have different physical properties due to different arrangement of atoms and different geometrical form. Different elements show allotropy e.g. carbon exist in there allotropic forms.

(a) Crystalline forms:

i. Diamond ii. Graphite iii. Bucky Balls

(b) Amophous or Non-Crystalline Forms:

i. Lamp Black ii. Coal iii. Coke
iv. Charcoal v. Gas-Carbon

Similarly, Sulphur exist in three allotropic forms.

i. Rhombic Sulphur (Octahedral Sulphur)
ii. Monoclinic Sulphur (Prismatic Sulphur)
iii. Plastic Sulphur

Allotropic form of carbon:

Carbon exists in three allotropic forms which are Diomond, graphite and bucky balls.

Diamond:

Diamond is the first and hardest allotropic form of carbon. In diamond each carbon atom is covalently bounded to four other carbon atoms making a tetrahedral geometry. In turn each carbon-atom is covalently bonded to four other carbon atoms to form interlocked structure in three dimensional form making a tetrahedral geometry.

Structure in three dimensional form making a tetrahedral geometry.

As the covalent bond in Diamond is very strong and the atoms occupy fixed position having no free space between the atoms therefore, diamond is hard. It is also non-conductor of electricity due to absence of free electrons.

Physical States of Matter

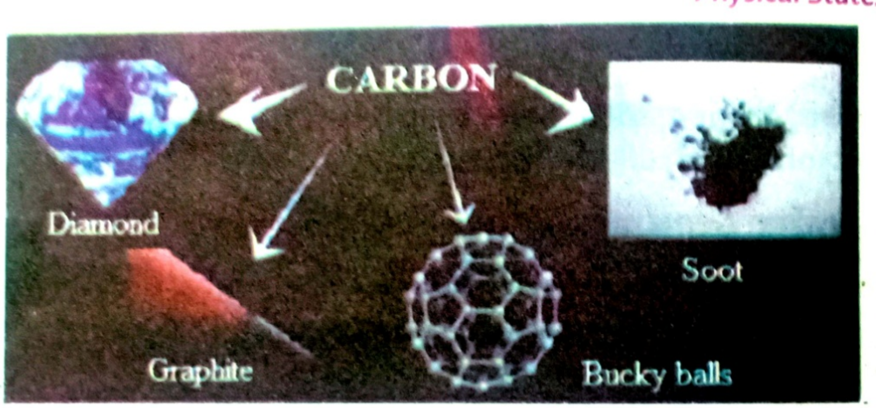


Fig. 5.17 Allotropes of Carbon

Graphite:

It is black, soft and 2nd allotropic form of carbon. It is also called plumage or black lead. It is formed naturally and also prepared artificially.

In graphite, each carbon atom is covalently bonded to three other carbon-atoms making a hexagonal structure. In turn each carbon atom is attached to three other carbon atoms making sheets. These sheets are bonded to each other through weak forces known as vander wall forces. Therefore, they can slide over one another that is why graphite is soft furthermore, in graphite each carbon atom has one free electron due to which it is good conductor of electricity.

Uses of graphite:

i. It is used as lubricant.
ii. It is used as electrode (anode)
iii. It is sued in lead pencil

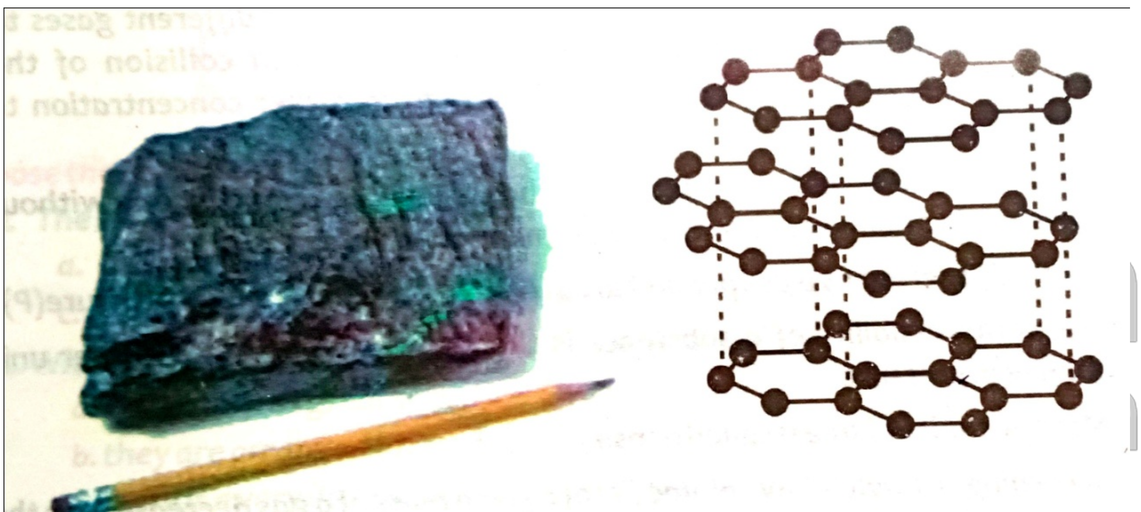


Fig. 5.19 Structure of Graphite

Bucky Balls:

Bucky Balls is third allotropic form of carbon, discovered in 1985. Their molecule contains 20 to 100 carbon atoms. These atoms are arranged in a shallow cage like structure, called Bucky balls. The simplest of these has a formula C_{60}. These carbon atoms are joined together making pentagonal and hexagonal structure.

Bucky Balls are used in a semi-conductor, super-conductors and as a lubricant.

Q4. What are solids? Differentiate between amorphous and crystalline solids.

Ans: **<u>Solids:</u>**

That state of matter which has definite shape and volume is called solids.

Explanation:

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

Types of Solids:

Solid are of two types depending upon the arrangement of the particles.

i. Crystalline Solids ii. Amorphous Solids

i. Crystalline Solids:

Crystalline solids are the solids in which particles (ions, atoms and molecules) are arranging in a regular pattern in three dimensions. Pure crystalline solids have sharp melting points.

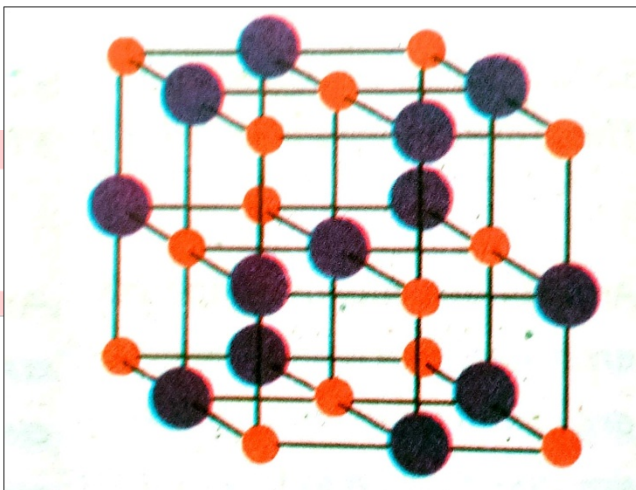


Fig. 5.16 Crystalline structure of NaCl

For example sodium chloride and naphthalene etc.

ii. Amorphous Solids:

Amorphous solids are the solids in which the particles are not properly arrange in three dimensions. They are also called non-crystalline solids. They are hard like true solids but they do not have sharp melting points rather they melt over a range of temperature.

For example

Glass, Waxes and plastics

Q5. Define Charles's Law and verify it graphically and diagrammatically.

Ans: **Charles Law:**

This law was presented in 1787 by two French Scientists Jac Tacques Charles measured the expansion of volumes of gases with rise in temperature. This law give a relationship between volume and temperature at a constant pressure.

Statement:

"The volume of given mass of a gas is directly proportional to its absolute temperature at a constant pressure".

Mathematically, this law can be expressed as:

$V \alpha T$ (at constant pressure)

= V = KcT (where K is the proportionality constant)

(OR) $\frac{V}{7} = Kc$

Where K 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 the relationship can be written in the following form:

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

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 gas is constant at constant pressure.

Experimental Verification:

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 the Charles's law.

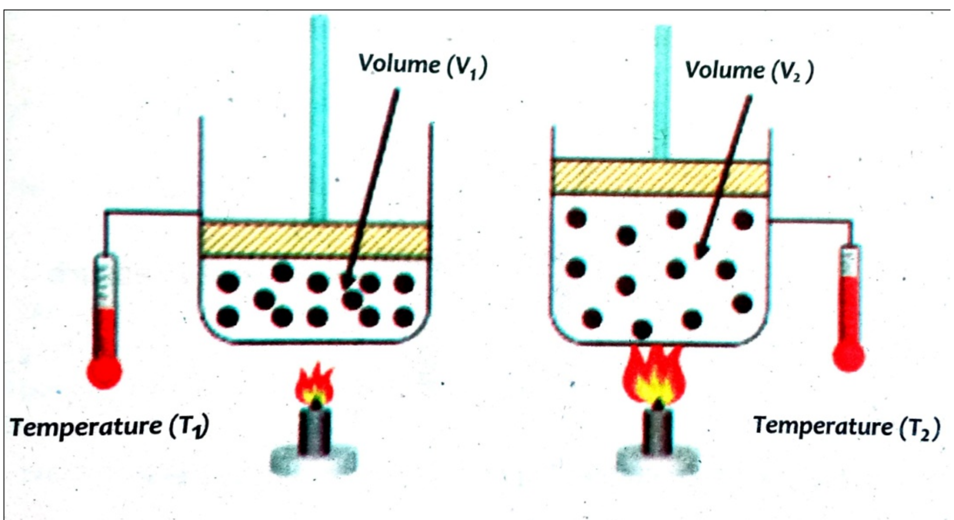


Fig 5.7 Experimental Verification of Charles's Law

Graphical Representation of Charles's law:

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 absolute temperature.

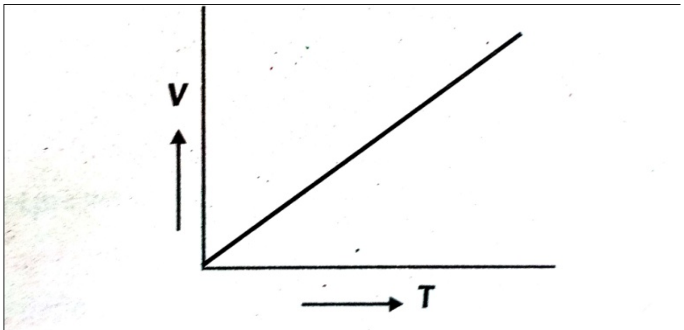


Fig 5.8 Graphic Representation of Charles's Law

Diagrammatic Representation:

The Charles's law can be diagrammatically represented as,

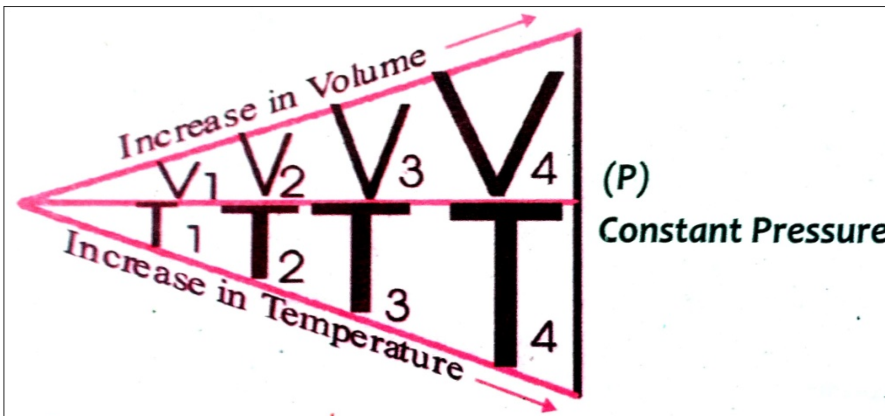


Fig 5.9 Diagrammatic Representation of Charles's law

PHYSICAL STATES OF MATTER

(Numerical Problems)

Q1. Calculate the final pressure of the sample of gas that is changed at constant temperature to 14.3 dm^3 from 7.5 dm^3 at 828 torr.

Ans: **Given Data:**

Initial volume = V_1 = 7.55 dm^3

Final volume = V_2 = 14.3 dm^3

Initial pressure = P_1 = 828 torr

Required Data:

Final pressure = P_2 = ?

According to formula:

$P_1V_1 = P_2V_2$

Re-arraning:

$P_2 = P_1V_1/V_2$

P_2 = 828 torr x 7.55 dm^3 / 14.3 dm^3

P_2 = 6251.4 / 14.3

Hence final pressure = P2 = 437.160 torr.

Q2. Calculate the final volume at 302 k of 5.41 dm^3 sample of gas originally at 353 k if the pressure does not change.

Ans: **Given Data:**

Initial volume = V_2 = 5.41 dm^3

Initial temperature = T_1 = 302k

Final temperature = T_2 = 353k

Required Data:

Final volume = V_2 = ?

Rearranging:

V_2 = V2 T2/T1

V_2 = 5.41 dm3 x 353 k / 302k

V_2 = 1909.73 / 302

Hence Final volume = V_2 = 6.323 dm^3

Q3. Calculate the initial volume of 0^0C of a sample of gas that is changed to 731 cm^3 by cooling to -140^0C at constant pressure.

Ans: **Given Data:**

Final volume = V_2 = 731 cm3

Initial temperature = T_1 = 0^0C = 0^0C + 273 = 273 k

Final temperature = T_2 = -14^0C = -14^0C + 273 = 259 k

Required Data:

Initial volume = V_1 = ?

According to formula:

$V_2/T_1 = V_2/T_2$

Rearranging:

$V_1 = V_2\ T/T_2$

$$\frac{V1}{T1} = \frac{V2}{T2}$$

➡

$$V1 = \frac{V2T1}{T2}$$

V_1 = 731 cm^3 x 273 k / 259 k

V_1 = 1909.73/302

Hence initial volume = V_1 = 770.513 cm^3

Q4. 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: **Given data:**

Initial volume: V_1 = 0.80 dm^3

Initial pressure = P_1 = 1.5 atm

Final pressure = P_2 = 2.1 atm

Required Data:

Final volume = V_2 = ?

According to formula

$P_1V_1 = P_2V_2$

Rearranging:

$V_2 = P_1V_1/P_2$

V_2 = 1.5 atm x 0.80 dm^3 / 2.1 atm

V_2 = 1.2 / 2.1

Hence final volume = V_2 = 0.571 dm^3

Q5. Calculate the final volume of 319^{0}C of a sample of gas 5.13 dm^3 at 171^{0}C, if the pressure does not change.

Ans: Initial volume = V_1 = 5.13 dm^3

Initial temperature = T_1 = 3190C = 3190C + 273 = 592 k

Final temperature = T_2 = 1710C = 171C + 273 = 444 k

Required Data:

$V_1 / T_1 = V_2 / T_2$

Rearranging:

$V_2 = V_1 T_2/T_1$

V_2 = 5.13 dm^3 x 444 k / 592 k

V_2 = 2277.72/592

Hence final volume = V_2 = 3.84dm^3 = 4 dm^3

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CHEMISTRY

Class 9th (KPK)

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SOLUTIONS
(Topic Wise Questions)

Q1. Write note on:

(a) Solution **(b) Solute** **(c) Solvent**

Ans: **Solution:**

A solution is a homogenous mixture of two or more than two substances the composition of which may be varies within definite limits. In solution the composition and appearance of all parts are uniform and there are no visible boundaries between the components of solution particles. The components of the solution are atoms, ions and molecules. Solution are existing in three physical states in solid, liquid and in gaseous form.

Example:

i. Gas solution: Air ii. Liquid Solution: Sea water iii. Solid Solution: Brass

Solute:

The component present in smaller proportion in a solution and dissolved in a solution is called solute. It may exist in three physical states solid, liquid and gaseous form.

Examples:

Solid solute = in steel carbon is a solute
Liquid solute = in sea water salt is solute
Gas solute = in air O_2 and CO_2 are solute

Solvent:

The component present in the larger proportion in a solution and dissolved the solute is called solvent. It may also exist in three physical state solid, liquid and gas.
Solid solvent = in steel iron is a solvent
Liquid solvent = in sea water, water is a solvent
Gas solvent = in air N_2 are solvent.

Q2. Define the following:

i. Aqueous Solution **ii. Binary Solution**
iii. Dilute Solution **iv. Concentrated Solution**

Ans: **Aqueous Solution:**

The word aqueous is taken from "aqua" aqua means water. A solution in which water is used as a solvent is called aqueous solution. In short cut way it is represented by (aq)

Examples:

Aqueous solution of Nacl is represented as NaCl (aq) similarly H_2SO_4 (aq), HCl (aq), NaOH (aq) etc.

Binary Solution:

Binary means two a solution which has two components one is solute and the other is solvent.

Examples:

i. NaCl + Water ii. Sugar + Water

Dilute Solution:

A solution in which less amount of solute is dissolved in a definite amount of solvent is called dilute solution. In a short cut way, it is represented by "dil"

Examples:

Dilute solution of NaCl is represented as "NaCl (dil)"
Similarly, H_2SO_4 (dil) etc.

Concentrated Solution:

A solution in which more amount of solute is dissolved in a definite amount of solvent is called concentrated solution.
Hence in a short cut way it is represented by "conc"

Examples:

Concentrated solution of NaCl is represented as NaCl (Conc)
H_2SO_4 (conc), HCl (conc) etc.

Q3. What is a solute-solvent interaction? Explain these interactions on the basis of solute and solvents.

Ans: **Solute-solvent interaction:**

There are forces of interaction between the molecules or ions of solute and also between the solute and solvent molecules. These types of interactions are called solute-solvent interaction.
Lithium chloride is highly soluble in water but gasoline is not. On other hand, gasoline mixes readily with benzene, but lithium chloride does not. Why are these such differences in solubility? Like dissolve like is a rough but useful for predicting whether one substance will dissolve in another or not.

Types on the basis of solute and solvent

i. Dissolving ionic compounds in polar solution:

The polarity of water molecules plays an important role in the formation 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 other ions in the solution.
For, example we add a few crystals of sodium chloride (NaCl) into a beaker of water. The water molecules come into contact with Na and Cl ions. The positive ends of the water molecules are attracted to Cl ions, while the negative ends area attracted to Na ions. The attraction between water molecules and ions is so strong enough to draw the ions away from the crystal and form solution. 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 solution, other ions are exposed and drawn away from the crystal surface by the solvent. The entire crystal gradually dissolves and hydrated ions.

ii. Dissolving ionic compounds in non-polar solvents:

Ionic compounds are generally not soluble 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 solute and solvents:

a. Liquid solute and solvents that are not soluble in each other are immiscible.
For example, benzene and water are immiscible and the components of this system exists in two distinct phases.
b. Non-polar substances, such as fats, oils and greases are generally quite soluble in non-polar liquids, such as carbon tetrachloride and gasoline.
Liquids that dissolve freely in on another in any proportion are said to be completely miscible.
Benzene and carbon tetrachloride are completely miscible. The non-polar molecules of these substances expert no strong forces of attraction or repulsion and the molecules mix freely with one another.
c. Ethanol and water 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 in the pure liquids are mutually soluble in all proportion. The components of this system exist in a single phase with a uniform arrangement. Hydrogen between the solute and solvent enhances the solubility of ethanol in water

Q4. Define and explain with example.
i. Suspension ii. Colloidal Solution

Ans: **Suspension (Turbidity):**

It is heterogeneous mixture of two or more than two different substance.

Explanation:

In suspension the solute particles do not dissolved in a solvent and remain suspended into the liquid (solvent) and can be seen with naked eye without the help of an instrument.

Its composition is not uniform throughout the solution.

The particles (solute) present in a solvent cannot pass through the filter paper and stay behind on the filter paper during filtration. They settle down at the bottom of the container it allowed to stand for some time. Suspension are the result of solute which do not dissolve in true sense and remain suspended or settle down at bottom and the solute is not uniform.

Examples:

A mixture of chalk in water, mud in water etc.

Suspension among the medicine like antibiotic, disiprin and paracetamol are the examples of suspension.

Colloids or colloidal solutions:

A solution in which solute particles are bigger in size than the solute particles of a solution but smaller than the solute particles of suspension is called colloid or colloidal solution. In colloids, solute particles are not homogenized with the solvent.

Explanation:

These are a little but bigger than the solute particles of a solution but not so bigger that can be seen by naked eye. Particles between 1nm-1000nm in diameter may form colloids. These can pass through the filter paper during filtration. They do not settle down at the bottom of the container of allowed to stand for some time.

A mixture of starch in water, milk, fog, smoke, blood and water-soluble paints.

CLASS OF COLLOID	PHASES	EXAMPLE
Sol	Solid dispersed in liquid	Paints, mud
Gel	Solid network extending throughout liquid	Gelatin
Liquid Emulsion	Liquid dispersed in liquid	Milk, Mayonnaise
Foam	Gas dispersed in gas	Smoke, airborne particles, matter, exhaust
Liquid aerosol	Liquid dispersed in gas	Fog, mist, clouds, aerosol, spray
Solid emulsion	Liquid dispersed in solid	Cheese, butter

Q5. Write the properties of solution, colloids and suspensions.

Ans: **Properties of solution, colloids and suspension:**

S.No	Solution	Colloids	Suspension
1	Homogenous	Heterogeneous	Heterogeneous
2	Their particle size is in between 0.01 – 1 nm. It can be atoms, ions or molecules.	Their particle size in between 1 – 1000 nm and are dispersed. It can be aggregates, or large molecules.	Their particles size is over 1000 nm and are suspended. It can be large particles or aggregates.

3	They cannot be separated on standing for long time	They cannot be separated on standing for ling timing.	Their particles are settling down on standing.
4	They cannot be separated by filtration.	They cannot be separated by filtration	They can be separated by filtration
5	They do not scatter light	They scatter light showing the Tyndall effect	They may scatter light, but are not transparent
6	Their particles are so small that they can't be seen with naked eye	Their particles are big but can't be seen with naked eye.	Their particles are big enough to be seen with naked eye.
7	e.g. table salt in water	e.g. milk in water	e.g. flour in water

SOLUTIONS

(Long Questions Answers)

Q1. Define solution? Explain types of solution on the basis of states of matter.

Ans: **Solution:**

A solution is a homogenous mixture of two or more than two substances the composition of which may be varies within definite limits. In solution the composition and appearance of all parts are uniform and there are no visible boundaries between the components of solution particles. The components of the solution are atoms, ions and molecules. Solution are existing in three physical states in solid, liquid and in gaseous form.

Example:

i. Gas solution: Air
ii. Liquid solution: sea water
iii. Solid solution: Steel, Brass

Types of solution on the basis of states of matter:

Solution exists in gas, liquid and solid states. On this basis there are nine classes of solutions.

Types	Examples	Solute	Solvent
Gas solutions Gas in gas Solid in gas	Air Fog Smoke	O_2 Water vapour Carbon particle	N_2 Air Air
Liquid solutions Liquid in liquid Gas in liquid Solid in liquid	Vinegar HCl Solution Sea water	Acetic acid HCl Gas NaCl	Water Water Water
Solid Solutions Solid in Solid Gas in Solid Liquid in Solid	Brass Hydrogen absorbed Palladium palate Dental amalgam	Zinc H_2 Gas Mercury (Hg)	Copper Palladium Silver (Ag)

Q2. a. Discuss the solubility of a substance?

Ans: **a. Solubility:**

The amount of solute in grams required to saturate 100 gm of the solvent at a particular temperature is called solubility.

(OR) The maximum amount of solute in gram required to saturate 100gm of the solvent at a particular temperature is called solubility.

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 a solute is taken to be the quantity required to make a saturated solution in a given quantity of the solvent.

Solubility = $\frac{wt\ of\ solute}{wt\ of\ solvent}$ x 100

Q2. b. Explain the factors that are responsible for the solubility of a substance?

Ans: **b. Factors affecting solubility:**

i. Temperature
ii. Pressure
iii. Nature of the solute
iv. Nature of the solvent

Temperature:

The solubility of many substances is affected with temperature. Increasing the temperature, usually decrease gases solubility. As the temperature increases, the average kinetic energy of the molecules in

the solution increase. A greater number of solute molecules escape from the attraction of solvents molecules and return to gas phase. At higher temperature, gases are generally less soluble.
The effect of temperature on the solubility of solids and liquids is more difficult to predict. It is often observed that solubility of many solutes in solution generally increases with the increase in temperature, but this not always happens. When a solute is added into solvent, there are different possibilities with respect to temperature.

These possibilities are given below:

i. The solubility of some solutes in solution generally increases with the increase 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 also with the 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. The solubility of some solids increases up to a certain temperature and then decrease (Na_2SO_4. $10H_2O$) at 32.4^0C. Above this temperature, it forms anhydrous Na_2SO_4. The maximum solubility of Sodium Sulphate is 32.4^0C.

Pressure:

Since 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.

Example:

CO_2 is filled in soda water bottles under 4 atmospheric pressure. When a bottle of soda water is opened, CO_2, comes out with effervescence (bubbles) because pressure in the bottle is released and reduced to 1 atm resulting in decrease in the solubility of the gas.

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 in a non-polar solvent like benzene.

Nature of Solute:

It also affects the solubility, if a solute is changed in the same solvent, the solubility changes. If a solute is changed and solvent remain the same, the solubility also changes.
For example, sodium chloride has high solubility in water and sugar has comparatively low solubility.
“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 compound are soluble in water.
For example, Nacl, Kcl, Na_2CO_3, sugar, glucose and alcohol are soluble in water.
ii. Non-polar substances are not soluble 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.

Q3. a. What is the difference between a concentrated and dilute solutions? Give example of each.

Ans: **a. Difference between a concentrated and dilute solutions:**

Dilute Solution:

A solution in which less amount of solute is dissolved in a definite amount of solvent is called dilute solution. In a short cut way it is represented by “dil”.

Example:

Dilute solution of NaCl is represented as “NaCl (dil)”.

Similarly, H_2SO_4 (dil) etc.

Concentrated Solution:

A solution in which more amount of solute is dissolved in a definite amount of solvent is called concentrated solution.

Hence in a short cut way it is represented by "conc".

Example:

Concentrated solution of NaCl is represented as "NaCl (conc)"

H_2SO_4, (Conc), HCl (Conc) etc.

Q.3.b Differentiate between unsaturated, saturated and supersaturated solutions.

Ans: **b. Saturated Solution:**

A solution which contain maximum amount of solute and cannot dissolve any more solute at a particular temperature is called saturated solution.

Explanation:

If more of the solute is added in it, it remains un-dissolved and will be settle down at the bottom of the container at a room temperature.

Example:

Take some water in a beaker and add NaCl in small amount at interval with constant stirring. If more quantity is added it remain insoluble. The insoluble salt will be settle at the bottom of the beaker at room temperature and water is incapable to dissolve more salt, such a solution is known as saturated solution.

Unsaturated Solution:

A solution which can dissolve further amount of a solute and have the capacity to dissolve more solute at a particular temperature is called as unsaturated solution.

Example:

Take a glass half filled with water. Add a spoon of table salt (NaCl) in it. It will dissolve. Add another spoon of salt in it. It will also dissolve in it. Such a solution is unsaturated because it can still dissolve more amount of solute in it at a particular temperature.

Super Saturated Solution:

A solution which contain more amount of solute than a saturated solution at a given temperature is called super saturated solution. (OR)

A solution which is more concentrated than a saturated solution is known as a super saturated solution. Such a solution is unable to dissolve more solute and the excess solute is separated out in the form of solid particles or crystals.

Example:

Fill half a test tube water and add sufficient crystals of $Na_2S_2O_3$ (Sodium thio sulphates) in it. Cool the test tube water under a tap water without shaking. A super saturated solution is formed having no crystals. Another example are custard, curd and kheer.

Q4. Describe one way to prove that a mixture of sugar and water is solution and that a mixture of sand and water is not a solution.

Ans: Mixing a little amount of sand in the glass of water, while the sand will get mixed in the Water it will not dissolve and will form a heterogeneous mixture. Now according to definition, solution is a homogeneous mixture which has uniform composition in which solute particles cannot be seen by naked eyes. Hence mixture of sand in water is not a solution because:

i. It forms heterogeneous mixture in which components are not in uniform composition in which sand is present in the bottom of glass.

ii. Sand particles can be seen in glass of water by naked eyes.

iii. Sand can be easily filtered by simple filtration while the solute particles in solution cannot be filtered by simple filtration.

Now mixing of a little amount of sugar in glass of water, When the sugar is mixed into the glass, it dissolved completely, which means that the individual particles of sugar cannot be seen in water. Hence the mixture of sugar in water is a solution because:
i. It forms a homogenous mixture which has uniform composition.
ii. The sugar cannot be seen in glass of water by naked eyes.
iii. Sugar cannot be separated from water easily by means of simple mechanical methods. The only separation method is mostly used is distillation.

Q5. Explain the following concentration units.
a. Percentage composition **b. Molarity**

Ans: **a. Percentage composition:**
The percentage composition of a solution can be expressed as:

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 by mass means, 10 gram of sugar inn 90 grams of solvent, so that solution weight 100 grams.

%mass / mass = $\frac{mass\ of\ solute\ (g)}{Mass\ of\ solute\ (g) + mass\ of\ solvent\ (g)}$ x 100

%mass / mass = $\frac{mass\ of\ solute\ (g)}{Mass\ of\ solution\ (g)}$ x 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 in solvent to make 100 cm^3 of the solution. In this case, the total mass of the solution is not considered.

% volume / mss = $\frac{mass\ of\ solute\ (g)}{volume\ of\ solution\ (cm)^3} \times 100$

iii. Percentage volume by mass (volume – mass relationship v/m):
It is the volume in cm^3 of a solute dissolved in 100 grams 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 in 100 grams of solvent. In this case the total volume of the solution is not considered.

% volume / mass = $\frac{volume\ of\ solute\ (cm^3)}{volume\ of\ solution\ (g)}$ x 100

iv. Percentage volume by volume (volume – volume Relationship v / v):
It is the volume in cm^3 of 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 volume of solution is 100 cm^3.

% volume / volume = $\frac{volume\ of\ solute\ (cm^3)}{volume\ of\ solution\ (cm^3)}$ x 100

b. Molarity:
Molarity is the number of moles of solute present in one dm^3 of the solution. It is defined as the number of moles of solutes dissolved per dm^3 of solution. Molarity is represented by M.

M = $\frac{Number\ of\ moles\ of\ solute}{volume\ of\ solution\ in\ dm^3 or\ litre}$

By definition

Mole = $\frac{Amount\ of\ solute\ in\ gram}{Molecular\ weight\ of\ solute}$

So molarity can also be given as:

M = $\frac{Amount\ of\ solute\ in\ gram}{Molecular\ weight\ of\ solute\ x\ volume\ of\ solution\ in\ dm^3\ or\ litre}$

SOLUTIONS
(Short Questions Answers)

Q1. Is sea water a solution? How would you prove with a simple experiment whether it is pure water or solution?

Ans: Yes, sea water is a solution of salts in water. We can differentiate between pure water and solution by doing two simple experiments.

Experiment 1:

As we know that pure water is tasteless. Now simply taste a little water from sea. If it is tasteless it is pure and if it has a taste then it is a solution.

Experiment 2:

Pure water is bad conductor of electricity. Now take a sample of water from the sea, allow the current through the water. The current will passed "it means that the sea water is solution".

Q2. A bottle in a drug store contains a label "3 percent Hydrogen peroxide". What does it mean?

Ans: It shows the percentage composition of Hydrogen peroxide solution i.e. percentage volume by volume. 3 % Hydrogen peroxide by volume means 3 cm^3 of Hydrogen peroxide in sufficient volume of solvent, so that volume of solution in 100 cm^3 i.e. 3% of hydrogen peroxide dissolve in 97% of water by volume-volume composition.

Q3. Classify the following as a solution, colloids 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 v. Cheese viii. Mist ix. Bottled water

Ans:

SOLUTIONS	COLLOIDS	SUSPENSION
Listerine mouthwash bottled water	Milk Hot cup of tea Mayonnaise Milk of magnesia Cheese Mist	Orange juice with pulp

Reason:

i. Solution is homogenous mixture of two or more components. Listerine mouthwash and Bottled water are homogenous mixtures so they are solution.

ii. Colloids is solution in which particles are intermediate in size between those in solutions and suspensions. Milk, Hot cup of tea, Mayonnaise, Milk of magnesia, Cheese and Mist have larger solute particles than true solution but not enough than suspension so they are colloids.

iii. Suspension is a heterogeneous mixture of undissolved particles in a given medium. The solute particles of Orange juice with pulp can be seen with naked eyes so it is suspension.

Q4. Why we stir paints thoroughly before using it?

Ans: Paint is a colloid, made of many components like pigments, colour in solvent. The solute particles of paint are not homogenized with the solvent so, it is better to stir paints thoroughly with the stirrer to get a uniform composition. By using it will cover the surface uniformly. So that's why we stir paints before using it.

Q5. Why suspension and solution do not show Tyndall effect, while colloids do?

Ans: When light is scattered by colloidal particles dispersed in a transparent medium is called Tyndall effects.

Reason:

In case of Colloids:

In colloids solute particles are intermediate in size between those in solutions and suspensions. Hence the particles size is large comparable to that of wavelength of light. That's the light rays are scattered and produce Tyndall effect.

In case of Suspension:

In suspension the solute particles are large in size even they can be seen with naked eyes and the wavelength of the visible is smaller than the solute particles. They block the light rays instead of scattering it and hence do not shows the Tyndall effect.

In case of solution:

Solutions are the homogenous mixture of two or more substance in which the solute particles are very small than the wavelength of the visible light. Due to very small solute particles the light rays are passed in straight line without showing the Tyndall effect.

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CHEMISTRY

Class 9th (KPK)

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ELECTROCHEMISTRY
(Topic Wise Questions)

Q1. What is electro chemistry? Discuss briefly

Ans: **Electro Chemistry:**

The branch of chemistry which deals with the study of conversion of electrical into chemical energy and chemical energy into electrical energy is called electro chemistry.

Explanation:

In electro chemistry we study the chemical change take place. When electric current is pass through a particular type of material.

Electrochemistry is used in this modern world to generate electricity that is used in different electronic devices like cell in remote control, mobile phones, camera and telecommunication. In industry, many important chemical and useful products are manufactured by electrochemical methods.

Q2. What is oxidation number? What are the rules of assigning oxidation number?

Ans: **2. Oxidation State:**

The apparent charge (+ive or –ive) on an atom of an element in a molecule or compound is called oxidation state.

Explanation:

Oxidation state of an element is described by its oxidation number. This number enables us to describe oxidation-reduction reaction, and a balancing Redox chemical reaction. An increase of oxidation number is oxidation where a decrease of oxidation state is reduction.

Rules for assigning oxidation number:

i. The oxidation number of all elements in the Free state is zero

Example: N_2, Cl_2, Na, Fe, P_4 and S_8 etc. is zero

ii. The oxidation number of a simple ions is the same as the charge on it.

Examples: $Na^+ = 1$, $Al^{+3} = +3$, Ca^{2+}, $Br^{-1} = -1$

iii. The oxidation number of hydrogens in its compounds is +1 except in the case of metal hydrides, where it is -1. E.g. Na^+ , H^{-1}

Example:

In HCl^+, $H2O^+$ is +1 where as in metal hydride NaH^{-1}, CaH_2^{-1}, LiH^{-1} etc. is -1

iv. The oxidation number of oxygen in its compounds in -2 except in this case of peroxide, where it is -1 and in case of OF_2, It is +2

Examples:

In Zn^{-2} O^{-2} is -2 and in peroxide $H_2O_2^{-1}$ is -1 while in $O+2F_2^{-2}$ is +2

v. The oxidation number of each element of group I, II and III are +1, +2 and +3 respectively

Examples:

Group IA Li^{+1} is +1 Group IIA Ca^{+2} is +2 Group IIIA Al^{+3} is +3

vi. The oxidation number of each element of group VIIA (Halogen) in their binary compound is -1.

Examples: Br^{-1} is -1

vii. In neutral molecules, the algebraic sum of the oxidation number of all the element is zero.

Examples: In H^{+1} Cl^{-1} ➔ +1 -1 = 0, H^{+1} N^{+5} O^{-6} ➔ (1+5-6) = 0

viii. In ions, the algebraic sum of oxidation number is equal to the charge on the ion.

Example:

CO_3^{-2} = (+4 x 1) + (-2 x 3) = -4 -6 = -2

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.

Example: CO($C^{+2}O^{-2}$), CO_2 (C^{+4}, O_2^{-4})

Q3. What is the oxidation state of C in CO_2, O in CO_2, Sn in $SnCl_4$ and S in K_2SO_4?

Ans: i. C in CO_2

x + 2 (-2) = 0

x – 4 = 0

x = +4

ii. O in CO_2

4 + 2 (x) = 0

2x = -4 ÷ both sides on 2

$\frac{Zx}{2} = \frac{A}{2}2$

x = -2

iii. Sn in $SnCl_4$

Solution:

Oxidation No of Sn = x

Oxidation No. of Cl = -1

SnCl4

x + 4 (-1) = m0

x – 4 = 0

x - +4

Oxidation No. of Sn in $SnCl_4$ is +4

iv. S in K_2SO_4

Solution:

Oxidation No. of K = +1

Oxidation No. of S = x

Oxidation Number of O = -2

K_2SO_4

2(+1) + x + 4 (-2) = 0

2 + x -8 = 0

x -6 = 0

x = +6

Oxidation No. of S in K_2SO_4 is +6

Q4. What is oxidizing and reducing agents? Explain with examples.

Ans: **i. Oxidizing Agent:**

An oxidizing agent is the specie that oxidizes other substances and itself get reduced. For example, KMnO4, $K_2Cr_2O_7$, HNO3 and Cl2 etc.

According to 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. The oxidation number of an oxidizing agent is decreased during a redox reaction.
Consider a following reaction:
Zn (S) + CuO (S) ➔ ZnO (S) + Cu (S)
Oxidation number of Cu in CuO is +2
Oxidation number of Cu (s) is = 0
There is decrease in oxidation number Cu from +2 to 0. So, Cuo is an oxidizing agent in the given example.

ii. Reducing agent:

a reducing agent is the specie that reduces other substances and itself get oxidized.
For example, H_2S, SO_2, Na, Al and Mg etc.

According to classical concept, a reducing agent may be,

a. The acceptor of oxygen to a substance.
b. The donor of hydrogen forms a substance.
c. The donor of an electron from a substance.
d. The oxidation number of an oxidizing agent is increased during a redox reaction.
Consider a following reaction:
Br_2 (s) + H_2S (s) ➔ 2HBr (s) + S (s)
Oxidation number of S in H_2S is -2
Oxidation number of free S is 0
There is increase in oxidation number of S from -2 to 0. H_2S is a reducing agent in the given example:

OXIDIZING AGENTS	REDUCING AGENTS
Bromine (Br_2)	Carbon (C)
Chlorine (Cl_2)	Carbon monoxide (CO)
Concentrated sulfuric acid (H2SO4)	Hydrogen (H_2)
Nitric Acid (HNO_3)	Hydrogen Sulphide (H_2S)
Oxygen (O_2)	Metals
Potassium permanganate ($KMnO_4$)	Potassium iodide (Kl)
Potassium dichromate ($K_2Cr_2O_7$)	Sulphur dioxide (SO_2)

Q5. What are the oxidation and reduction reactions?

Ans: **Oxidation Reduction reactions:**

Redox reaction (Red means reduction, Oxi means oxidation)
Those chemical reactions in which Oxidation-reduction (loss and gain of electrons) takes place simultaneously are called oxidation-reactions or Redox reactions.
$CuSO_4$ (aq) + Zn (s) ➔ $ZnSO_4$ (aq) + Cu (s)
The overall ionic equation is
The Redox reactions are actually made up two half reaction, for example consider the reaction.
$2Na^0 + l_2^0$ ➔ $2Na^{+1} + 21^{-1}$
The reaction can be written in form of two half reaction as,
$2Na^0$ ➔ $2Na^{+1} + 2e^-$ (oxidation)
$l_2^0 + 2e$ ➔ 21 (reduction)

In the 1st half reaction, loss of electrons take 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.

Q6. What are electrochemical cells? Explain electrolytic cell in detail.

Ans: **Electrochemical Cells:**

A device in which inter conversion of electrical and chemical energies take place is called an electrochemical cell.

It is an energy device, in which either a chemical reaction take place by using electric current (such as electrolysis) or chemical reaction produces electric current (such as electric conductance).

Electro chemical cells are of two types.

i. Electrolytic Cell
ii. Galvanic or Voltaic cell

Electrolytic Cell:

A device in which electric current is used to produce Redox reaction.

(OR)

It is a device in which electrical energy is converted into chemical energy by non-spontaneous redox reaction.

Explanation:

In this cell electric current is produced due to the presence of potential difference. There reaction involve the gaining of electrons (reduction) and the losing of electrons (oxidation).

Construction:

An electrolytic cell consists of solution of an electrolytes. Two metallic called electrodes i.e. anode and cathode are dipped in the electrolytic solution. These 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 electrode which is connected with negative terminal of battery is called cathode.

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 form the anode. Te anions move toward anode and discharge their electron (s) there and thus oxidation take place at anode. The cations move towards cathode and gain the electron (s) there are thus oxidation takes place at cathode.

For example, when electric current passed from the fused sodium chloride (NaCl), the following reactions take place during the process.

$$2NaCl\ (l) \rightarrow 2Na^+\ (l) + 2Cl^-\ (l)$$

At anode:

$$2Cl^-\ (l) \rightarrow Cl_2\ (g) + 2e$$

At cathode:

$2Na^{+}(l) + 2e \rightarrow 2Na\,(l)$

Overall reaction:

$2Na^{+}(l) + 2Cl\,(l) \rightarrow 2Na\,(l) + Cl_2\,(g)$

Diagram:

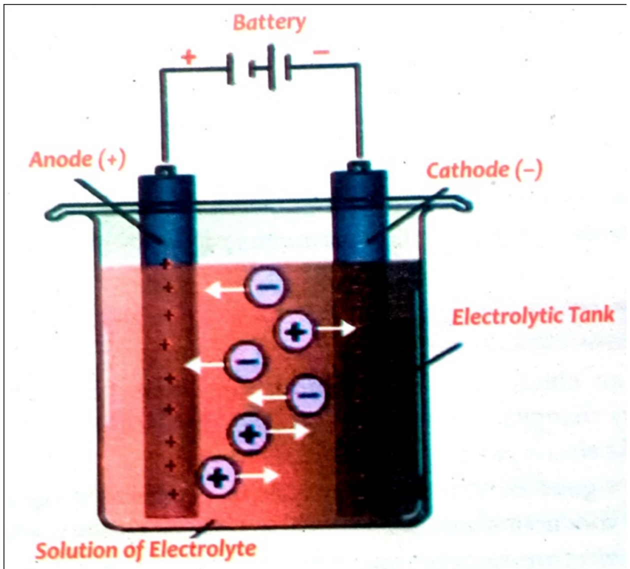


Fig. 7.4 Electrolytic Cell

Q7. What is galvanic or Voltaic cell? Explain in detail.

Ans: **Galvanic Cell:**

The cell in which the chemical energy is converted into electrical energy by spontaneous redox reaction is called voltaic cell (Galvanic cell)

Explanation:

Voltaic (Galvanic) cell consist of two separate compartments called "half cells" containing electrolytic solution and electrodes one is for oxidation and other is for reduction. The two half-cells are connected internally by a salt bridge and externally by a wire to which galvanometer or voltmeter is connected which detect or record the current. The best example of voltaic or galvanic cell in the Daniel cell.

Daniel Cell:

Construction of Daniel Cell:

A Galvanic cell consist 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. The left half-cell consists of zinc electrode dipped in 1M solution of zinc sulphate ($ZnSo_4$) and right half copper electrode dipped in 1m of $CuSo_4$ connected to a wire to an external circuit to which a galvanometer or voltmeter is attached. The solutions in different containers are connected with the bridge. This bridge is known as "salt bridge". A salt bridge is a U-shaped tube. This tube is filled with the electrolyte gel, such as K_2SO_4 or Na_2SO_4 is called as "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 the tendency to lose electron more readily than copper when the circuit is completed. As a result, oxidation take place at Zn electrode. The electron flows from Zn electrode through the external circuit to copper electrode. These electrons are gained by the copper ions of the solution at the cathode and deposited as a copper atom at the cathode.

Cell Reaction:

The flow of electrons forms one electrode to 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 cell reaction. Thus, we have half-cell reaction i.e. oxidation and reduction processes, going on two electrodes simultaneously. Electrons travel in external circuit, while ions move through the salt bridge and this way electric current is produced. These reactions are as follow:

At anode:

$Zn\ (s) \rightarrow Zn^{+}\ (aq) + 2e$

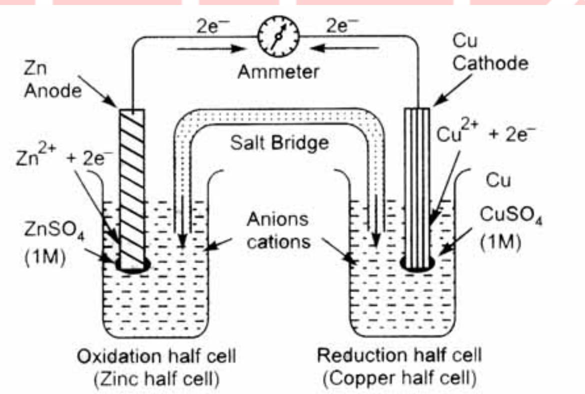
At cathode:

$Cu^{+}\ (aq) + 2e^{-} \rightarrow Cu\ (s)$

Overall reaction:

$Zn\ (s) + Cu^{+}\ (Aq) \rightarrow Zn^{+}\ (aq) + Cu\ (s)$

Diagram:



Daniell cell

Q8. Define electrolytes, Non electrolytes and electrodes.

Ans: **Electrolytes:**

The substance which dissociate into (+ive) and (-ive) ions in an aqueous solution and conduct electric current easily are called electrolytes.

Those substances which in aqueous solution or in molten state allow the electric current to pass through them are called electrolytes.

Examples:

Acids, Bases and salts,

Acids = Hcl, H_2SO_4, HNO_3, CH_3COOH

Bases = KOH, NaOH, NH_4OH

Salts = NaCl, N_4HCl etc. are electrolytes

Types of electrolytes:

Weak Electrolytes:

Those substance which ionize partially (upto limit extent) and conduct electric current poorly due to small ionization in aqueous solutions are called weak electrolytes

Examples:

NH_4OH, CH_3COOH, Ca $(OH)_2$, Fe $(OH)_3$ etc.

NH_4OH (aq) <===> NH_4 (aq) + OH (aq) (1.52%)

(Weak base)

CH_3COOH (aq) <===> CH_3COO (aq) + H^+ (aq) (1.5%)

(weak acid)

Examples:

NaOH, HCL, KOH, H2SO4 etc.

= NaOH (aq) ➔ Na+ (Aq) + OH (aq) (93%) (Strong base)

= HCl (aq) ➔ H+ (aq) + Cl (aq) (84%) (strong acid)

Strong Electrolytes:

Those substance which ionize completely (a large extent) in aqueous solution and conduct electricity to a large extent are called strong electrolytes.

Examples:

NaOH, HCL, KOH, H_2SO_4, etc.

= NaOH (aq) ➔ Na^+ (aq) + OH^- (aq) (93%) (strong base)

= HCL (aq) ➔ H^+ (Aq) + Cl^- (aq) (84%) (strong acid)

Non-electrolytes:

The substance which do not ionized in aqueous solution and do not allow the electric current to pass through them are called non-electrolytes. Most of the organic compounds are non-electrolytes.

Examples: Petrol, Sugar, Glucose, Urea Pure water etc.

Electrodes:

Electrodes are the conductor i.e. metallic plates, wires or rods through which electrons enter or leave the electrolyte in a cell are called electrodes.

There are two types of electrodes, Anode and Cathode.

a. Anode:

The anode is the positive electrode at which anion gathers and leaves the electron in the electrolytic cell.

Anions are the negatively charged particles e.g. Cl, OH etc.

b. Cathode:

The cathode is the negative electrode at which cations gathers and gains the electron in the electrolytic cell.

Cations are the positively charged particles e.g. Na^+, NH_4^+, etc.

Q9. Write the electrolytic refining of copper Cu.

Ans: **Electrolytic refining of copper Cu.**

Pure copper is very good conductor of electricity and used in electrical instruments.

Copper is purified by electro-refining.

Construction:

Large block of the blistered (impure) copper (99% pure) are suspended as anode in the copper Sulphate ($CuSO_4$) and sulphuric acid (H_2SO_4) between the thin sheets of pure copper which acts as cathode. The operation is performed at 50^0C and applied voltage of about 0.3 volts and optimum current density used is 160-400 A/m^2.

Procedure:

When the electric current is passed copper dissolves from the impure copper anode to give Cu^{+2} ions.

Reaction at anode (oxidation):

$$Cu\ (s) \rightarrow Cu^{+2}\ (aq) + 2e^-$$

At the cathode, all the cu^{+2} ions from the solution are reduced to the metallic copper and get deposited at the cathode.

Reaction at anode (oxidation):

$$Cu\ (s) \rightarrow Cu^{+2}\ (aq) + 2e^-$$

As the electrolysis continued, copper from the anode 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.

Reaction at Cathode (reduction)

$$Cu^{+2}\ (aq) + 2e^- \rightarrow Cu\ (s)$$

Diagram:

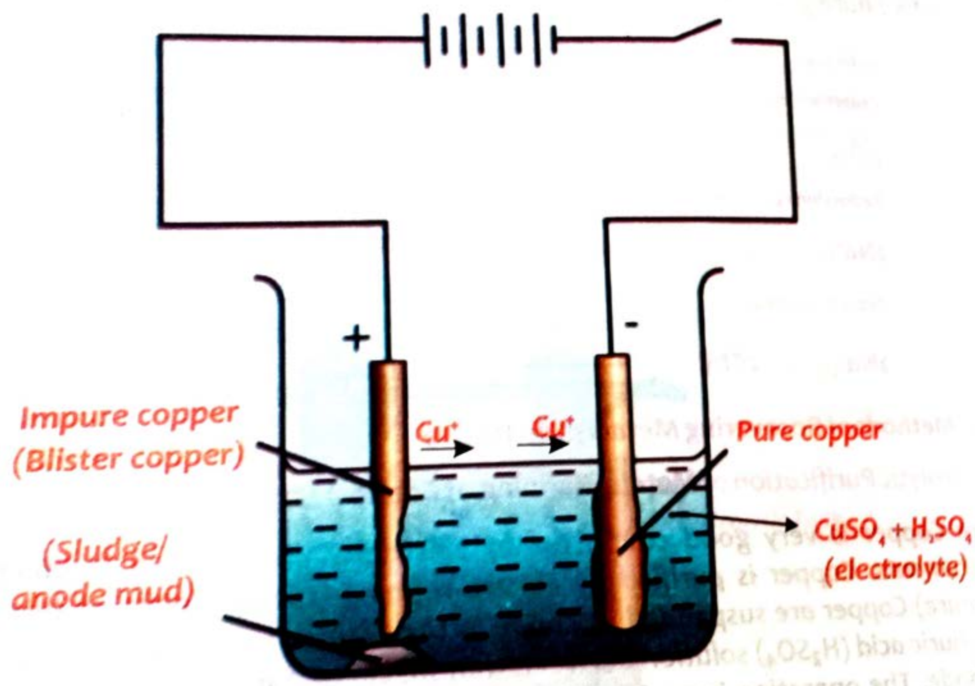


Fig. 7.5: Refining of Copper

Q10. What is battery? Explain dry cell.

Ans: **Battery:**

A group of combination of galvanic cell joined in a series is called battery.

Explanation:

A battery is a self-contained, chemical power pack than can produce a limited amount of electrical energy, whenever it is needed. It covers chemical energy into electrical energy, for specific period of time. Car batteries consist of six or more identical voltaic cells connected in series.

Dry Cell:

The dry cell was prepared by Leclanche in 1887.

Construction of dry cell:

The dry cell consists of metallic container. Its container is made up of Zinc (zn) which acts as anode. The Zinc casing is consumed during the chemical reaction. A graphite rod is placed in the center of container which acts as cathode. This container is also filled with the mixture of Ammonium chloride (NH_4Cl), Magnesium dioxide (MnO_2) and Carbon (C) which is in the form of paste. The cell is water proofed with the wax. The voltage produces by the dry cell is 1.25v.

Reaction of the Cell:

Oxidation and reduction reactions occurs in the cell to produce the electric current.

At Anode:

The zinc acts as anode in the cell. The zinc is oxidized by losing two electrons.

$Zn \rightarrow Zn^{+2} + 2e^-$

At Cathode:

The graphite acts as cathode in the cell. In the cell NH_4Cl and MnO_2 are reduced to Mn_2O_3 and NH_4.

$2MnO_2 + 2NH4 + 2e \rightarrow Mn_2O_3 + 2NH_3 + H_2O$

$2NH_4Cl \rightarrow 2NH_4^+ + 2Cl$

Overall reaction:

$Zn^{+2} + 2Cl ==. ZnCl_2$

Q11. Write a note on electrochemical industries.

Ans: **Electrochemical industries:**

Electrochemical industries are based on many electrochemical operations. Some of these are as follow:

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.

Q12. Write a note on prevention of corrosion and its techniques.

Ans: **Prevention of Corrosion:**

Corrosion can be prevented by a number of methods depending on the circumstances of 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 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 these modern paints contain a combination of chemical called stabilizers. These stabilizers provide prevention against not only corrosion but also 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 a nickel. Stainless steel strongly resists the corrosion. The development of new alloys are constantly under production.

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 on other metal by dipping) or electroplating, for example iron articles are protected from rusting by Nickel (Nil), 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 to protect iron 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. It is because the magnesium or zinc is a better reducing agent than iron, electrons are supplied by the magnesium or zinc than by iron, keeping the iron form 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 steel hull, in salt water the titanium acts as anode and is oxidized instead of the steel hull (the cathode).

Q13. Write a note on zinc plating, chrome plating and tin plating.

Ans: **a. Electroplating of Tin:**
The target metal is cleaned with caustic soda, treated with acids, in order to remove the rusts 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 acts as anode and is dipped in sodium stannate ($Na_2SnO_3.3H_2O$) 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 the cathode they discharged a deposited on the object.
The following reaction occurs:

At Anode:

$$Sn^{+2} (aq) + 2e^- \rightarrow Sn (s)$$

b. Electroplating of Zinc

The target is cleaned and washed thoroughly with water. This object is dipped in zinc sulphate ($ZnSO_4$) container a small amount of sulphuric acid (H_2SO_4) solutions which acts 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 dissolves and converted into Zn^{-2} ions. The electron moves to the cathode, where Zn^{+2} ions discharged and deposited as Zn metal.

Reactions taking place at electrodes are:

Anode: $Zn (s) \rightarrow Zn^{+2} (aq) + 2e$

Cathode: $Zn^{+2} (aq) + 2e \rightarrow Zn (S)$

c. Electroplating of chromium:

The electroplating of chromium is carried out in electrolytic cell. In this process pure sheet of chromium act as anode a dipped in chromium Acid (H_2CrO_4) solution containing small amount of sulphuric acid (H_2SO4). The object to be electroplated acts as cathode. The electrolyte ionizes and produces Cr+3 ions, at cathode they discharged and deposited on the object.

Reaction taking place at electrodes are:

At Anode: $Cr(s) \rightarrow Cr^{+3} (aq) + 3e$

At Cathode: $Cr^{+3} (Aq) + 3e \rightarrow Cr\ s(s)$

ELECTROCHEMISTRY

(Long Questions Answers)

Q1. a. What is electroplating?

a. Electroplating:

It is the process in which a thin layer of one metal is deposited on another metal by using electricity is called electroplating.

(OR)

The process in which a layer of superior metal is to be deposited on inferior metal by using electricity is called electroplating.

Procedure of electroplating:

i. The metallic substance must be deposited must be cleaned well and pure, washed with sand paper or with water or caustic soda.

ii. The object on which the layer of another metal is to be deposited is made cathode (electrode) by connecting it with negative terminal of battery.

iii. A pure sheet of metal which is to be electroplated is made the anode by connecting it with positive terminal of battery.

iv. A soluble salt of the metal to be electroplated is used as an electrolyte in the cell.

v. Electroplating is carried out in a tank made of cement, wood or glass.

vi. When current is passed, the metal atoms from anode are deposited over the cathode.

Q1. b. Distinguish between the nature of the anode and cathode in such a process.

Ans: **b. Nature of Anode:**

A sheet or rod of pure metal which is to be deposited is made anode. It is connected to the positive terminal of battery. When the electric current is passed, from anode converted into ions in solution. Thus, the mass of anode starts decreasing.

Examples:

$Sn(s) \rightarrow Sn^{+2}(aq) + 2e$

$Zn(s) \rightarrow Zn^{+2}(aq) + 2e$

$Cr(s) \rightarrow Cr^{+3}(aq) + 3e$

Nature of cathode:

The cathode is made of metal which is to be deposit with a superior metal. The metallic ions of superior metal move towards the cathode and are deposited on the object to be coated.

Examples:

$Sn^{+2}(aq) + 2e \rightarrow Sn(s)$

$Zn^{+2}(aq) + 2e \rightarrow Zn(s)$

$Cr^{+3}(aq) + 3e \rightarrow Cr(s)$

Q2. Differentiate between the process of oxidation and reduction. Write an equation to illustrate each.

Ans: There are different concepts to explain the oxidation and reduction.

Oxidation:

1. Classical concept:

According to the classical concept oxidation is a chemical reaction which involves the addition of oxygen.

Examples:

In a given reactions addition of oxygen take place to carbon and hydrogen to form CO_2 and H_2O.

$C + O_2 \rightarrow CO_2$

$2H_2 + O_2 \rightarrow 2H_2O$

(OR)

Removal of hydrogen is also called oxidation according to the classical concept.

Examples:

In a given reactions NH_3 is oxidized to N_2 and H_2S is to S.

$2NH_3 + 3Cl_2 \rightarrow N_2 + 6HCL$

$2H_2S + 2HNO_3 \rightarrow 3S + 2No + 4H_2O$

2. Modern electronic concept:

The modern electronic concept states that a substance which loses electrons in a chemical reaction is said to be oxidized and the reaction is called oxidation.

Example:

In a given reaction ferrous ion is oxidized to ferric ion by losing one electron.

$Fe^{+2} \rightarrow Fe^{+3} + e^-$

Reduction:

Classical concept:

According to the classical concept reduction is a chemical reaction which involves the removal of oxygen.

It is an opposite phenomenon of oxidation

Examples: In a given reactions removal of oxygen take place to form metals.

$Cuo + H_2 \rightarrow Cu + H_2O$

$2HgO \rightarrow 2Hg + O_2$ (OR)

ii. When addition of hydrogen occurs in a chemical reaction is also called reduction according to the classical concept.

Examples:

$N_2 + 3H_2 \rightarrow 2NH_3$

$C + 2H_2 \rightarrow CH_4$

Modern electronic concept:

The modern electronic concept states that a substance which gains electrons in a chemical reaction is said to be reduced and the reaction is called reduction.

Example:

In a given reaction stannic ions is reduced to stannous ion by gaining two electrons.

$Sn^{+4} + 2e \rightarrow Sn^{+2}$

Q3.a. What is corrosion? Explain the rusting of iron as an example of corrosion.

Ans: **a. Corrosion:**

The conversion of any metal into its oxide by the action of environment is called corrosion.

(OR)

It is an oxidation-reduction process which takes place by the action of air in the presence of moisture with the metals.

Explanation:

The meaning of corrosion to a number of people is rust. The term rust and corrosion are almost synonymous. Corrosion is usually starting at the exposed surface of the metal. Corrosion process involve oxidation reduction reaction. Corrosion is a naturally occurring spontaneous phenomenon and it drives the materials to its lowest possible energy state.

Rusting of iron:

Pure iron is silvery metal but when exposed to moist air its surface is corroded and converted to a reddish-brown mass known as rust. Chemically rust is hydrate iron (III) oxide. For rust information there must be a thin film of water on the surface of the metal and air in surrounding. The impurities or the stained portions are responsible for the formation of small electrolytic cells, with anode of pure iron and cathode of impure or strained portions. Iron is oxidized at the anode, producing Fe (II) ions and electrons. It moves along the surface of the metal to cathode where it reacts with water and oxygen to form hydroxide ions.

At anode $2Fe \rightarrow 2Fe^{+2}\,(aq) + 4e$

At Cathode $2H_2O + O2\ 4e \rightarrow 4OH^-$

Fe (II) hydroxide which is further oxidized by atmospheric oxygen to form hydrated Fe (III) oxide, rust

$Fe^{+2}\,(aq) \rightarrow Fe^{+3}\,(aq) + e$

$Fe^{+3}\,(aq) + 3OH\,(aq) \rightarrow Fe(OH)_3(s)$ (rust)

The rust mass is soft and porous in nature and therefore cannot prevent further deeper atmospheric action.

Q3. b. Differentiate between electrolytic cell and Galvanic cell.

Ans: **b.**

Electrolytic cell	Galvanic cell (voltaic cell)
An electrolytic cell converts electrical energy into chemical energy.	A voltaic cell converts chemical energy into electrical energy.
The redox reaction is non spontaneous and electrical energy has to be supplied to initiate the reaction.	The redox reaction is spontaneous and is responsible for the production of electrical energy.
Both the electrodes are placed in a same container in the solution of molten electrolyte. Salt bridge is not required.	The two half cells are set up in different cells containers, being connected through the salt bridge or porous partition. Salt bridge is required.
The anode is positive and cathode is the negative electrode.	The anode is negative and cathode is the positive electrode.
The reaction at the anode is oxidation and the reaction at cathode is reduction.	The reaction at the anode is oxidation and the reaction at cathode is reduction
The external battery supplies the electrons. They enter through the cathode and come out through the anode.	The electrons are supplied by the species getting oxidized. They move from the anode to the cathode in the external circuit.

Q3. **c. Discuss the method of recovering/extracting of metal from its ore.**

Ans: c. The process of extracting metals from their ores is called metallurgy. Electrolytic cell is used as a device for this purpose. Down cell is one the example of recovering of metal from is its ore is which sodium is extracted from NaCl.

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). Commercially sodium metal is obtained from the electrolysis of molten Sodium Chloride (NaCl) in the down cell.

Construction of Down's Cell:

Sodium chloride being a strong electrolyte in the molten state give us Na and Cl ions which are free to move towards their respective electrodes. The cell used for the electrolysis of fused NaCl is called Down's cell. Graphite is used as a "Anode" and a steel electrode at both sides are used as a cathode. Anode is present in the center of the cell. The electrolysis yield sodium (Na) metal and chlorine gas as product. So, it is necessary to keep these products separated, otherwise they will react to give sodium chloride (NaCl) again.

At the cathode, one Na^+ ion pick up one electron and is changed into Na atom. Cl ion will move towards the Anode and give an electron and will change into chlorine atom.

Reaction:

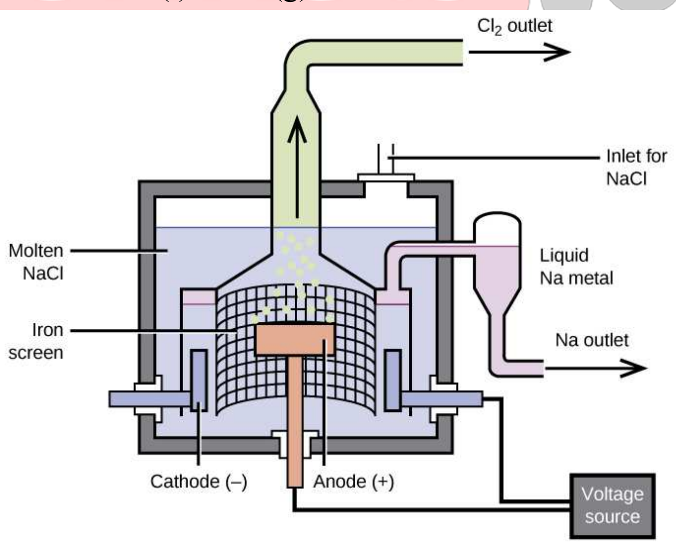
$2Na^+ + 2e^- \rightarrow 2Na$ (reduction)

$2Cl^- \rightarrow 2Cl + 2e^-$ (oxidation)

Net Reaction:

2Nacl (fused) + 2Cl $\rightarrow$ 2Na (l) + Cl_2 (g)

Diagram:



Q4. **Discuss the preparation of Sodium Hydroxide (NaOH) from brine along with diagram and reactions at cathode and anode.**

Ans: **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.

Construction of 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 the 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 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.

2Nacl ➔ 2Na (aq) + 2Cl (aq)

$2H_2O$ (l) + 2e ⇔ H^+ (aq) + OH^- (aq)

Working of Nelson cell:

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 pick up electrons to form H_2 gas. Na^+ ions are not reduced instead with OH ions, present in the solution to form caustic soda (NaOH) which make the solution alkaline, while Cl ions move towards anode where thy give electrons to the electrode.

Reaction at Anode:

2Cl- (aq) ➔ Cl_2 (g) + 2e (oxidation)

Cl (g) + Cl (g) == Cl_2 (g)

At Cathode:

$2H^+$ (aq) + 2e ➔ 2H (aq) (reduction)

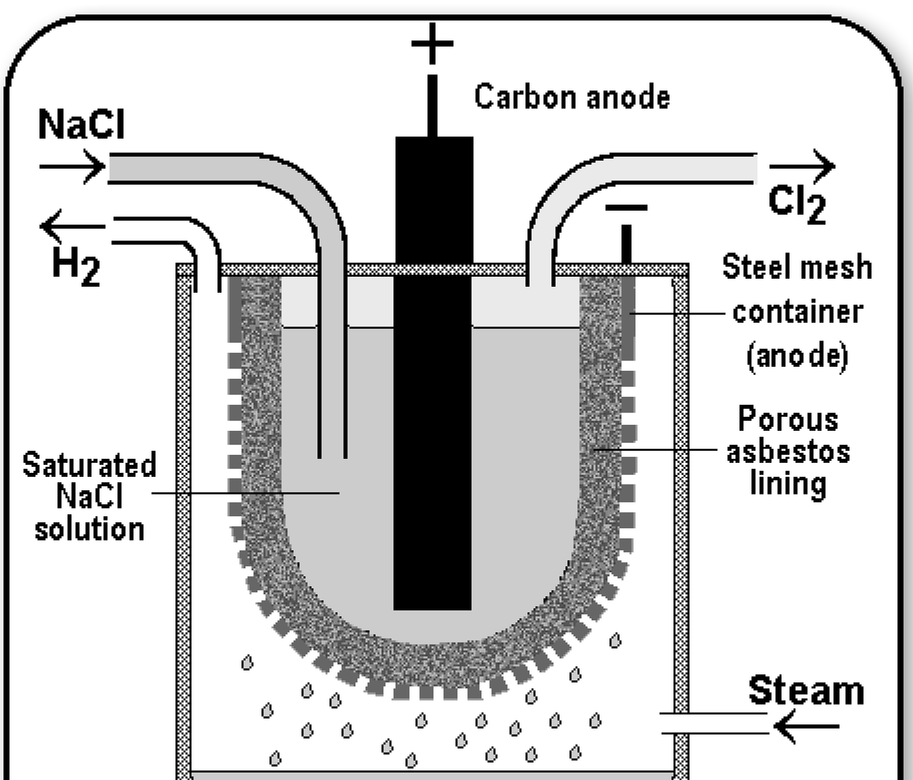
H + H ➔ H_2

$2Na^+$ (aq) + 2OH (aq) ➔ 2NaOH (aq)

The overall reaction:

2NaCl (aq) + $2H_2O$ (l) – 2NaOH (aq) + Cl_2 (g) + H_2 (g)

Diagram:



ELECTROCHEMISTRY

(Short Questions Answers)

Answer briefly the following questions.

Q1. Indicate which element is reduced in the following reactions.

a. $Ca\ (s) + Br_2\ (g) \rightarrow CaBr_2(s)$

b. $8H + MnO_4 + 5e \rightarrow Mn^2 + 4H_2O$

Ans: **a. $Ca\ (s) + Br_2\ (g) \rightarrow CaBr_2(s)$**

Reduction is decreased in oxidation number of an element in a chemical reaction. Hence in above reaction Bromine is reduced because its oxidation number is reduced from 0 to -1.

b. $8H + MnO_4 + 5e \rightarrow Mn^2 + 4H_2O$

In above equation Manganese is reduced because its oxidation number is reduced from +7 to +2.

Q2. What is Oxidation number of silvers on each side of the following equation?

$4Ag\ (s) + O_2\ (g) \rightarrow 2Ag_2O\ (s)$

Ans: **$4Ag\ (s) + O_2\ (g) \rightarrow 2Ag_2O\ (s)$**

Oxidation number of silvers on reactant hand side is zero because it is present in Free state. While oxidation number of silvers on right hand side is +1 due to loss of one electron.

Q3. Why NaOH is a strong but NH_4OH is weak electrolytes?

Ans: NaOH (Sodium Hydroxide) is a strong electrolyte. Strong electrolytes are those electrolytes which

Are dissociate completely in aqueous solution and can conduct electricity to a large extent.

$NaOH (aq) \rightarrow Na^+ (aq) + OH (aq)$ (84%)

While NH4OH (sodium hydroxide) is weak electrolyte which do not dissociate completely in water.

$NH_4OH (aq) \rightleftharpoons NH4^+ (Aq) + OH (aq)$ (1.5%)

Q4. How to prevent corrosion? Enlist few of methods.

Ans: **Prevention of corrosion:**

Corrosion can be prevented by:

i. Metallic Coating:

Metallic coating is used to prevent the metal from the atmosphere effect

Example:

Electroplating (iron particles are protected from rusting by Ni, Cr or tin plating)

ii. Paint Coating:

We can prevent corrosion by paint which as a protective coating.

iii. Alloying:

By alloying the metal.

Example:

Stainless steel is the alloy of iron with Ni, Cr and Si

iv. Corrosion inhibitors:

Corrosion inhibitors are chemicals that react with the metal's surface or with the environmental gases which cause corrosion.

Example: Glycine, polyethylene etc.

v. Cathodic Protection:

Cathodic protection is a method usually used to protect iron in buried fuel tanks and pipelines. For this purpose mostly active metals like zinc or magnesium is used.

Q5. Write chemical reactions that occur in Nelson's cell.

Ans: **Reactions that occur in Nelson's are:**

The saturated brine ionized as follows:

$2NaCl \rightarrow 2Na (aq) + 2Cl (aq)$

$2H_2O (l) \leftrightarrow H + (aq) + OH^- (aq)$

At Anode:

$2Cl- (Aq) \rightarrow Cl_2 (g) + 2e$ (oxidation)

$Cl (g) + Cl (g) \rightarrow Cl_2 (g)$

At Cathode:

$2H^+ (aq) + 2e \rightarrow 2H (aq)$

$H + H \rightarrow H_2$

$2Na^+ (aq) + 2_0H (aq) \rightarrow 2NaOH (aq)$

The overall reaction:

$2NaCl (aq) + 2H_2O (l) \rightarrow 2NaOH (aq) + Cl_2 (g) + H_2 (g)$

Q6. Write an example from daily life which involves the oxidation-reduction reaction.

Ans: Photosynthesis in plants is the example of oxidation-reduction reaction. It converts the light energy into chemical energy. This process consists of series of chemical reactions that require carbon dioxide, water and store chemical energy in the form of sugar, light energy from light drives the reactions.

$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

Photosynthesis transfers electrons from water to carbon dioxide molecules. This transfer of electrons is the example of an oxidation-reduction process. Water is oxidized by losses electrons and carbon dioxide is reduced by gaining electrons.

Q7. Assign oxidation numbers to each atom in the following compounds.

a. Hl b. PBr_3 c. $CaCo_3$ d. H_2PO_4 e. As_3O_5 f. H_2SO_4

Ans: a. Hl – H = +1, 1 = -1

b. PBr_3. P = +3, Br = -1 (3Br = -3)

c. $CaCo_3$ – Ca = +2, C = +4, O = -2 (30 = -6)

d. H_3PO_4. H = +1 (3H = +3), P = +5, 0 = -2 (40 = -8)

e. As_3O_5. As = +3.33 (3As = +1U), (50 = -10)

f. H_2SO_4. H = +1 (2H = +2), S = +6, 0 = -2 (40 = -8)

Q8. Why oxygen is necessary for rusting?

Ans: Oxygen is necessary for rusting. It converts metal into hydrated rust. Dents and strains present on the surface of iron acts as an anode which oxidizes the iron.

$2Fe\ (s) \rightarrow 2Fe^{+2}\ (aq) + 4e$

While that part on the surface of iron where oxygen and water are present acts as cathode. The free electrons move to this part and following reaction occurs.

$2H_2O + O_2 + 4e^- \rightarrow 4OH$

$Fe^{+2}$is further oxidized by atmospheric oxygen to form hydrated Fe (III) oxide which is known as rust.

$Fe^{+2}\ (aq) \rightarrow Fe+3 + e^-$

$Fe^{+3} + 3OH^-\ (aq) \rightarrow Fe\ (OH)_3\ (s)$.

Q9. Sketch the Daniel cell, labeling the cathode, anode and the direction of flow of electrons.

Ans:

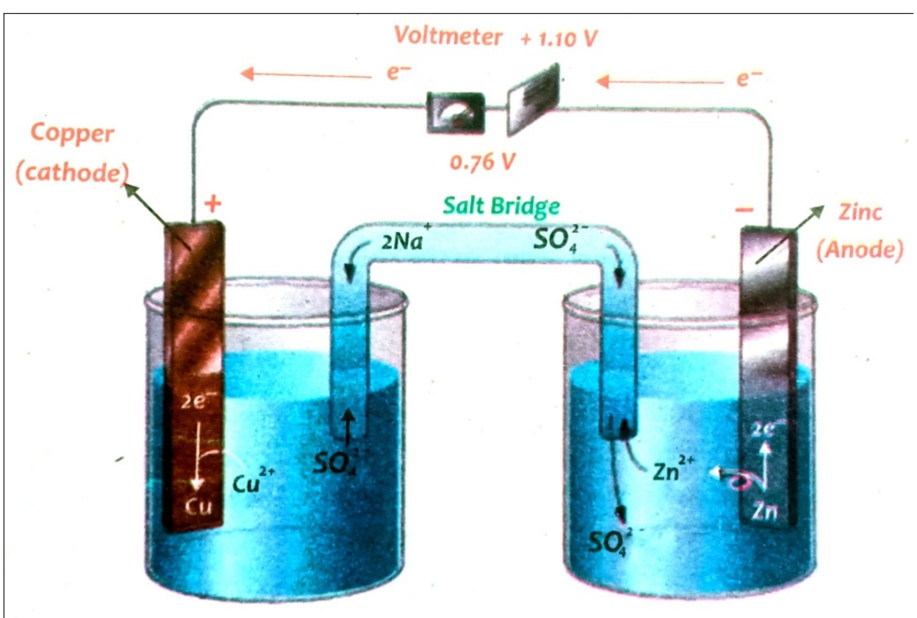


Fig. 7.6: Daniel cell

Q10. Write down some possible uses of an electrolytic cell.

Ans: **<u>Uses of an electrolytic cell:</u>**

i. It is used for the commercial preparation of sodium metal.
ii. It is used for the conversion of electrical energy into chemical energy.
iii. It is used for the purification of impure metals.
iv. It is used in the extraction of metals from their ores.
v. It is also used in the process of electroplating.

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CHEMISTRY

Class 9th (KPK)

NAME: ______________________________

F.NAME: ______________________________

CLASS:_____________ SECTION: _________

ROLL #: ______ SUBJECT: ______________

ADDRESS: ____

____

SCHOOL: ____

CHEMICAL REACTIVITY
(Topic Wise Questions Answers)

Q1. Write the characteristic of Metal and Non-Metals.

Ans: **Electropositive Character:**

All elements have the ability to lose electrons easily from their valence shells and get (+ive) charged to form cation. Electron losing ability is called electro-positivity.

Example:

$Na + energy \rightarrow Na^+ + 1e^-$ $\Delta E = 496$ kj/mol

On other hand, non-metals have the ability to accept electrons in their valence shell to get (-ive) charged particle called Anion.

Example: $Cl + e \rightarrow Cl^-$

2. Electrical Conductance:

Metals are good conductor of heat and electricity. While non-metals are insulator.

The conductance in metal is due to mobile sea of electrons which are loosely held are responsible for the conduction of electric current.

3. Nature of Oxides:

Metal are basic in nature while non-metals oxide are acidic nature.

Example:

$Na + O_2 \rightarrow Na_2O$

$Na_2O + H_2O \rightarrow 2NaOH$

Similarly

$2S + 3O_2 \rightarrow 2SO3$

$SO3 + H_2O \rightarrow H_2SO_4$

Q2. What are Alkali Metals? Also explain occurrence of alkali metals.

Ans: **Alkali Metals:**

The elements of the group IA except Hydrogen are called alkali metals.

The name Alkali came from Arabic language. It means Ashes. These metal were first found in the ashes of plants.

Some chemist had the opinion that the word alkali is given due to the fact that these elements react with water and forming the strong Alkalies. Alkali metal 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 IA form ionic compounds with elements of group VIA and group VIIA.

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 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 laboratory.

Q3. **What are Alkaline earth Metals? Also explain occurrence of alkaline earth metals.**

Ans: **Alkaline Earth Metals:**

The elements of 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 valence 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 (Br), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). They become stable by gaining the electronic configuration of noble gases by losing their outermost electrons. These metals are often found in the form of sulphate in nature, Examples include the minerals such as gypsum (calcium sulphate), epsomite (magnesium sulphate) and barite (barium sulphate).

Occurrence of Alkaline earth metals:

Aklaline 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 the sea waters. 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 radioactive in nature.

Q4. **How ionization potential values vary for Group I and group II elements on descending the group?**

Ans: **i. Energies of Group I and II elements:**

Ionization Energy:

The amount of energy required to remove an electron from an isolated gaseous atom of an element is called ionization energy.

Example: (Alkali Metals)

The alkali metals have one electron in their outer most shell. E.g.

$Na \rightarrow Na^{+} + 1e$ $\Delta E = 496$ kj/mole
(2.8.1) (2.8)

In alkali metals, sodium (Na) has the highest I. Energy in its own group due to smaller size and the small distance b/w the nuclear charge and valence electron. Down the group I.E is decrease due to increasing number of shells by increasing in atomic number.

Now the distance b/w the nuclear charge and the valence shell electrons are also increase. So, it is easier to remove an electron due to less bonded.

I.E of Group IA

Elements	Atomic No	Atomic Radius	I.E Jk/mol
Li	3	$1.52A^0$	520
Na	11	$1.86A^0$	496
K	19	$2.27A^0$	419
Rb	37	$2.48A^0$	403
Cs	55	$2.68A^0$	375

Alkaline earth metals:

Alkaline earth metals have two electrons in their valence shell. Since atomic radii decrease due to increase of nuclear charge therefore high amount of energy will be required to remove an electron from the valence shell.

$Mg \longrightarrow Mg^{+1} + 1e \quad \Delta E = 738$ kj/mol
Removal of second e after first one then
$Mg^{+1} \longrightarrow Mg^{+2} + 1e \quad \Delta E = 1450$ kj/mol
Ionization energy of element increase from left to right in a period due to smaller size and increasing nuclear charges.

Similarly:

It decreases down the group due to increasing of shell and the distance b/w the nucleus and valence shell electron also increase, so down the group ionization energy decreases. There is general decreasing order to melting and boiling points, hardness, conductivity and ionization energy down the group.

Group I	Group II	Properties			
Li	Be	Decreasing ↓	Melting point Boiling Point Hardness Conductivity Ionization Potential	Increasing	Electro positivity Atomic radii Atomic volume Reactivity Reducing power density
Na	Mg				
K	Cs				
Rb	Sr				
Cs	Ba				

Periods

		Decreasing →	Increasing →
Group I	Group II	Electro positivity	Melting point
Li	Be	Atomic radii	Boiling point
Na	Mg	Conductivity	Density
K	Ca	Atomic volume	Hardness
Rb	Sr	Reactivity	
Cs	Ba	Reducing agent	

Q5. What is the difference in the reactivities of Group I and Group II elements? Describe with Respect to the variation in atomic number and ionization potential.

Ans: **Differences in the reactivity of group IA and IIA.**

The reactivity of element shows that how much the element is reactive when it is reacted with other substances especially air, acids and water. The differences in the reactivities of group IA and group IIA with respective to atomic number and I.P is as follow.

Differences in the reactivities with respect variation in atomic number and I.P

As we go from group IA to group IIA, along a period, the atomic number increases, but the number of shells remain the same. Thus, the nuclear charge increases and the atomic size decreases.

Therefore, the valence electrons of group IIA are more tightly bound to the nucleus as compared to alkali metals and hence group IA elements are more reactive as compared to group IIA. Down the group both in group IA and IIA, the atomic size increases due to the addition of new shells although the atomic number increases which increase the nuclear charge. Thus, the valence electrons become farther from the nucleus and they can be easily removed. Therefore, the IE values decrease down the group due to which the reactivity is increased down the group.

Q6. Describe the position, properties and uses of Sodium.

Ans: **Sodium (Na):**

Sodium does not occur as a free metal in nature because it is too reactive metal and readily combines with other elements and compounds. It is found in sea as sodium chloride, sodium bromide and sodium iodide. It is also found in deposits as rock salt.

Its Latin name is “Natrium”

Position of Sodium in Periodic table:

Sodium belongs to alkali metals. Sodium atomic number is "11" and mass number is "23" and its symbol is "Na". It occupies first position in 3rd period and 3rd position in Alkali metal (Group IA) it has three electronic shell have only one electron in their valence shell.

Physical Properties:

i. It is silvery white solid.
ii. Na is soft metal and can be cut with a knife
iii. Its density $(\frac{m}{y})$ is 0.971g/cm^3
iv. Its melting point in 97.6^{0}C and boiling point is 880^{0}C.
v. It has relatively low tensile strength
vi. It is lighter than water and therefore floats on the surface of water.
vii. It ductile (which can be drawn to form wires) and malleable.
viii. It is good conductor of electricity due to the free movement of valence electrons.

Chemical Properties:

Sodium (Na) is highly reactive and can react with water (H_2O), hydrogen ((H_2), Oxygen (O_2) and halogens (Group VIIA)

1. Reaction with water:

Sodium react vigorously with cold water forming metal hydroxide with the liberation of hydrogen gas.

$2Na(s) + 2H_2O \rightarrow 2NaOH\ (aq) + H_2\ (g)$ + heat

The reaction is exothermic (heat released) as a result hydrogen produced catch fire on the surface of water.

2. Reaction with hydrogen:

Sodium react with hydrogen to form hydrides.

$2Na + H_2 \rightarrow 2NaH($ $\rightarrow$ metal hydrides)

3. Reaction with oxygen:

Sodium (which is metal) react with oxygen from basic oxide and react with water form Alkali (Base).

Examples:

$4Na\ (s) + O2\ (g) \rightarrow 2Na_2O$ (sodium oxide)

$Na_2O + H_2O \rightarrow 2NaOH$ (sodium hydroxide)

4. Reaction with halogens:

Sodium react with halogens (Group VIIA) to form sodium halide

$2Na(s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$

$2Na\ (s) + Br_2(l) \rightarrow 2Na\ Br\ (s)$

5. Reaction with Sulphur:

Sodium is powerful reducing agent. It reduces the other substance but itself oxide.

$2Na\ (s) + S\ (s) \rightarrow Na2S(s)$

6. As a reducing agent:

Sodium is powerful reducing agent. It reduces the other but itself oxide.

$2Na^0(s) + Mg^{+2}O^{-2}\ (s) \rightarrow Na_2^{+1}\ O\text{-}2\ (s) + Mg^0\ (s)$

$4Na^0\ (s) + Ti^{+4}\ Cl_4\text{-}4\ (s) \rightarrow 4Na^{+1}\ Cl^{-1}$ (s) of Ti^0

In above case, the oxidation state of Na^0 is zero and in Na_2O, the oxidation state change to (+2) increasing oxidation state occur is reducing agent.

Uses of Sodium:

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 (NaNH2).

ii. It is used in sodium vapour lamps (which gave a bright orange-yellow light) for street lighting.
iii. It is used as coolant in nuclear reactors.
iv. It is used in 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 (Ti), Zirconium (Zr) from chlorides or oxides.
vi. It forms alloys with other metal. Its most useful alloy is with mercury (Hg) called sodium amalgam and with metal silver.

Q7. Write the position, properties and uses of Magnesium and Calcium.

Ans: **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 essential constituent of chlorophyll in green plants.

Position of Magnesium in Periodic Table:

Magnesium atomic number is 12 and its symbol is "Mg". It occupies second position in 3rd period and second in group IIA as it has three electronic shells and two electrons in their valence shell.

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.

Position of Calcium in Periodic table:

Calcium atomic number is 20 and its symbols is "Ca". It occupies second position in 4th period and third position in group IIA as it has four electronic shells and two electrons in their valence shell.

Physical Properties of Magnesium

i. Magnesium is silvery grey solid.
ii. Its density is 1.74g/cm^3.
iii. Its melting point 651^{0}C and boiling point is 1106^{0}C.
iv. It is malleable and ductie.
v. It is good conductor of heat and electricity.

Physical properties of Calcium:

i. Calcium is silvery white solid.
ii. Its density is 1.55g/cm^3.
iii. Its melting point is 851C^0 and boiling point is 1106C^0.
iv. It is malleable and ductile.
v. It is good conductor of heat and electricity.

Chemical properties of Magnesium and calcium:

1. Reaction with H2:

Both "Mg" and "Ca" combined directly with hydrogen formed hydrides.

Mg (s) + H_2 $\rightarrow$ MgH_2 (Magnesium hydride)

Ca (s) + H_2 (g) $\rightarrow$ CaH_2 (Calcium hydride)

2. Reaction with Oxygen:

Both (Mg, Ca) burn in air. Magnesium burns with a dazzling flame forming MgO called magnesia.

$2Mg (S) + O_2(g) \rightarrow 2MgO (s)$

While calcium form CaO produce brick red coloured flame.

$2Ca (s) + O_2(g)$ heating $2CaO (s)$

Both (Mg, Ca) form base, when dissolved in water.

$2MgO (s) + H_2O (I) \rightarrow 2Mg (OH)_2$

$2CaO (s) + H_2O (I) \rightarrow 2Ca (OH)_2 (aq)$

3. Reaction with Nitrogen:

Both (Mg, Ca) react with "N_2" form nitrides

$3Mg (s) + N_2 (g) \rightarrow Mg_3N_2 (s)$

$3Ca (s) + N_2 (g) \rightarrow$ heat $Ca_3N_2 (s)$

4. Reaction with Acids:

Both (Mg, Ca) react with strong & dill acids give us hydrogen (H2) gas.

$Mg (s) + 2HCl (aq) \rightarrow MgCl_2 (s) + H_2 (g)$

$Mg (s) + H_2SO_4 (dil) \rightarrow MgSO_4 (s) + H_2 (g)$

$Ca (s) + 2HCl (dil) \rightarrow Cacl_2 (s) + H_2 (g)$

$Ca (s) + H_2SO_4 (aq) \rightarrow CaSO_4 (s) + H_2 (g)$

Reaction with Halogens:

They react with halogens form halides

$Mg (s) + Cl_2 (g) \rightarrow MgCl_2 (s)$

$Mg (s) + Br_2 (l) \rightarrow MgBr_2 (s)$

$Ca (s) + Cl_2(g) \rightarrow CaCl_2 (s)$

$Ca (s) + Br_2(l) \rightarrow CaBr_2 (s)$

Uses of Magnesium:

vi. Mg is low density metal, so it is used in the formation of light but tough alloys, such as Duralumin (a mixture of Al, Cu, Mg and Mn) Magnesium (a mixture of Al, Mg). These alloys are used for construction of aircrafts, cars and moving parts of machines.

vii. It is also used in phogrophic flashlight powder, flames and fireworks.

viii. It is used deoxidant in metallurgy and the extraction of Titanium and Uranium.

ix. Its compounds such as magnesium oxide (MgO) are mixed with clay, to make refractory bricks for furnace lining.

x. Magnesium sulphate (MgSO4) is used in textile, paper industry, soap formation and pharmaceutical industries etc.

Uses of Calcium:

i. Calcium is used as dioxide in steel coatings and copper alloys.

ii. It is used in the making of calcium and calcium hydride and in extraction of uranium.

iii. Their compounds such as lime CaO is added to soil in the form of fertilizers to decrease its acidity. It is also used for softening, pollution control and in pulp, paper, sugar and glass manufacturing industries.

iv. It is used in steel making.

Q8. What are Soft and Hard metals?

Ans: **Soft and Hard metals:**

Soft Metals:

Metals of group IA elements are quite soft, they react quickly with H2 and O2 and violently with H2o, and such metals are called soft metals. They are soft and have low melting and boiling point.

Example: Na, Li, k etc.

Hard metals:

The metal of "d" and "f" block elements are hard metals. They are hard in their physical appearance. Iron (fe), Copper (Cu) Silver (Ag), Cobalt (Co), Nickel (Ni), Tungsten (W) are hard, their melting point, boiling points and density show much higher values.

They do not react readily under ordinary conditions of temperature and pressure.

Both soft metals and hard have their own importance. Such as iron is used to prepare steel which is harder form of iron also used in heavy machinery locomotives, railway tracks in the construction of bridges.

Q9. Write down the comparison properties of Sodium (Na) and Iron (Fe).

Ans: Comparison properties of 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 outermost shell and is very soft and can be cut with knife.	It is hard and requires great energy to break.
It has weak attractive force between the atoms of sodium.	It has strong attractive force between the atoms of iron.
The melting point of sodium is 97.6^0C	The melting point of iron is 1538^0C.
The boiling point of sodium is 880^0C.	The boiling point of iron is 2862^0C.
It has low density $0.927g/cm^3$	It has higher density $7.874g/cm^3$
It is lighter and floats on the surface of water.	It is heavy and settles at the bottom of water.
It has low tensile strength, cannot be used where stress is required.	It has high tensile strength which can be used in construction of builing and bridges. It is also used to prepare steel.
It is very reactive, stored in kerosene oil.	It is less reactive than sodium.

Q10. Write a note on commercial value of silver (Ag), Platinum (Pt) and Gold (Au)?

Ans: Those metals which are expensive and have great commercial and economic value are precious metals. In noble metals particularly silver (Ag), Platinum (Pt) and Gold (Au) are considered as precious metals.

i. Silver (Ag):

Silver is soft, white metal that usually occurs in nature in one of four forms, a) A native element, b) as a primary constituent in silver mineral, c) as a natural alloy with other metals, d) as a minor constituent in the ores of other metals.

Silver is known as a precious metal because it is rare and 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 combine 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 patina meaning little silver. Platinum is the 72^{nd} most common element in earth crust. That is why, platinum is an expensive metal.
Platinum is heavy, soft, malleable, and ductile and has a fairly high melting point (1770^0C). It is noble metal because it is un-reactive. It does not even react with oxygen in air and resistant to react with acids.
Platinum is used in the catalytic converters to remove pollutants from the car engine exhaust gases. But as an expensive metal, so metals such as palladium etc. are used in its place.
The ease with which platinum can be shaped, its strength, colour, hardness and inertness make it suitable for jewellery and gem setting. Un-reactivity also makes it useful in dental fillings, making surgical tools 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:

Gold has been used to make ornamental objects and jewellery for thousands of years. Special properties of gold like 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 stress applied to many jewellery items. Alloying gold with other metals such as copper, silver and platinum increases its durability. Older than the coins were made of gold. Gold coins were commonly used in transactions up to paper currency because a more common form of exchange.
Gold is using a 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 reliable performance justifies the high cost. Gold alloys are used for dental filling, tooth crowns and orthodontic appliance. Gold is used in dentistry because it is chemically inert, non-allergic and easy for the dentist to work.

Q11. Write the electronegative characters of non-metals.

Ans: **Electronegative character of non-metals:**

The tendency of an atom of an element to gain electron from other element in order to become stable electronic configuration is called electronegative character.
Every element tries to complete its valence shell to become stable. The electronegative character increase across the period so all elements of group VIIA are most electronegative in their respective period. They accepted electron from less electronegative elements to complete its valence shell by octet rule. The general electronic configuration of group VIIA is ns^2np^5.

$X + e^- \rightarrow x$

Where x is called halide ion

$F + e^- \rightarrow F^-$ (Fluorine ion)

Group VIIA are collectively called halogen which mean salt formers. They contain fluorine (f), chlorine (cl), bromine (Br), Iodine (I) and Astatine (At)

Electro configuration and physical state of halogens

Name	AT No.	Symbol	Physical State	Electron Affinity	Electro negativity	Atomic Radius
Fluorine	9	F	Pale yellow gas	-322	4.0	72×10^{-9} m
Chlorine	17	Cl	Green Gas	-349	3.0	99×10^{-9} m
Bromine	35	Br	Dark red liquid	-325	2.80	114pm
Iodine	53	I	Dark crumble solid	-295	2.50	133 pm
Astatine	85	At	Black solid	-270	2.20	150pm

Q12. Write the physical properties of Halogens.

Ans: **Physical Properties:**

1. Fluorine and chlorine are gases.
2. Bromine is a fuming gas.
3. Fluorine is yellowish gas.
4. Cl_2 is greenish yellow colour gas.
5. Bromine is radish brown liquid.
6. Iodine is deep violet solid.
7. Astatine is Black solid.

They all are electronegative in all over the periodic table. They form (-ive) ion when react with metals. Fluorine is the most electronegative atom and form hydrogen bonding. Their melting and boiling point increase down the group.

CHEMICAL REACTIVITY
(Long Questions Answers)

Q1. **Compare and contrast the properties of alkali and alkaline earth metals, with reactions.**

Ans: **Comparison between physical properties of alkali and alkaline earth metals:**

ALKALI METALS	ALKALINE EARTH METALS
They are all silvery white metals.	They are all silvery white metals but be is grayish white.
They are soft metals.	They are soft metals but harder than alkali metals.
They have large atomic sizes	They have small atomic sizes
They have large atomic radii and ionic radii	They have small atomic radii and ionic radii
They have lower melting points and boiling points.	They have higher melting and boiling points.
They have lower densities	They have higher densities
They have low ionization energies and electronegativity values	They have higher ionization energies and electronegativity values.
They have lower electron affinity but higher than alkaline earth metals.	They have exceptionally lower electron affinity than alkali metals.
They are less conductor of heat and electricity	They are more conductor of heat and electricity

Comparison between chemical properties of alkali and alkaline earth metals:

Alkali metals are more reactive than the alkaline than the alkaline earth metals because they have one electron in valence shell while alkaline earth metals have two electrons. Therefore, alkali metals can easily loss their electron and are more reactive.

i. Reactive with hydrogen:

Elements of group 1A and group IIA react with hydrogen and forms their respective hydrides.

$2Na + H_2 \rightarrow 2NaH$

$Mg + H_2 \rightarrow MgH_2$

ii. Reactive with Oxygen:

Elements of group IA and group IIA with oxygen and forms their respective oxides.

$2Na + O_2 \rightarrow 2NaO$

$Mg + O_2 \rightarrow MgO_2$

iii. Reaction with Halogens:

Both reacts with halogens forming halides.

$2Na + Cl_2 \rightarrow 2NaCl$

$Mg + Cl_2 \rightarrow MgCl_2$

iv. Reaction with water:

Most of alkali metals reacts with water liberating hydrogen gas while alkaline earth metals react slowly except beryllium which do not react with water.

$2Na + H_2O \rightarrow 2NaOH + H_2$

$Mg + H_2O \rightarrow MgO + H_2$

v. Reaction with Nitrogen:

Among the alkali metals only lithium reacts with nitrogen, while all alkaline earth metals react with nitrogen forming nitrides.

$6Li + N_2 \rightarrow 2Li_2N$

$3Mg + N_2 \rightarrow Mg_3N_2$

Q2. a **Differentiate between soft and hard metals.**

Ans: **a. Differences between soft and hard metals.**

SOFT METALS	HARD METALS
The metals present in group IA and IIA are very soft, very reactive, 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.
They are very soft in nature even some of them can be cut with knife.	They are very hard in nature and requires greater energy to break.
They are very reactive in nature.	They are less reactive in nature
They have low ionization energies	They have high ionization energies
They are very electropositive in nature	They are less electropositive in nature.
They have low densities and are lighter which can float on the surface of water.	They have higher densities and are heavy which settles as the bottom of water.
They have low melting points and boiling points.	They have higher melting and boiling points.
They readily react with H_2 and O_2	They do not readily react with H_2 and O_2 at normal condition.
They violently react with H_2O	They may or may not react with H_2O, some react with H_2O but very slowly
They have weak attractive force between the atoms.	They have strong attractive force between the atoms.

Q2. b **Give the reaction of magnesium with: i. H_2 ii. HCl iii. O_2 iv. H_2O v. Cl_2**

Ans: **b. Chemical properties of Magnesium:**

1. Reaction with H2:

Magnesium combined directly with hydrogen formed hydrides.

Mg (s) + H_2 (g) ➔ MgH_2 (s) (Magnesium hydride)

2. Reaction with Oxygen:

Magnesium burns in air with a dazzling flame forming MgO called magnesia.

2Mg (s) + O_2 ➔ 2MgO (s)

3. Reaction with H2O:

When magnesium oxide (MgO) is dissolved in water, it forms basic solution.

2MgO (s) + H_2O (I) ➔ 2Mg $(OH)_2$ (aq)

4. Reaction with HCl:

Magnesium (Mg) reacts with HCl give us hydrogen (H_2) gas.

Mg (s) + 2HCl (aq) ➔ $MgCl_2$ (s) + H_2 (g)

5. Reaction with Cl2:

Magnesium reacts with chlorine and form halides

Mg (s) + Cl_2 (g) ➔ $MgCl_2$ (s)

Q3. **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.**

Ans: **Elements exist in Free state:**

Some elements exist in free state because they have completed their outermost shells and are stable. These elements have high ionization energies and having low reactivity. So that's why they cannot easily take part in a chemical reaction and do not form chemical bond with other elements, and exist freely.

Examples:
Group VIIIA (noble gases)
Noble material like Ag, Cu, Hg and Au.

Elements in combined state:
Some elements exist in combined state because they have incomplete their outermost shells and are unstable. These elements have low ionization energies and large atomic size. They having high reactivity. So, to complete their outermost shell they easily take part in a chemical reaction and forms chemical bond with other elements, and cannot exist freely i.e. occur in combined state.

Examples:
Alkali metals and alkaline earth metals.
Halogens, Carbon family and Oxygen family etc.

Q4. Define metal and non-metal and compare the properties (both physical and chemical) of metals and non-metals.

Ans: **Metal:**
A metal is an element which loses an electron and forms a cation.

Explanation:
Metals are those substances which are good conductor of heat and electricity. Their oxides and hydroxides are basic in nature. When a metal reacts with in oxygen it produces a basic oxide. When it is dissolved in water it forms an alkaline solution which turns red litmus paper into blue.

Examples:
Elements of group IA except hydrogen, Group IIA, transition elements, lanthanides and actinides.

Non-metal:
A non-metal is an element which gains an electron and forms an anion.

Explanation:
Non-metals are those substances which are non-conductor of heat and electricity. Their oxides and hydroxides are acidic in nature. When a non-metal reacts with in oxygen it produces an acidic oxide. When it is dissolved in water it forms an acidic solution which turns blue litmus paper into red.

Examples:
Hydrogen, boron of group IIIA, C and Si of group IVA, N and P of group VA, Group VIA, Group VIIA and Group VIIIA.

Comparison between the properties (both physical and chemical) of metals and non-metals.

METALS	NON-METALS
They are good conductor of heat and electricity.	They are non-conductor of heat and electricity.
Their oxides and hydroxides are basic in nature.	Their oxides and hydroxides are acidic in nature.
They are ductile, malleable and sonorous.	They are not ductile, malleable and sonorous.
They are usually solids at room temperature except mercury.	They are present in all three states of matter i.e. solid, liquids and gases.
They electron donor in chemical reactions. They are reducing agents.	They are electron acceptor in chemical reactions. They are oxidizing agents.
They become positively charged ion in solution.	They become negatively charged ion in solution.
They are electropositive in nature.	They are electronegative in nature.
They form electrovalent (ionic) chlorides.	They form covalent chlorides.

Some metals can replace hydrogen from acids to form salts	Non-metals cannot replace hydrogen from acids.
They do not combine easily with Hydrogen. Few hydrides are formed are electrovalent. $2Na + H_2 \rightarrow 2NaH$ $2K + H_2 \rightarrow 2KH$	They combine easily with hydrogen to form many stable hydrides. $H_2 + Cl_2 \rightarrow 2HCl$ $H_2 + F_2 \rightarrow 2HF$ $H_2 + L_2 \rightarrow 2Hl$

Q5. Halogens are very reactive elements, write down halogen's reactions with hydrogen, oxygen,

metals, non-metals and other compounds along with displacement reaction.

Ans: **Chemical Properties of Halogens:**

All halogens are very reactive elements and exist in diatomic state with single covalent bond.

i. Reaction with H2:

$H_2 (g) + F_2 (g)$ quick $2HF (aq)$

Fluorine react with hydrogen vigorously

$H_2 (g) + Cl_2 (g)$ quick $2HCl (aq)$

$H_2 (g) + Br_2 (I)$ light $2HBr (aq)$

$H_2 (g) + l_2 (s) <===> 2Hl (aq)$

ii. Reaction O2:

Fluorine react with oxygen to form monoxide and dioxide (di-oxygen – di fluoride)

$O_2 (g) + 2F_2 (g) \rightarrow 2OF_2 = (g)$ (monoxide)

$O_2 (g) + F_2 (g) \rightarrow O2F_2 (g)$ (dioxide)

iii. Reaction with metals:

Halogens react with metals and form corresponding halides.

$Cu (s) + Br_2 (I) \rightarrow CuBr_2$ (Copper bromide)

$2K (s) + l_2 (s) \rightarrow 2Kl (s)$ (Potassium iodide)

iv. Reaction with non-metals:

Halogens react with non-metals such as phosphorous to form PCl3 (tri-chloride) and PCl5 (Penta chloride)

$2P(s) + 3Cl_2 (g) \rightarrow 2PCl_3 (aq)$

$2P(s) + 3Br_2(g) \rightarrow 2PBr_3(aq)$

$2P(s) + 5Cl_2 (g) \rightarrow 2PCl_5 (aq)$

$2P(s) + 5Br_2 (g) \rightarrow 2PBrs (aq)$

v. Reaction with other compounds:

Halogens oxidized other compounds but itself reduce, during reaction

$H_2S (aq) + Cl_2 (g) \longrightarrow 2HCl (Aq) + S (s)$

$2NH_3 (aq) + 3Cl_2 (g) \longrightarrow 6HCl (aq) + N_2 (g)$

vi. Displacement Reactions:

During displacement reaction a more reactive halogen will displace a less reactive halogen from its halide solution. They reactivity of halogen decreases down the group.

Order of reactivity F > Cl > Br > I > As.

Examples:

i. $2NaBr + Cl_2 \rightarrow 2Nacl + Br_2$

ii. $2Kl + Br_2 \rightarrow 2Kbr + l_2$

CHEMICAL REACTIVITY

(Short Questions Answers)

Q1. Identify at least wo groups which contain only metallic elements.

Ans: In periodic table most of the metal elements are present at the left side of the periodic table. All The elements of group IA (Except hydrogen) and group IIA are metals. Group IA contain Li, Na, K, Rb, Cs and Fr while group IIA contains Be, Mg, Ca, Sr, Ba and Ra.

Q2. Write the reaction of group IA metals with oxygen, with balance equations.

Ans: **Alkali metals react with oxygen and forms various types of oxides:**

i. In presence of oxygen lithium burns with red flame and give lithium oxide, which is white solid.

1. $4Li\ (s) + O_2\ (g) \rightarrow 2Li_2O\ (s)$

ii. In presence of oxygen sodium burns with bright yellow flame and give white sodium oxide.

2. $4Na\ (s) + O_2\ (g) \rightarrow Na_2O_2\ (s)$

iii. In presence of oxygen potassium burns violently with a little coloured flame and give white potassium oxide.

3. $K\ (s) + O_2\ (g) \rightarrow KO_2\ (s)$

iv. Similarly rubidium and cesium catch fire in air and produce superoxide.

4. $4Rb\ (s) + O_2\ (g) \rightarrow 2RbO_2\ (s)$
5. $4Cs\ (s) + O_2\ (g) \rightarrow 2CsO_2\ (s)$

Q3. State the physical properties of metals.

Ans: **Physical properties of metals:**

i. All metals are solid at room temperature and on atmospheric pressure except mercury.
ii. Metals are malleable i.e. they can be beaten into sheets and foils.
iii. Metals are ductile i.e. they can be drawn into wires.
iv. All the metals are good conductors of heat and electricity.
v. Metals are lustrous i.e. they have shiny surfaces.
vi. Metals are sonorous i.e. they produce ringing sound when struck.
vii. They have high melting points and boiling points.
viii. They have low I.E, E.A and E.N.
ix. They have large atomic masses as compare to nonmetals.

Q4. How does sodium act as reducing agent and write down its reaction also?

Ans: Sodium is powerful reducing agent. It reduces the metal oxides into metals and itself oxidize.

$2Na^0\ (s) + Mg^{+2}\,O\text{-}2\ (s) \rightarrow Na_2^{+1}\,O^{-2}\ (s) + Mg^0\ (s)$

$4Na^0\ (s) + Ti^{+4}\,Cl_4^{-4}\ (s) \rightarrow 4Na^{+1}\,Cl^{-1}\ (s) + Ti^0\ (s)$

In above case, the oxidation state of Na0 is zero and in Na_2O, the oxidation state change to (+2) increasing oxidation state occur is reducing agent.

Q5. Ionization energy of Alkaline earth metals is higher than alkali metals, why?

Ans: The amount of energy required to remove an electron from isolated gaseous atom of an element is called ionization energy.

I.E of alkaline earth metals (group IIA) is higher than alkali metals (group IA). Because the atomic size of alkaline earth metal is smaller than the atomic size of alkali metals.

Akali metals have one electron in their outer most shell while alkaline earth metals have two electrons. Therefore, higher energy is needed to remove two electrons from alkaline earth metals as compare to alkali metals.

Example: Group 1A

After removing one electron 1e
Na ➔ Na^+ + 1e
I.E = 496 kj/mol

after removing one electron
Mg ➔ Mg^+ 1e
I.E = 738 kj/mol

Q6. Pure gold is not used for ornaments, why?

Ans: Gold has been used to make ornamental objects for thousands of years. Special properties of gold Like very high luster, attractive colour, inertness, resistivity etc. makes it perfect for manufacturing jewelry. Pure gold is not used for ornaments because pure gold is too soft and malleable to resist the stresses applied due to which it can easily be, deshaped, by applying a little force. Therefore, alloying gold with other metals such as copper, silver and platinum increase its strength, hardness and durability.

Q7. What are the uses of Magnesium?

Ans: **Uses of Magnesium (Mg):**

i. Mg is low density metal, so it is used in the formation of light but tough alloys, such as Duralumin (a mixture of Al, Cu, Mg and Mn) Magnesium (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 (MgSO4) is used in textile, paper industry, soap formation and pharmaceutical industries etc.

Q8. Write down the reaction of chlorine with sodium hydroxide with balance equation.

Ans: **Reaction of chlorine with sodium hydroxide:**

Chlorine reacts with sodium hydroxide into two ways:
i. When chlorine is passed through the cold solution of sodium hydroxide, then the sodium hypochlorite is formed.
$Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$
ii. When chlorine is passed through the hot solution of sodium hydroxide, then the sodium chlorate is formed.
$3Cl_2 + 6NaOH \rightarrow 5NaCl + NaClO_3 + 3H_2O$

Q9. How does ionization energies values vary in a group?

Ans: As we know that the atomic size increases down the group due to increasing number of shell and the distance between the nucleus and valence shell electrons. As a result of force of attraction between the nucleus and valence electrons decreases so, less amount of energy will be required to remove and electron. So down the group ionization energy decreases.

Examples:

LE of Li = 520 kj/mol
I.E of Na = 496 kj/mol

Q10. What happens during displacement reaction in halogens?

Ans: During 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.

Examples:

$2NaBr + Cl_2 \rightarrow 2Nacl + Br_2$
$2Kl + Br_2 \rightarrow 2Kbr + I_2$