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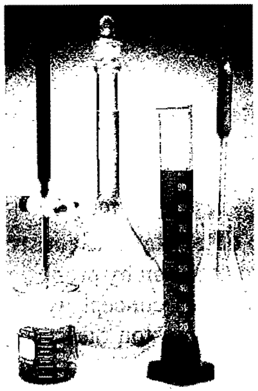


CHAPTER 1

BASIC PRINCIPLES

1.1 CHEMISTRY AND ITS SCOPE

Chemistry is a branch of physical science which deals with the study of matter, its physical and chemical properties, its chemical composition, the physical and chemical changes which it undergoes and the energy changes that accompany these processes.

All objects in this universe are composed of matter. Most of these objects are visible (solids and liquids) but some are invisible. Chemistry is termed as a material science because it is concerned with all material substances such as air, water, rocks, minerals, plants, animals including man, the earth on which we all live, and other planets. According to one of the famous scientists of twentieth century, **Linus Pauling,** *Chemistry is the science of substances, their properties, their structure and their transformations*.

Chemistry is a very interesting subject which touches almost every aspect of our lives, our culture and our environment. It has changed our civilization to a great extent. The present day chemistry has provided man with more comforts for a healthier and happier life. A large number of materials which we use these days were unknown at the turn of the century. A few decades back, our clothes and footwears were exclusively of natural origin such as vegetable fibres, wool, hair, skin of animals, etc., but, now the synthetic fibres produced in chemical factories have largely replaced them. Modern chemistry has given man new plastics, fuels, metal alloys, fertilizers, building materials, drugs, energy sources, etc.

During the last few decades, the expansion of chemistry has been tremendous. The field has become wide and complex. For convenience and better understanding of the subject, it has been divided into various branches. The four main branches of chemistry are:

(1) Organic chemistry; (2) Inorganic chemistry;

(3) Physical chemistry; (4) Analytical chemistry.

(1) Organic chemistry: It is concerned with the study of compounds of carbon except carbonates, bicarbonates, cyanides, isocyanides, carbides and oxides of carbon. It is actually the study of hydrocarbons and their derivatives.

(2) Inorganic chemistry: It deals with the study of all known elements and their compounds except organic compounds. It is concerned with the materials obtained from minerals, air, sea and soil.

(3) Physical chemistry: It is concerned with the physical properties and constitution of matter, the laws of chemical combination and theories governing reactions. The effect of temperature, pressure, light, concentration, etc., on reactions come under the scope of physical chemistry.

(4) Analytical chemistry: It deals with various methods of analysis of chemical substances both qualitative and quantitative. It includes chemical and physical methods of analysis.

A number of specialised branches have been introduced as to cope with the extraordinary expansion in the subject of chemistry. Some of the specialised branches are:

(i) Biochemistry: It comprises the studies of the substances related to living organisms and life processes.

(ii) Medicinal chemistry: It deals with the application of chemical substances for the prevention and cure of various diseases in living beings.

(iii) Soil and agriculture chemistry: It deals with the analysis and treatment of soils so as to increase its fertility for the better yields of crops. It is concerned with the chemicals used as fertilizers, insecticides, germicides, herbicides, etc.

(iv) Geochemistry: It includes the study of natural substances like ores and minerals, coal, petroleum, etc.

(v) Industrial chemistry: It deals with the study of chemical processes for the production of useful chemicals on a large scale at relatively low costs.

(vi) Nuclear chemistry: It is the most recent branch. It includes the study of nuclear reactions, the production of radioactive isotopes and their applications in various fields.

(vii) Structural chemistry: It deals with various techniques used for elucidation of the structure of chemical substances. It is concerned with the properties of substances in terms of their structure.

(viii) Polymer chemistry: It includes the study of chemical substances of very high molecular masses of the order of 100,000 or greater, called polymers—natural or artificial. This branch is gaining popularity as the use of plastics, rubber, synthetic fibres, silicones, etc., is on the increase these days.

(ix) Limnochemistry: It deals with the study of chemistry involved in the river water or water reservoirs.

(x) Phytochemistry: It includes the study of chemistry of plants.

Thus, it can be said that there is no other branch of science which is so wide in its scope as chemistry.

1.2 BRIEF HISTORY OF CHEMISTRY

It is difficult to specify the date when science of chemistry came into existence; however, its growth must have gone side by side with the growth of civilization. Broadly, the history of chemistry can be studied under five periods of its development.

(i) Ancient period up to 350 A.D.: In ancient times, many chemical operations such as souring of milk, conversion of sweet juices into wines, the conversion of wines into vinegar, etc., were known. Around 3000 B.C., techniques of making glass, pottery, pigments, dyes, perfumes and extraction of metals especially gold* and silver were known in China, India, Egypt and Greece. The beginning of chemistry as a science could probably be set about 400 B.C., when the theory was proposed that everything is composed of four elements: earth, air, fire and water. The first book of chemistry was written in Egypt around 300 A.D. The term chemistry meant the Egyptian art.

(ii) The alchemical period (350–1500): During this period, scientists called alchemists tried to discover two things: an elixir of life which could make man eternally young and a philospher's stone which could transmute base metals like zinc, copper, iron, etc., into gold. The alchemists failed in their efforts because no philospher's stone and elixir of life actually existed but we are indebted to them for designing new types of apparatus and for discovering new chemical operations such as distillation, sublimation, extraction of gold by amalgamation process and preparation of caustic alkalies from ashes of plants.

(iii) †Iatro chemistry period (1500–1650): During this era, chemists paid their attention towards medico-chemical problems. They believed that the primary object of chemistry was to prepare medicines and not to make gold from base metals. During this period, the study of gases was begun and quantitative experiments were undertaken for the first time. **Robert Boyle** (1627–1691) found that when a metal is heated in air, the mass increases. He also established the relationship between volume and pressure of a gas. In 1661, Boyle wrote the book 'The Skeptical Chymist' in which he criticised the basic ideas of alchemy.

(iv) The phlogiston period (1650–1774): The phlogiston theory was proposed by **Ernst Stahl** (1660–1734). Phlogiston was described as a substance in a combustible material which is given off when the material burns. This theory persisted for about 100 years and was a centre of much controversy. During the end of the eighteenth century, much work was done with gases, especially by **Joseph Black, Henry Cavendish, Joseph Priestley** and **Carl Scheele. Priestley** was a very conservative scientist. Even after his discovery of oxygen, he still believed in phlogiston theory.

(v) Modern period: Lavoisier (1743–1793), a French chemist, is regarded as the **father of modern chemistry.** He presented the exact explanation of combustion by proposing that oxygen is necessary for combustion. This concept was largely responsible for the overthrow of the phlogiston theory. Among his other contributions, he showed that water is composed of hydrogen and oxygen, proposed the theory of indestructibility of matter, presented a clear definition of an element and proposed a system of chemical nomenclature.

Another major step towards modern chemistry was taken in the first decade of the nineteenth century when the English chemist, **John Dalton,** postulated that all elements are made up of atoms. He pictured atoms as tiny, indestructible units that could combine to form **compound atoms** or **molecules.** Dalton proposed that each element has its own kind of atoms and the atoms of different elements differ in essentially nothing but their masses. He determined the relative masses of atoms of many elements. Thus, a new era had begun. The other important chemists of this period are:

(a) **Richter**—Law of Reciprocal Proportions (1794)
(b) **Proust**—Law of Definite Proportions (1799)
(c) **Gay-Lussac**—Law of Combining Volumes of Gases (1808)
(d) **Avogadro**—Avogadro Hypothesis (1811)
(e) **Berzelius**—Introduced the Modern Symbols for Elements (1813)
(f) **Faraday**—Laws of Electrolysis (1833)
(g) **Thomas Graham**—Law of Gaseous Diffusion (1861)
(h) **Mendeleev**—Periodic Law and Periodic Table (1869)
(i) **Arrhenius**—Theory of Ionization (1887)
(j) **Henry Becquerel**—Discovery of Radioactivity (1896)
(k) **Madam Curie**—Discovered Radium and Polonium (1898)

The twentieth century is regarded as an active era of chemistry. During this period, chemistry has made many contributions to human knowledge and civilization. Now, we live in a world of synthetic materials. Chemistry of today is actually helping in solving major problems of our present day civilization such as population explosion, food and diseases, depletion of sources of energy, depletion of natural sources and environmental pollution.

1.3 MATTER AND ENERGY

Besides life, matter and energy are regarded the two fundamental entities with which whole of the universe is composed of. **Matter is anything that has mass and occupies space.** All bodies in the universe conform to this definition. Mass is the quantity of matter in a particular sample of matter. Mass of a body is constant and does not change regardless of where it is measured. The mass of a

*Gold was probably the first metal to be used because it occurred as a free metal in the earth.
†*Iatro* is a Greek word meaning a physician.

body would be the same on the moon as it is on the earth. Our senses of sight and touch usually tell us that an object occupies space except in the case of colourless, odourless and tasteless gases where some other evidence is required to satisfy the definition of matter.

The term **weight** should not be used in place of mass as it has a different meaning. The term **weight** refers to the force with which an object is attracted towards earth. An object resting on earth experiences a force called its weight, W, that is equal to its mass m, multiplied by the acceleration due to gravity g, that is,

$$W = mg$$

The weight of an object thus depends on the value of 'g' which varies from place to place. However, the mass of an object is determined by comparing the weights of two objects, one of known mass, the other of unknown mass in the same location on earth as both experience the same gravitational acceleration.

Matter is indestructible, *i.e.*, it can neither be created, nor destroyed, but it can change its form; thus, the total quantity of matter of the universe is constant.

Energy is defined as the capacity of doing work. Anything which has the capacity to push the matter from one place to another possesses energy. There are various forms of energy such as heat, light, etc. Energy is neither created, nor destroyed, but can only be transformed from one form of energy to another.

The world became aware of the fact that matter can be converted into energy with the discovery of nuclear reactions, especially nuclear fission and nuclear fusion. The relationship between mass and energy was given by **Einstein.** The famous relation is:

$$E = mc^2$$

where, E = energy, m = mass and c = velocity of light.

On account of this equation, the above two laws are amalgamated into a single statement:

"The total amount of matter and energy available in the universe is fixed."

Example 1. *Calculate the amount of energy released in ergs, calories and in joules when 0.001 kg of mass disappears.*

[Given, Velocity of light $= 3 \times 10^8\ ms^{-1}$]

Solution: According to Einstein equation, $E = mc^2$

$$m = 0.001\text{ kg} = 1 \times 10^{-3}\text{ kg};\ c = 3 \times 10^8\text{ ms}^{-1}$$

$$E = (1 \times 10^{-3})(3 \times 10^8)^2 = 9 \times 10^{13}\text{ J}$$

$$1\text{ J} = 10^7\text{ erg} = 0.24\text{ cal}$$

$$9 \times 10^{13}\text{ J} = 9 \times 10^{13} \times 10^7\text{ erg} = 9 \times 10^{13} \times 0.24\text{ cal}$$

$$= 9 \times 10^{20}\text{ erg} = 2.16 \times 10^{13}\text{ cal}$$

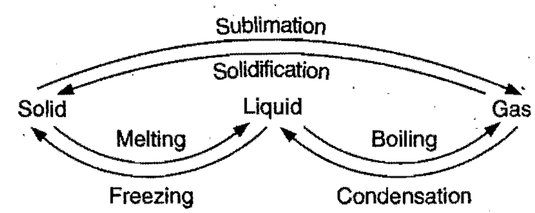
Classification of Matter

(i) Physical classification: Matter can exist in any one of three forms, (a) solid, (b) liquid and (c) gas.

In the solid state, substances are rigid. They have a definite shape and fixed volume. There is negligible effect of changes in pressure and temperature on their volumes. The individual particles that make up a solid occupy definite positions in the structure and are very near to one another. This form of matter is associated with minimum amount of energy.

In liquid state, substances have no definite shape but possess a fixed volume. There is slight effect of pressure and temperature on their volumes. They have the property of flowing. The particles are nearer to one another than in a gas and this form of matter is associated with energy more than solids.

In a gaseous state, substances have no definite shape and volume. Gases fill completely any vessel in which they are confined and thus occupy the whole space available to them. There is a large effect of pressure and temperature on their volumes. The particles are far apart from one another and move with very high speeds in all possible directions. This form of matter is associated with maximum amount of energy.

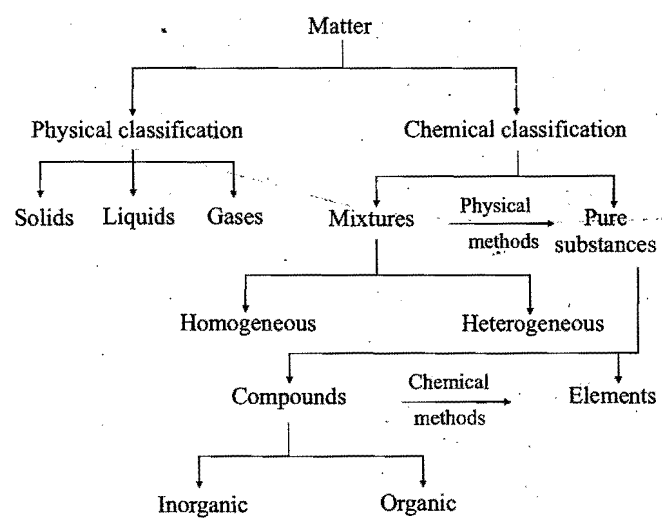


$$\text{Solid} \underset{-\text{ energy}}{\overset{+\text{ energy}}{\rightleftharpoons}} \text{Liquid} \underset{-\text{ energy}}{\overset{+\text{ energy}}{\rightleftharpoons}} \text{Gas}$$

$$\underset{\text{Release of energy}}{\overset{\text{Absorption of energy}}{\rightleftharpoons}}$$

Depending on temperature and pressure, a substance can exist in any one of the three forms of matter.

(ii) Chemical classification: Matter exists in nature in the form of chemical substances. **A pure substance is defined as a variety of matter, all samples of which have same composition and properties.** Pure substances are divided into elements and compounds. Most of the materials found in nature are in the form of mixtures consisting of two or more substances. There are two types of mixtures—Homogeneous and Heterogeneous. Both types of mixtures can be separated into their components (pure substances) by mechanical and physical methods. The classification can be summarized in the following way:



Properties of Matter: Properties are the characteristic qualities with the help of which different kinds of matter can be commonly recognised. In chemistry, substances are distinguished by two types of properties, *viz* (i) Chemical properties and (ii) Physical properties.

The chemical properties of substances are those in which they undergo change in composition either alone or by interactions with other substances, *i.e.*, to form new substances having different compositions from the substances which undergo change.

The properties of substances which are observed in absence of any change in composition under specific physical state are termed physical properties. Colour, density, melting point, boiling point, hardness, refractive index, thermal conductivity, electrical conductivity, malleability, ductility, etc., are some examples of the physical properties. The properties of matter can be further classified into two: (i) Intensive properties and (ii) Extensive properties. The **intensive properties** are those which do not depend upon the quantity of matter, *e.g.*, colour, density, melting point, boiling point, refractive index, etc. These properties are same irrespective of the quantity of the substance. Chemical properties are also intensive properties. The **extensive properties** of matter depend on the quantity of matter. Volume, mass, weight, energy, etc., are the extensive properties.

1.4 ELEMENTS AND COMPOUNDS

Elements are pure substances that cannot be decomposed into simpler substances by chemical changes. The smallest particles of an element possess the same properties as the bigger particles. An element can also be defined as **a pure substance which consists of only one type of atoms.** Due to discovery of isotopes, this definition does not seem to be correct. The modern definition of an element is that it is a simple individual which has a definite atomic number (see atomic structure) and has a definite position in the periodic table. It cannot be decomposed in a chemical change. In chemistry, the elements are the chemical alphabet and compounds are the words, *i.e.*, combinations of elements.

There are presently 117 different elements known. Every element has been given a definite name and for convenience a nick name which in chemical language is called a **symbol.** *Symbol is a small abbreviation to represent a full and lengthy name of the element.* Symbols have been derived:

(i) either by taking the first letter of the name of the element which is capitalized:

O—Oxygen	N—Nitrogen	F— Fluorine
C—Carbon	H—Hydrogen	U— Uranium
P—Phosphorus	S—Sulphur	I— Iodine

(ii) or by taking the first letter and one more letter from the name of the element. The first letter is always capitalized.

Ca—Calcium	Ni—Nickel	Al—Aluminium
Mg—Magnesium	Co—Cobalt	Bi—Bismuth
Cl—Chlorine	Br—Bromine	Ba—Barium

(iii) or from names of the elements in other languages such as Latin, German, etc.

Na—Sodium (Latin name Natrium)

Cu—Copper (Latin name Cuprum)

Fe—Iron (Latin name Ferrum)

Ag—Silver (Latin name Argentum)

Pb—Lead (Latin name Plumbum)

Au—Gold (Latin name Aurum)

K—Potassium (Latin name Kalium)

Hg—Mercury (Latin name Hydragyrum)

W—Tungsten (German name Wolfram)

Out of 117 elements known, 88 have been isolated from natural sources and the remaining have been prepared by artificial means. The man made elements are:

S.No.	Name	Symbol	S.No.	Name	Symbol
1.	Neptunium	Np	16.	Hassium or Unniloctium	Hs or Uno
2.	Plutonium	Pu	17.	Meitnerium or Unnilennium	Mt or Une
3.	Americium	Am	18.	Ununnilium	Uun
4.	Curium	Cm	19.	Unununium	Uuu
5.	Berkelium	Bk	20.	Ununbium	Uub
6.	Californium	Cf	21.	Ununtrium	Uut
7.	Einsteinium	Es	22.	Ununquadium	Uuq
8.	Fermium	Fm	23.	Ununpentium	Uup
9.	Mendelevium	Md	24.	Ununhexium	Uuh
10.	Nobelium	No	25.	Ununoctium	Uuo
11.	Lawrencium	Lr	26.	Technetium	Tc
12.	Kurchatovium	Ku	27.	Promethium	Pm
13.	Hahnium	Ha	28.	Astatine	At
14.	Seaborgium or Unnilhexium	Sg or Unh	29.	Francium	Fr
15.	Nielsbohrium or Unnilseptium	Bh or Uns			

The elements from S. No. 1 to 25 are called transuranic elements. The credit for the discovery of most of the transuranic elements goes to the scientist **G.T. Seaborg.** The first artificially produced element was technetium. It was synthesised in 1937 by scientists at the University of California at Berkley.

Most of the earth's crust is made up of a small number of elements. Only ten of the naturally occurring elements make up 99% mass of the earth's crust, oceans and atmosphere. The following table shows the abundance of highly abundant elements in nature:

Note : Among the naturally occurring elements, ${}_{1}^{1}H$ is lightest and ${}_{92}^{238}U$ is the heaviest atom.

Abundance of Elements (Earth's Crust, Oceans and Atmosphere)

Oxygen	49.5%	Chlorine	0.19%
Silicon	25.7%	Phosphorus	0.12%
Aluminium	7.5%	Manganese	0.09%
Iron	4.7%	Carbon	0.08%
Calcium	3.4%	Sulpur	0.06%
Sodium	2.6%	Barium	0.04%
Potassium	2.4%	Chromium	0.033%
Magnesium	1.9%	Nitrogen	0.030%
Hydrogen	0.87%	Fluorine	0.027%
Titanium	0.58%	Zirconium	0.023%
	99.2%		0.7%

If the entire universe is considered, then 90% of matter is hydrogen. Helium is the second most abundant element amounting to 9% and the remaining elements make up only 1% of the universe with oxygen, neon, carbon and nitrogen next in order of decreasing abundance.

The commercial use of an element depends not only upon its abundance but also upon its accessibility. Some of the common elements such as copper, zinc, tin and lead are not abundant but are found in nature in rich deposits from which these can be easily extracted. On the other hand, the elements such as titanium and zirconium which are found in abundance in nature are not widely used because their ores are not rich and their extraction is difficult and expensive.

Metals, Non-metals and Metalloids

All the elements may be classified into two groups, **metals** and **non-metals.** The division is based on both physical and chemical properties.

Metals are regarded as those elements which possess the following properties:

(i) They are generally solids at ordinary conditions. Mercury is an exception which is in liquid state.
(ii) They are lustrous in nature.
(iii) They possess high density.
(iv) They are good conductors of electricity and heat.
(v) They are malleable and ductile.
(vi) They possess generally high melting and boiling points.
(vii) They react with mineral acids liberating hydrogen.
(viii) They form basic oxides.
(ix) They form non-volatile hydrides if combine with hydrogen.
(x) They have molecules usually mono-atomic in the vapour state.

Sodium, calcium, aluminium, copper, silver, zinc, iron, nickel, gold, mercury, etc., are the examples of metals.

The non-metals do not show the above properties. Six of the non-metals, carbon, boron, phosphorus, sulphur, selenium and iodine, are solids. Bromine is the only liquid non-metal at room temperature and normal pressure. The remaining non-metals; nitrogen, oxygen, fluorine, chlorine, hydrogen, helium, argon, neon, krypton, xenon and radon are gases. Non-metals are generally (i) brittle, (ii) non-lustrous, (iii) having low melting and boiling points, (iv) non-conductors of heat, (v) capable of forming acidic oxides or neutral oxides, (vi) not capable of evolving hydrogen from acids, and (vii) capable of forming volatile hydrides.

There are some elements which do not fit completely into either the metal or non-metal class. Elements which have some properties of both metals and non-metals are called semi-metals or metalloids. The semi-metals are silicon, germanium, arsenic, antimony and tellurium.

The above classification of elements is a rough one as certain metals like lithium, sodium, potassium possess low density; certain non-metals like hydrogen and graphite (a form of carbon) are good conductors of electricity. Metals rarely combine with one another while non-metals combine with one another to form compounds. Metals and non-metals commonly combine with each other to form compounds.

Compounds

Compounds are also pure substances that are composed of two or more different elements in a fixed proportion by mass. Compounds containing more than four elements are rare. The properties of a compound are altogether different from the properties of the elements from which it has been constituted. The compound **water** has a definite composition, *i.e.*, 11.2% hydrogen and 88.8% oxygen, Thus, the two are present in the ratio of 1 : 8 by mass. The properties of water are totally different from the properties of hydrogen and oxygen both. Hydrogen and oxygen are in gaseous state while water is in liquid state under ordinary atmospheric conditions. Oxygen supports combustion while hydrogen is combustible but water is normally used for extinguishing fire. Component elements in compounds can be separated only by chemical means and not by physical methods.

Compounds are classified into two types:

(i) Organic compounds: The compounds obtained from living sources are termed organic compounds. The term organic is now applied to hydrocarbons (compounds of carbon and hydrogen) and their derivatives.

(ii) Inorganic compounds: The compounds obtained from non-living sources such as rocks and minerals are termed inorganic compounds. The compounds of all elements except hydrocarbons and their derivatives are included in this category. The number of organic compounds is very large in comparison to inorganic compounds.

Some Specific Properties of Substances: Some specific properties of substances are given below:

(i) Deliquescence: The property of certain compounds of taking up the moisture present in atmosphere and becoming wet when exposed, is known as deliquescence. These compounds are known as deliquescent. Sodium hydroxide, potassium hydroxide, anhydrous calcium chloride, anhydrous magnesium chloride, anhydrous ferric chloride, etc., are the examples of deliquescent compounds. Sodium chloride is not deliquescent but when common salt is placed in atmosphere it becomes wet due to presence of an impurity of magnesium chloride.

(ii) Hygroscopicity: Certain compounds combine with the moisture of atmosphere and are converted into hydroxides or hydrates. Such substances are called hygroscopic. Anhydrous copper sulphate, quick lime (CaO), anhydrous sodium carbonate, etc., are of hygroscopic nature.

(iii) Efflorescence: The property of some crystalline substances of losing their water of crystallisation on exposure and becoming powdery on the surface is called efflorescence. Such salts are known as efflorescent. The examples are:

Ferrous sulphate ($FeSO_4 \cdot 7H_2O$), sodium carbonate ($Na_2CO_3 \cdot 10H_2O$), sodium sulphate ($Na_2SO_4 \cdot 10H_2O$), potash alum [$K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$], etc.

(iv) Malleability: This property is shown by metals. When the solid is beaten and does not break but is converted into a thin sheet, it is said to possess the property of malleability. Copper, gold, silver, aluminium, lead, etc., can be easily hammered into sheets. Gold is the most malleable metal.

(v) Ductility: The property of a metal to be drawn into wires is termed ductility. Copper, silver, gold, aluminium, iron, etc., are ductile in nature. Platinum is the most ductile metal.

(vi) Elasticity: When the stress is small, the solid completely regains its original shape, size or volume after the deforming force is removed. The solid is then said to be elastic. Steel, glass, ivory, etc., are elastic bodies.

(vii) Plasticity: When stress is increased on a metal, a limit is reached beyond which, if the stress is removed, the solid does not come back to its original shape or size. It acquires a permanent deformation. Such materials can be given any shape without any difficulty.

(viii) Brittleness: The solid materials which break into small pieces on hammering are called brittle. The solids of non-metals are generally brittle in nature.

(ix) Hardness: A material is said to be harder than the other if it can scratch it. The hardness is measured on Mho's scale. For this purpose, ten minerals have been selected which have been assigned hardness from 1 to 10.

Hardness	Mineral	Hardness	Mineral
1	Talc	6	Orthoclase
2	Gypsum	7	Quartz
3	Calcite	8	Topaz
4	Fluorite	9	Corundum
5	Apatite	10	Diamond

On Mho's scale, hardness of diamond is maximum and that of talc is minimum. If a material can scratch topaz but cannot scratch corundum it possesses hardness equal to 8.

1.5 MIXTURES

A **mixture** is a material containing two or more substances either elements or compounds or both in any proportion. Substances which form a mixture are called **components.** Components are present in the mixture without loss of their identity. There are two types of mixtures—**homogeneous** and **heterogeneous.** In a homogeneous mixture, the components are mixed uniformly to microscopic level. The components cannot be seen by naked eye or with the help of a microscope. The mixture is uniform throughout having a single **phase*.** The homogeneous mixture is **isotropic** in nature, *i.e.*, every portion of it has the same composition and properties.

Alloys such as brass, steel, 22-carat gold; solutions such as common salt dissolved in water, sugar dissolved in water, iodine dissolved in carbon tetrachloride, benzene in toluene, methyl alcohol in water; gasoline (a mixture of hydrocarbons), air, etc., are some of the examples of homogeneous mixtures.

A heterogeneous mixture is not uniform. It can have two or more phases. The components can be seen by naked eye or with the help of a microscope. It has **anisotropic** properties, *i.e.*, properties are not uniform throughout the mixture. Soil, a mixture of sulphur and sand, a mixture of iron filings and sand, smoke, etc., are the examples of heterogeneous mixtures.

The components of a mixture differ in many of their physical and chemical properties. The advantage of this difference is taken in the separation of a mixture. The method of separation employed should not bring about the destruction of any one of the components. Some preliminary techniques based on physical properties are described here in brief.

(i) Filtration: This method is useful when one of the components is an insoluble solid in a solvent. The insoluble solid is obtained by filtration of the suspension through filter-paper. For example, common salt containing sand is separated by filtration. The mixture is mixed with water. It is shaken so as to dissolve common salt. The sand remains insoluble. The suspension is put to filtration. The sand collects on the filter-paper. It is taken in a basin and dried by heating. The filtrate is taken in evaporating dish and heated till whole of the water is evaporated. Solid common salt is thus obtained in the dish.

Sugar containing charcoal, potassium nitrate containing saw dust or mixtures having insoluble components can be separated by filtration.

(ii) Sublimation: It is a process in which a solid substance is directly converted into its vapours by application of heat and vapour is reconverted into solid by subsequent cooling. The method is used when one of the components undergoes sublimation and other components are not decomposed by heating. For example, naphthalene can be separated from common salt by sublimation. Similarly, a mixture of ammonium chloride and potassium chloride can be separated by sublimation as ammonium chloride sublimes on heating.

(iii) Distillation: It is a process of converting a liquid into its vapour by heating and then condensing the vapours again into the same liquid by cooling. Thus, distillation involves vaporisation and condensation both.

$$\text{Distillation} = \text{Vaporisation} + \text{Condensation}$$

This method is employed to separate liquids which have different boiling points or a liquid from non-volatile solid or solids either in solution or suspension. The mixture of copper sulphate and water or mixture of water (b.p. 100°C) and methyl alcohol (b.p. 45°C) can be separated by this method.

* Phase is defined as part of a system which has uniform properties and composition. A solution or mixture of sugar and water is a one phase system. Every drop of the solution has same properties and same composition.

(iv) Magnetic separation: If one of the components of a mixture has magnetic properties, it can be separated by using a magnet. Iron is separated from a mixture of iron filings and sulphur by moving a magnet through the mixture.

(v) Solvent extraction: This method is based on the preferential solubility of one of the components of the mixture in a particular solvent (usually a low-boiling organic solvent) which forms a distinctly separate layer with the other liquid if present in the mixture. For example, iodine present in water can be recovered with the help of ether or carbon disulphide. For this method, a separating funnel is utilized. The aqueous solution of iodine is taken in separating funnel to which ether is added. The funnel is shaken. Two layers are formed. The upper layer which is dark brown consists of ether and iodine and the colourless lower layer consists of only water. The lower layer is taken out. The coloured layer is then poured out and ether is removed cautiously by distillation when iodine is left behind.

Two immiscible liquids such as water and oil can also be separated by the use of a separating funnel.

Example 2. *How will you separate the following mixtures?*

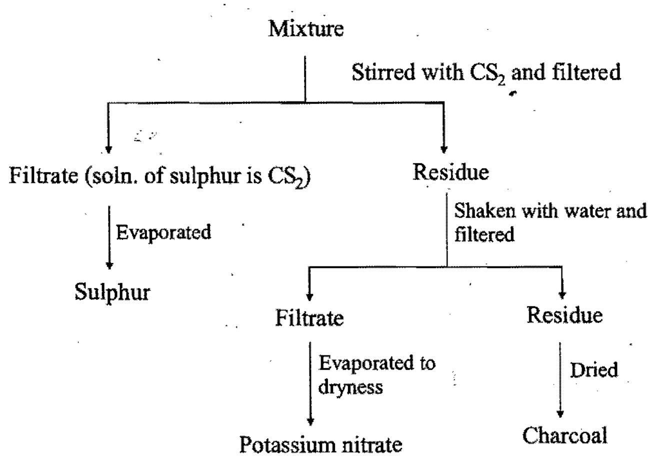
(*a*) *Sulphur, potassium nitrate and charcoal,*

(*b*) *Sand, common salt, iron filings and naphthalene,*

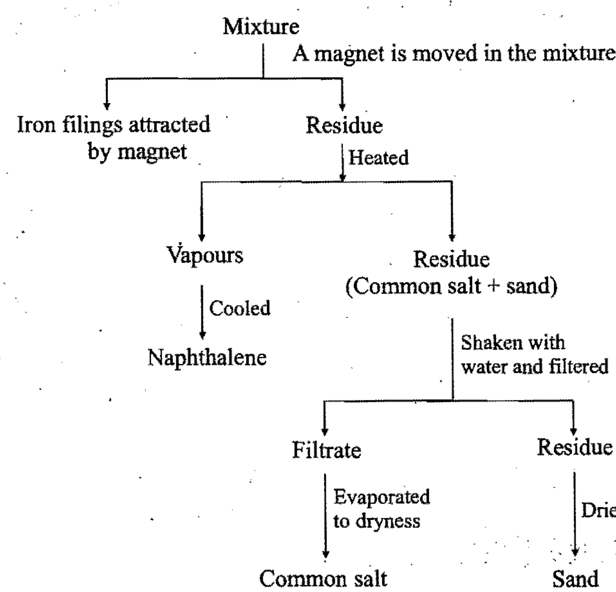
(*c*) *Powdered glass, ammonium chloride and potassium chloride.*

Solution:

(a) (i) Sulphur is soluble in carbon disulphide,
(ii) Potassium nitrate is soluble in water,
(iii) Charcoal is insoluble in carbon disulphide as well as in water.



(b) (i) Iron filings are separated by a magnet,
(ii) Naphthalene sublimes on heating,
(iii) Sand is insoluble in water.



(c) (i) Ammonium chloride sublimes on heating,
(ii) Potassium chloride is soluble in water,
(iii) Powdered glass is insoluble in water.

1.6 ALLOYS

When two or more elements are melted together and resulting liquid is allowed to solidify, the product so obtained is called an alloy if it possesses metallic properties. An alloy may consist of a mixture of a metal with another metal, a metal with a non-metal or a metal with both metal and non-metal.

Alloys are prepared because they have properties more suitable for certain applications than do the simple metals. Alloys are used because they are harder and stronger, have desirable casting properties, special physical properties such as magnetic properties and resistance to corrosion in certain environments. Melting point of an alloy is normally lower than the melting point of either of the pure components. Thermal and electrical conductivities are normally reduced in alloys.

An alloy containing one component mercury is called amalgam. Most of the metals form amalgams. Iron, platinum, tungsten, etc., are few metals which do not form amalgams.

Alloys are mainly classified into two distinct types, namely ferrous and non-ferrous. Ferrous alloys always contain iron, carbon and one or two of the other elements such as manganese, nickel, chromium, copper, vanadium, molybdenum, tungsten, etc. When the percentage of carbon in the alloy is below 0.1, the alloy is termed the iron alloy and if it is above 0.1, the alloys are called steels. When iron is not present in the alloy, it is termed a non-ferrous alloy. Some of the important alloys have been listed below:

Alloy	Composition	Main uses
1. Brass	Cu 60–80%, Zn 20–40%	Utensils, condenser tubes, electrical goods, cartridge shell
2. Bronze	Cu 75–90%, Sn 10–25%	Coins, statues, utensils.

3. German silver	Cu 56%, Zn 24%, Ni 20%	Utensils, resistance coils
4. Gun metal	Cu 87%, Sn 10%, Zn 3%	Machine parts, guns
5. Rolled gold	Cu 95%, Al 5%	Artificial jewellery
6. Magnalium	Al 94%, Mg 6%	Balance beams, light instruments
7. Electron	Mg 95%, Zn 5%	Construction of aircraft
8. Duralumin	Al 95%, Cu 4%, Mn 0.5%, Mg 0.5%	Making aeroplanes
9. Type metal	Pb 82%, Sb 15%, Sn 3%	Making printing types
10. Solder	Pb 50–70%, Sn 30–50%	Soldering
11. Britannia	Sn 93%, Sb 5%, Cu 2%	Tableware
12. Wood's metal	Bi 50%, Pb 25%, Sn 12.5%, Cd 12.5%	Electric fuses and other safety devices
13. Nichrome	Ni 60%, Cr 15%, Fe 25%	Electrical resistances
14. Constantan	Ni 40%, Cu 60%	Electrical resistances
15. Monel metal	Ni 70%, Cu 30%	Chemical plants
16. Invar	Ni 35%, Steel 65%	Surveying instruments, pendulums, chronometers
17. Stainless steel	Fe 89.4%, Cr 10%, Mn 0.35%, C 0.25%	Utensils, ornamental pieces.

1.7 PHYSICAL AND CHEMICAL CHANGES

Matter undergoes two types of changes; physical and chemical. A physical change is one in which a substance changes its physical state but keeps its chemical identity. In physical change, a new substance does not come into existence. Mass remains the same. Physical properties are altered. This is a temporary change. For example, water shows all of its chemical properties whether it is in the form of ice or water or steam. Ice melts to form water and water can be converted again into ice by placing it in a freezer. When 10 g of ice melts, 10 g of water is obtained. Melting, evaporation, condensation, freezing, sublimation, distillation, passing of electric current through metallic conductor, making of magnet from an iron piece, are some examples of physical changes.

In a chemical change, a new substance or substances come into existence. The starting materials called reactants, are used up and new substances called products, are formed. The composition of the new substances is different from that of the starting materials. It is a permanent change as it is not easy to obtain the starting materials again from the products.

Energy is always released or absorbed when chemical or physical changes occur. Energy is required to melt ice and energy is required to boil water. Conversely, the condensation of steam to form liquid water always liberates energy, as does the freezing of liquid water to form ice. Chemical changes either release energy (exothermic) or adsorb energy (endothermic).

Chemical changes are of various types. The important ones are:

(i) Combination: Two or more substances react to form one product. When a compound is obtained by the direct reaction between elements, it is termed direct union or synthesis.

$$H_2 + Cl_2 = 2HCl \quad ; \quad C + O_2 = CO_2$$

$$2Mg + O_2 = 2MgO \quad ; \quad SO_2 + H_2O = H_2SO_3$$

(ii) Decomposition: When a compound is broken down into two or more simple constituents, the change is called decomposition. Often heat is utilised for the decomposition. Such decomposition is termed **thermal decomposition.**

$$2HgO = 2Hg + O_2 \; ; \quad 2KClO_3 = 2KCl + 3O_2$$

$$CaCO_3 = CaO + CO_2 ; \; 2NaHCO_3 = Na_2CO_3 + H_2O + CO_2$$

(iii) Substitution: When one element enters into a compound by the replacement of the other element, the change is termed substitution.

$$CuSO_4 + Zn = Cu + ZnSO_4 ; \; C_2H_6 + Cl_2 = C_2H_5Cl + HCl$$

$$2KI + Cl_2 = 2KCl + I_2 \quad ; \quad Zn + 2HCl = ZnCl_2 + H_2$$

(iv) Addition: Something is added to a chemical substance without elimination.

$$KI + I_2 = KI_3$$

$$C_2H_4 + HBr = C_2H_5Br$$

$$CH_3CHO + HCN = CH_3CH\begin{matrix} OH \\ CN \end{matrix}$$

(v) Internal rearrangement: When nothing is added or nothing is eliminated from a chemical substance but due to rearrangement of the various atoms present in a molecule, a new compound comes into existence. When ammonium cyanate is heated, a new substance urea is formed.

$$NH_4CNO = NH_2CONH_2$$

The chemical change is termed isomerisation, when one isomer is converted into another.

$$\underset{n\text{-Butane}}{CH_3-CH_2-CH_2-CH_3} \xrightarrow[300°C]{AlCl_3/HCl} \underset{2\text{-Methyl propane}}{CH_3-\overset{\overset{\displaystyle CH_3}{|}}{CH}-CH_3}$$

(vi) Polymerization: Two or more molecules of a substance combine to form a giant molecule,

$$\underset{\text{Acetylene}}{3C_2H_2} = \underset{\text{Benzene}}{C_6H_6}$$

$$\underset{\text{Formaldehyde}}{nHCHO} = \underset{\text{Para formaldehyde}}{(HCHO)_n}$$

(vii) Double decomposition: An exchange of partners occurs between two compounds.

$$AB + CD = AD + CB$$

$$BaCl_2 + Na_2SO_4 = BaSO_4 + 2NaCl$$

$$KOH + HCl = KCl + H_2O$$

The reaction is also termed neutralisation, *i.e.*, a reaction between an acid and a base to form a salt and water molecule.

$$\underbrace{FeCl_3 + 3HOH}_{\text{Water}} = Fe(OH)_3 + 3HCl$$

Reaction of above type is termed hydrolysis.

1.8 LAWS OF CHEMICAL COMBINATION

In order to understand the composition of various compounds, it is necessary to have a theory which accounts for both qualitative and quantitative observations during chemical changes. Observations of chemical reactions were most significant in the development of a satisfactory theory of the nature of matter. These observations of chemical reactions are summarised in certain statements known as laws of chemical combination.

(i) Law of conservation of mass: The law was first stated by **Lavoisier** in 1774. It is also known as the law of indestructibility of matter. According to this law, **in all chemical changes, the total mass of a system remains constant or in a chemical change, mass is neither created nor destroyed.** This law was tested by **Landolt.** All chemical reactions follow this law. Thus, this law is the basis of all quantitative work in chemistry.

Example: *1.70 g of silver nitrate dissolved in 100 g of water is taken. 0.585 g of sodium chloride dissolved in 100 g of water is added to it and chemical reaction occurs. 1.435 g of silver chloride and 0.85 g of sodium nitrate are formed.*

Solution: Total masses before chemical change

$$= \text{Mass of } AgNO_3 + \text{Mass of NaCl} + \text{Mass of water}$$

$$= 1.70\,g + 0.585\,g + 200.0\,g$$

$$= 202.285\ g$$

Total masses after the chemical reaction,

$$= \text{Mass of AgCl} + \text{Mass of } NaNO_3 + \text{Mass of water}$$

$$= 1.435\,g + 0.85\,g + 200.0\,g$$

$$= 202.285\,g$$

Thus, in this chemical change,

Total masses of reactants = Total masses of products

This relationship holds good when reactants are completely converted into products.

In case, the reacting materials are not completely consumed, the relationship will be

Total masses of reactants = Total masses of products + Masses of unreacted reactants

(ii) Law of definite or constant proportions: This law was presented by **Proust** in 1799 and may be stated as follows:

A chemical compound always contains the same element combined together in fixed proportion by mass, *i.e.*, a chemical compound has a fixed composition and it does not depend on the method of its preparation or the source from which it has been obtained.

For example, carbon dioxide can be obtained by using any one of the following methods:

(a) by heating calcium carbonate,
(b) by heating sodium bicarbonate,
(c) by burning carbon in oxygen,
(d) by reacting calcium carbonate with hydrochloric acid.

Whatever sample of carbon dioxide is taken, it is observed that carbon and oxygen are always combined in the ratio of 12 : 32 or 3 : 8.

The converse of this law that when same elements combine in the same proportion, the same compound will be formed, is not always true. For example, carbon, hydrogen and oxygen when combine in the ratio of 12 : 3 : 8 may form either ethyl alcohol (C_2H_5OH) or dimethyl ether (CH_3OCH_3) under different experimental conditions.

(iii) Law of multiple proportions: This law was put forward by **Dalton** in 1808. According to this law, **if two elements combine to form more than one compound, then the different masses of one element which combine with a fixed mass of the other element, bear a simple ratio to one another.**

Hydrogen and oxygen combine to form two compounds H_2O (water) and H_2O_2 (hydrogen peroxide).

In water,	Hydrogen 2 parts	Oxygen 16 parts
In hydrogen peroxide,	Hydrogen 2 parts	Oxygen 32 parts

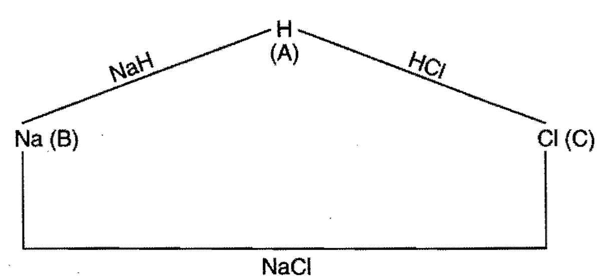
The masses of oxygen which combine with same mass of hydrogen in these two compounds bear a simple ratio 1 : 2.

Nitrogen forms five stable oxides.

N_2O	Nitrogen 28 parts	Oxygen 16 parts
N_2O_2	Nitrogen 28 parts	Oxygen 32 parts
N_2O_3	Nitrogen 28 parts	Oxygen 48 parts
N_2O_4	Nitrogen 28 parts	Oxygen 64 parts
N_2O_5	Nitrogen 28 parts	Oxygen 80 parts

The masses of oxygen which combine with same mass of nitrogen in the five compounds bear a ratio 16 : 32 : 48 : 64 : 80 or 1 : 2 : 3 : 4 : 5.

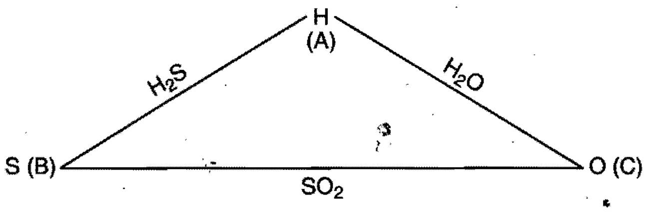
(iv) Law of reciprocal proportions: This law was given by **Richter** in 1794. The law states that **when definite mass of an element *A* combines with two other elements *B* and *C* to form two compounds and if *B* and *C* also combine to form a compound, their combining masses are in same proportion or bear a simple ratio to the masses of *B* and *C* which combine with a constant mass of *A*.**



For example, hydrogen combines with sodium and chlorine to form compounds NaH and HCl respectively.

In NaH,	Sodium 23 parts	Hydrogen one part
In HCl,	Chlorine 35.5 parts	Hydrogen one part

Sodium and chlorine also combine to form NaCl in which 23 parts of sodium and 35.5 parts of chlorine are present. These are the same parts which combine with one part of hydrogen in NaH and HCl respectively.



Hydrogen combines with sulphur and oxygen to form compounds H_2S and H_2O respectively.

In H_2S,	Hydrogen 2 parts	Sulphur 32 parts
In H_2O,	Hydrogen 2 parts	Oxygen 16 parts

Thus, according to this law, sulphur should combine with oxygen in the ratio of 32 : 16 or a simple multiple of it. Actually, both combine to form SO_2 in the ratio of 32 : 32 or 1 : 1.

The law of reciprocal proportions is a special case of a more general law, the law of equivalent masses, which can be stated as under :

"In all chemical reactions, substances always react in the ratio of their equivalent masses."

(v) Law of gaseous volumes: This law was enunciated by **Gay-Lussac** in 1808. According to this law, gases react with each other in the simple ratio of their volumes and if the product is also in gaseous state, the volume of the product also bears a simple ratio with the volumes of gaseous reactants when all volumes are measured under similar conditions of temperature and pressure.

$$\underset{1\text{ vol}}{H_2} + \underset{1\text{ vol}}{Cl_2} = \underset{2\text{ vol}}{2HCl} \qquad \text{ratio} \qquad 1:1:2$$

$$\underset{2\text{ vol}}{2H_2} + \underset{1\text{ vol}}{O_2} = \underset{2\text{ vol}}{2H_2O} \qquad \text{ratio} \qquad 2:1:2$$

$$\underset{2\text{ vol}}{2CO} + \underset{1\text{ vol}}{O_2} = \underset{2\text{ vol}}{2CO_2} \qquad \text{ratio} \qquad 2:1:2$$

$$\underset{1\text{ vol}}{N_2} + \underset{3\text{ vol}}{3H_2} = \underset{2\text{ vol}}{2NH_3} \qquad \text{ratio} \qquad 1:3:2$$

SOME SOLVED EXAMPLES

Example 3. *What mass of sodium chloride would be decomposed by 9.8 g of sulphuric acid, if 12 g of sodium bisulphate and 2.75 g of hydrogen chloride were produced in a reaction assuming that the law of conservation of mass is true?*

Solution: $NaCl + H_2SO_4 = NaHSO_4 + HCl$

According to law of conservation of mass,

Total masses of reactants = Total masses of products

Let the mass of NaCl decomposed be x g, so

$$x + 9.8 = 12.0 + 2.75$$
$$= 14.75$$
$$x = 4.95\text{ g}$$

Example 4. *In an experiment, 2.4 g of iron oxide on reduction with hydrogen yield 1.68 g of iron. In another experiment, 2.9 g of iron oxide give 2.03 g of iron on reduction with hydrogen. Show that the above data illustrate the law of constant proportions.*

Solution:

In the first experiment

The mass of iron oxide = 2.4 g

The mass of iron after reduction = 1.68 g

The mass of oxygen = Mass of iron oxide − Mass of iron

$= (2.4 - 1.68) = 0.72\text{ g}$

Ratio of oxygen and iron = 0.72 : 1.68 = 1 : 2.33

In the second experiment

The mass of iron oxide = 2.9 g

The mass of iron after reduction = 2.03 g

The mass of oxygen = (2.9 − 2.03) = 0.87 g

Ratio of oxygen and iron = 0.87 : 2.03 = 1 : 2.33

Thus, the data illustrate the law of constant proportions, as in both the experiments the ratio of oxygen and iron is the same.

Example 5. *Carbon combines with hydrogen to form three compounds A, B and C. The percentages of hydrogen in A, B and C are 25, 14.3 and 7.7 respectively. Which law of chemical combination is illustrated?*

Solution:

Compound	% of Hydrogen	% of Carbon
A	25.0	(100 − 25.0) = 75.0
B	14.3	(100 − 14.3) = 85.7
C	7.7	(100 − 7.7) = 92.3

In Compound *A*

25 parts of hydrogen combine with 75 parts of carbon

1 part of hydrogen combines with 75/25

= 3 parts of carbon

In Compound *B*

14.3 parts of hydrogen combine with 85.7 parts of carbon

1 part of hydrogen combines with 85.7/14.3

= 6.0 parts of carbon

In Compound *C*

7.7 parts of hydrogen combine with 92.3 parts of carbon

1 part of hydrogen combines with 92.3/7.7

= 12.0 parts of carbon

Thus, the masses of carbon in three compounds A, B and C, which combine with a fixed mass of hydrogen are in the ratio of 3 : 6 : 12 or 1 : 2 : 4. This is a simple ratio. Hence, the data illustrate the law of multiple proportions.

Example 6. *Two compounds each containing only tin and oxygen had the following composition:*

	Mass % of tin	*Mass % of oxygen*
Compound A	78.77	21.23
Compound B	88.12	11.88

Show how this data illustrate the law of multiple proportions?

Solution:

In Compound A

21.23 parts of oxygen combine with 78.77 parts of tin

1 part of oxygen combines with 78.77/21.23

= 3.7 parts of tin

In Compound B

11.88 parts of oxygen combine with 88.12 parts of tin

1 part of oxygen combines with 88.12/11.88

= 7.4 parts of tin

Thus, the mass of tin in compounds A and B which combine with a fixed mass of oxygen are in the ratio of 3.7 : 7.4 or 1 : 2. This is a simple ratio. Hence, the data illustrate the law of multiple proportions.

Example 7. *Illustrate the law of reciprocal proportions from the following data: KCl contains 52.0% potassium, KI contains 23.6% potassium and ICl contains 78.2% iodine.*

Solution: **In KCl:** Potassium 52.0%,

Chlorine $(100-52)=48\%$

In KI: Potassium 23.6%;

Iodine $(100-23.6)=76.4\%$

23.6 parts of potassium combine with 76.4 parts of iodine

52.0 parts of potassium will combine with

$(76.4/23.6)\times 52.0 = 168.3$ parts of iodine.

The ratio of masses of chlorine and iodine which combines with same mass of potassium $= 48:168.3$ or $1:3.5$

In ICl: Iodine = 78.2% and chlorine

$=(100-78.2)=21.8\%$

The ratio of chlorine and iodine in ICl $=21.8:78.2=1:3.5$. Hence, the data illustrate the law of reciprocal proportions.

Example 8. *Zinc sulphate crystals contain 22.6% of zinc and 43.9% of water. Assuming the law of constant proportions to be true, how much zinc should be used to produce 13.7 g of zinc sulphate and how much water will they contain?*

Solution: 100 g of zinc sulphate crystals are obtained from

= 22.6 g zinc

1 g of zinc sulphate crystals will be obtained from

= 22.6/100 g zinc

13.7 g of zinc sulphate crystals will be obtained from

$$=\frac{22.6}{100}\times 13.7$$

= 3.0962 g of zinc

100 g of zinc sulphate crystals contain water

= 43.9 g

1 g of zinc sulphate crystals contain water

= 43.9/100 g

13.7 g of zinc sulphate crystals shall contain water

$$=\frac{43.9}{100}\times 13.7 = 6.0143\ \text{g}$$

Example 9. *Carbon monoxide reacts with oxygen to form carbon dioxide according to the equation, $2CO + O_2 = 2CO_2$. In an experiment, 400 mL of carbon monoxide and 180 mL of oxygen were allowed to react, when 80% of carbon monoxide was transformed to carbon dioxide.*

All the volumes were measured under the same conditions of temperature and pressure. Find out the composition of the final mixture.

Solution:
$$\underset{2\ \text{vol}}{2CO} + \underset{1\ \text{vol}}{O_2} = \underset{2\ \text{vol}}{2CO_2}$$

From the above equation, it is observed that volume of oxygen required for the transformation of carbon monoxide into carbon dioxide is half the volume of carbon monoxide and the volume of carbon dioxide produced is same as that of carbon monoxide.

Volume of carbon monoxide transformed

$$=\frac{80\times 400}{100}=320\ \text{mL}$$

Hence, volume of oxygen required for transformation

$$=\frac{1}{2}\times 320 = 160\ \text{mL}$$

Volume of carbon dioxide produced

= 320 mL

So, the composition of final mixture is

Carbon monoxide = (400 − 320)

= 80 mL

Carbon dioxide = 320 mL

Oxygen = 180 − 160 = 20 mL

Example 10. *How much volume of oxygen will be required for complete combustion of 40 mL of acetylene (C_2H_2) and how much volume of carbon dioxide will be formed? All volumes are measured at NTP.*

Solution:
$$\begin{array}{cccccc} 2C_2H_2 & + & 5O_2 & = & 4CO_2 & +2H_2O \\ 2\ \text{vol} & & 5\ \text{vol} & & 4\ \text{vol} & \\ 40\ \text{mL} & & \frac{5}{2}\times 40\ \text{mL} & & \frac{4}{2}\times 40\ \text{mL} & \\ 40\ \text{mL} & & 100\ \text{mL} & & 80\ \text{mL} & \end{array}$$

So, for complete combustion of 40 mL of acetylene, 100 mL of oxygen are required and 80 mL of carbon dioxide is formed.

1.9 DALTON'S ATOMIC THEORY

The concept that matter is composed of very small particles was given by Indian and Greek philosophers. As early as 400 to 500 B.C. the Greek philosopher **Democritus** suggested that matter cannot be forever divided into smaller and smaller parts. The ultimate particles were considered as indivisible. These particles were called atoms. The word atom has been derived from the Greek word '*atomos*' meaning 'indivisible'. These early ideas, however, were not based on experiments but were mere speculations. The existence of atoms was accepted by Boyle in his book '*The Sceptical Chymist*' (1661) and by **Newton** in his books '*Principia*' and '*Opticks*' (1704). The old ideas were put on a scientific scale by **John Dalton** in the years 1803 to 1808 in the form of a theory known as **Dalton's Atomic Theory** which is a landmark in the history of chemistry. The main points of Dalton's atomic theory are:

(i) Elements consist of minute, indivisible, indestructible particles called atoms.
(ii) Atoms of an element are identical to each other. They have the same mass and size.
(iii) Atoms of different elements differ in properties and have different masses and sizes.
(iv) Compounds are formed when atoms of different elements combine with each other in simple numerical ratios such as one-to-one, one-to-two, two-to-three and so on.
(v) Atoms cannot be created, destroyed or transformed into atoms of other elements.
(vi) The relative numbers and kind of atoms are always the same in a given compound.

The theory convincingly explained the various laws of chemical combination, but the theory has undergone a complete shake up with the modern concept of structure of atom. However, the Daltonian atom still retains its significance as the unit participating in chemical reactions. The following are the modified views regarding Dalton's atomic theory:

(i) The atom is no longer supposed to be indivisible. The atom is not a simple particle but a complex one.
(ii) Atoms of the element may not necessarily possess the same mass but possess the same atomic number and show similar chemical properties **(Discovery of isotopes).**
(iii) Atoms of the different elements may possess the same mass but they always have different atomic numbers and differ in chemical properties **(Discovery of isobars).**
(iv) Atoms of one element can be transmuted into atoms of other element. **(Discovery of artificial transmutation).**
(v) In certain organic compounds, like proteins, starch, cellulose, etc., the ratio in which atoms of different elements combine cannot be regarded as simple. There are a number of compounds which do not follow the law of constant proportions. Such compounds are called **non-stoichiometric compounds.**

1.10 ATOMS, MOLECULES AND FORMULAE

An atom is the smallest particle of an element. The atom of hydrogen is the smallest and the lightest. Atoms take part in chemical combination and remain as indivisible. All atoms do not occur free in nature. **Avogadro** introduced the idea of another kind of particles called the molecules. **A molecule is the smallest particle of an element or compound that can have a stable and independent existence.** A molecule of an element consists of one or more atoms of the same element. Certain elements are capable of existence as single atoms and their atoms can be regarded as molecules. A molecule of an element that consists of one atom only is called monoatomic molecule as in the case of inert gases. Oxygen is not stable in atomic form but is stable in molecular form. A molecule of oxygen is diatomic in nature, *i.e.*, its molecule consists of two oxygen atoms. Hydrogen, nitrogen, fluorine, chlorine, bromine, iodine are also diatomic like oxygen. Some elements exist in more complex molecular forms. The molecule of phosphorus consists of four phosphorus atoms and the molecule of sulphur consists of eight sulphur atoms. Such molecules having more than two atoms are said to be polyatomic. A representation of the molecule of an element involves use of a subscript to the right of the elemental symbol. The diatomic molecule of chlorine is represented as Cl_2, whereas molecules of phosphorus and sulphur are represented as P_4 and S_8, respectively.

The molecule is the smallest possible unit of a compound which shows the properties of the compound. The molecules of all compounds contain two or more different types of atoms. These differ from the molecules of elements which contain only one type of atoms.

Thus, it becomes clear that atoms are the components of molecules and the molecules are components of elements or compounds.

The formula is a group of symbols of elements which represents one molecule of a substance. The formula of a substance represents its chemical composition. Water consists of molecules containing two hydrogen atoms and one oxygen atom which are represented as H_2O. The subscript to the right of the symbol for hydrogen indicates the number of hydrogen atoms contained in a molecule. No subscript follows the symbol for oxygen which means, by convention, that only one atom of oxygen is contained in the molecule.

The subscripts representing the number of atoms contained in a molecule of a compound are in no way related to the number of atoms present in the molecule of a free element. Although both hydrogen and oxygen are composed of diatomic molecules, a water molecule contains only one atom of oxygen and two atoms of hydrogen. The two hydrogen atoms present in H_2O are not molecular hydrogen but rather two hydrogen atoms that have chemically combined with an oxygen atom.

For a chemical formula to be correct, it must contain two pieces of information: (i) it must indicate the elements in the make up of the compound, and (ii) it must indicate the combining ratio of atoms of these elements in the particular compound. The first information is provided by including in the formula correct chemical symbols for all the elements in the compound. The second piece of information is provided by subscripts, *i.e.*, numbers written to the right slightly below the chemical symbols of the elements.

Nitric acid is a combination of hydrogen, nitrogen and oxygen giving a base formula HNO. These elements combine in the ratio 1 : 1 : 3. Therefore, the correct formula for nitric acid is HNO_3.

Some compounds are composed of ions rather than of molecules. Ions differ from atoms and molecules by being electrically charged particles of matter. The charges may be positive or negative and generally vary in magnitude. The positively charged ions are called cations and negatively charged ions are called anions. Simple cations and anions come into existence by loss and acceptance of an electron or electrons by neutral atoms respectively. Ions that consist of several atoms held together by chemical bonds similar to those involved in the molecules are called polyatomic ions or complex ions. These complex ions differ from molecules in the sense that they bear a charge. Some of the common complex ions are:

NO_3^-	Nitrate	PO_4^{3-}	Phosphate	NH_4^+	Ammonium
SO_4^{2-}	Sulphate	ClO_4^-	Perchlorate	PH_4^+	Phosphonium
SO_3^{2-}	Sulphite	CO_3^{2-}	Carbonate	MnO_4^-	Permanganate

When ions are present in a compound, the number of positive charges on a cation must balance with the negative charges on an anion to produce electrically neutral matter. Since, the charge on the anion may not always be equal to that on the cation, the number of anions will not always be equal to the number of cations.

Calcium nitrate consists of calcium and nitrate ions. Each calcium ion carries 2 units positive charge while each nitrate ion carries 1 unit negative charge. Thus, to make net charge zero, two nitrate ions will link with one calcium ion and the formula will be $Ca(NO_3)_2, [Ca^{2+} + 2NO_3^-]$. Names and formulae of some common chemical compounds are listed below :

Common Name	Chemical Name	Chemical Formula
Alum	Ammonium aluminium sulphate	$(NH_4)_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$
Aspirin	Acetyl salicylic acid	$C_9H_8O_4$
Battery acid or oil of vitriol	Sulphuric acid	H_2SO_4
Blue vitriol	Copper sulphate	$CuSO_4 \cdot 5H_2O$
Baking soda	Sodium bicarbonate	$NaHCO_3$
Bleaching powder	Calcium chlorohypochlorite	$CaOCl_2$
Borax	Sodium tetraborate	$Na_2B_4O_7 \cdot 10H_2O$
Butter of tin	Stannic chloride	$SnCl_4 \cdot 5H_2O$
Caustic soda	Sodium hydroxide	NaOH
Caustic potash	Potassium hydroxide	KOH
Carbolic acid	Phenol	C_6H_5OH
Chile saltpetre	Sodium nitrate	$NaNO_3$
Carborundum	Silicon carbide	SiC
Corrosive sublimate	Mercuric chloride	$HgCl_2$
Calomel	Mercurous chloride	Hg_2Cl_2
Dry ice	Carbon dioxide (solid)	CO_2
Formalin	Formaldehyde (40% solution)	HCHO
Grain alcohol (Spirit)	Ethyl alcohol	C_2H_5OH
Green vitriol	Ferrous sulphate	$FeSO_4 \cdot 7H_2O$
Gypsum	Calcium sulphate	$CaSO_4 \cdot 2H_2O$
Gammexane (BHC)	Benzene hexachloride	$C_6H_6Cl_6$
Hydrolith	Calcium hydride	CaH_2
Hypo (Antichlor)	Sodium thiosulphate	$Na_2S_2O_3 \cdot 5H_2O$
Indian nitre	Potassium nitrate	KNO_3
Limestone	Calcium carbonate	$CaCO_3$
Lunar caustic	Silver nitrate	$AgNO_3$
Laughing gas	Nitrous oxide	N_2O
Litharge	Lead monoxide	PbO
Muratic acid	Hydrochloric acid	HCl
Mohr's salt	Ferrous ammonium sulphate	$FeSO_4(NH_4)_2SO_4 \cdot 6H_2O$
Milk of magnesia	Magnesium hydroxide	$Mg(OH)_2$
Microcosmic salt	Sodium ammonium hydrogen ortho-phosphate	$Na(NH_4)HPO_4$
Marsh gas (Damp fire)	Methane	CH_4
Oleum	Sulphuric acid (Fuming)	$H_2S_2O_7$
Oxone	Sodium peroxide	Na_2O_2
Plaster of Paris	Calcium sulphate hemihydrate	$CaSO_4 \cdot \frac{1}{2}H_2O$
Philosphers's wool	Zinc oxide	ZnO
Phosgene	Carbonyl chloride	$COCl_2$
Pearl ash	Potassium carbonate	K_2CO_3
Pyrene	Carbon tetrachloride	CCl_4
Picric acid	2,4,6-Trinitrophenol	$C_6H_2(OH)(NO_2)_3$
Quick lime	Calcium oxide	CaO
Red lead (Minium)	Lead tetroxide	Pb_3O_4
Sugar	Sucrose	$C_{12}H_{22}O_{11}$
Slaked lime (Milk of lime)	Calcium hydroxide	$Ca(OH)_2$
Sal ammoniac	Ammonium chloride	NH_4Cl
Sugar of lead	Lead acetate	$(CH_3COO)_2Pb$
Sand	Silicon dioxide	SiO_2
Table salt (Common salt)	Sodium chloride	NaCl
TEL	Tetra-ethyl lead	$Pb(C_2H_5)_4$
Tear gas	Chloropicrin	CCl_3NO_2
Washing soda	Sodium carbonate	$Na_2CO_3 \cdot 10H_2O$
Water glass	Sodium silicate	Na_2SiO_3
White vitriol	Zinc sulphate	$ZnSO_4 \cdot 7H_2O$

1.11 ATOMIC AND MOLECULAR MASS

One of the most important concepts derived from Dalton's atomic theory is that of atomic mass, *i.e.*, each element has a characteristic atomic mass. As atoms are very tiny particles, their absolute masses are difficult to measure. However, it is possible to determine the relative masses of different atoms if a small unit of mass is taken as a standard. For this purpose, mass of one atom of hydrogen was assumed as unity and was accepted as standard. **The atomic mass of an element can be defined as the number which indicates how many times the mass of one atom of the element is heavier in comparison to the mass of one atom of hydrogen.**

$$A = \text{Atomic mass of an element}$$

$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of hydrogen}}$$

In 1858, oxygen atom was adopted as a standard on account of the following reasons:

(i) It is much easier to obtain compounds of elements with oxygen than with hydrogen as oxygen is more reactive than hydrogen.

(ii) The atomic masses of most of the elements become approximately whole numbers but with hydrogen as standard the atomic masses of most of the elements are fractional.

The mass of one atom of natural oxygen was taken to be 16.0.

Thus, atomic mass of an element

$$= \frac{\text{Mass of one atom of the element}}{\frac{1}{16}\text{th part of the mass of one atom of oxygen}}$$

$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of oxygen}} \times 16$$

By accepting oxygen as a standard, the atomic mass of hydrogen comes as 1.008, sodium 22.991 and sulphur 32.066.

In 1961, the International Union of Chemists selected a new unit for expressing the atomic masses. They accepted the stable isotope of carbon (^{12}C) with mass number of 12 as the standard. **Atomic mass of an element can be defined as the number which indicates how many times the mass of one atom of the element is heavier in comparison to $\frac{1}{12}$th part of the mass of one atom of carbon-12 (^{12}C).**

$$A = \text{Atomic mass of an element}$$

$$= \frac{\text{Mass of one atom of the element}}{\frac{1}{12}\text{th part of the mass of one atom of carbon-12}}$$

$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of carbon-12}} \times 12$$

[The quantity 'A' was formerly known as atomic weight. However, this term is no longer used as the word 'weight' means gravitational force.]

Atomic mass unit: The quantity $\frac{1}{12}$ mass of an atom of carbon-12 (^{12}C) is known as the atomic mass unit and is abbreviated as amu. The actual mass of one atom of carbon-12 is 1.9924×10^{-23} g or 1.9924×10^{-26} kg.

*The term Dalton is used for one atomic mass unit, 1 Dalton = 1 amu.

Thus,

$$1\,\text{amu}^* = \frac{1.9924 \times 10^{-23}}{12} = 1.66 \times 10^{-24}\text{ g or } 1.66 \times 10^{-27}\text{ kg}$$

$$A = \text{Atomic mass of an element}$$

$$= \frac{\text{Mass of one atom of the element}}{1\,\text{amu}}$$

The atomic masses of some elements on the basis of carbon-12 are given below:

Hydrogen	1.008 amu	Iron	55.847 amu
Oxygen	16.00 amu	Sodium	22.989 amu
Chlorine	35.453 amu	Zinc	65.38 amu
Magnesium	24.305 amu	Silver	107.868 amu
Copper	63.546 amu		

The actual mass of an atom of an element

= The atomic mass of an element in amu $\times 1.66 \times 10^{-24}$ g

So, the actual mass of hydrogen atom

$$= 1.008 \times 1.66 \times 10^{-24} = 1.6736 \times 10^{-24}\text{ g}$$

Similarly, the actual mass of oxygen atom

$$= 16 \times 1.66 \times 10^{-24} = 2.656 \times 10^{-23}\text{ g}$$

It is clear from the above list of atomic masses that atomic masses of a number of elements are not nearly whole numbers. Actually, the above values are average relative masses. Most of the elements occur in nature as a mixture of isotopes. (Isotopes—the atoms of the same element having different atomic masses). With very few exceptions, however, elements have constant mixtures of isotopes. Chlorine is found in nature as a mixture containing two isotopes Cl-35 (34.969 amu) and Cl-37 (36.966 amu). These are found in the ratio of 75.53% (Cl-35) and 24.47% (Cl-37). Therefore, the average relative mass of chlorine is calculated as:

$$(34.969 \times 0.7553) + (36.966 \times 0.2447) = 35.46\text{ amu}$$

Based on the average mass, the atomic mass of chlorine is 35.46 or 35.5 amu but it is never possible to have an atom having a relative mass 35.5 amu. It can have relative mass of about 35.0 or 37.0 amu depending on the particular isotope. Thus, average relative mass of any naturally occurring sample of chlorine is 35.46 or 35.5 amu as it is a mixture of two isotopes present in definite proportion. The same reasoning applies to all other elements.

The average atomic masses of various elements are determined by multiplying the atomic mass of each isotope by its fractional abundance and adding the values thus obtained. The fractional abundance is determined by dividing percentage abundance by hundred.

$$\text{Average isotopic mass} = \frac{m \times a + n \times b}{m + n}$$

here, a, b are atomic masses of isotopes in the ratio $m : n$.

$$\text{Average isotopic mass} = \frac{x}{100} \times a + \frac{y}{100} \times b$$

here, x, y are percentage abundance of the two isotopes $(y = 100 - x)$.

Example 11. *Boron has two isotopes boron-10 and boron-11 whose percentage abundances are 19.6% and 80.4% respectively. What is the average atomic mass of boron?*

Solution:

$$\text{Contribution of boron-10} = 10.0 \times 0.196 = 1.96 \text{ amu}$$

$$\text{Contribution of boron-11} = 11.0 \times 0.804 = 8.844 \text{ amu}$$

$$\text{Adding both} = 1.96 + 8.844 = 10.804 \text{ amu}$$

Thus, the average atomic mass of boron is 10.804 amu.

Example 12. *Carbon occurs in nature as a mixture of carbon-12 and carbon-13. The average atomic mass of carbon is 12.011. What is the percentage abundance of carbon-12 in nature?*

Solution: Let x be the percentage abundance of carbon-12; then $(100 - x)$ will be the percentage abundance of carbon-13.

Therefore, $$\frac{12x}{100} + \frac{13(100 - x)}{100} = 12.011$$

or $$12x + 1300 - 13x = 1201.1$$

or $$x = 98.9$$

Abundance of carbon-12 is 98.9%.

Gram-atomic Mass or Gram Atom

When numerical value of atomic mass of an element is expressed in grams, the value becomes gram-atomic mass or gram atom. The atomic mass of oxygen is 16 while gram-atomic mass or gram atom of oxygen is 16 g. Similarly, the gram-atomic masses of hydrogen, chlorine and nitrogen are 1.008 g, 35.5 g and 14.0 g respectively. **Gram-atomic mass or gram atom of every element consists of same number of atoms.** This number is called **Avogadro's number.** The value of Avogadro's number is 6.02×10^{23}.

Absolute mass of one oxygen atom

$$= 16 \text{ amu} = 16 \times 1.66 \times 10^{-24} \text{ g}$$

Therefore, the mass of 6.02×10^{23} atoms of oxygen will be

$$= 16 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23}$$

$$= 16 \text{ g (gram-atomic mass)}$$

Thus, gram-atomic mass can be defined as the absolute mass in grams of 6.02×10^{23} atoms of any element.

Number of gram atoms of any element can be calculated with the help of the following formula:

$$\text{No. of gram atoms} = \frac{\text{Mass of the element in grams}}{\text{Atomic mass of the element in grams}}$$

Molecular Mass

Like an atom, a molecule of a substance is also a very small particle possessing a mass of the order of 10^{-24} to 10^{-22} g. Similar to atomic mass, molecular mass is also expressed as a relative mass with respect to the mass of the standard substance which is an atom of hydrogen or an atom of oxygen or an atom of carbon-12. The molecular mass of a substance may be defined as the mass of a molecule of a substance relative to the mass of an atom of hydrogen as 1.008 or of oxygen taken as 16.00 or the mass of one atom of carbon taken as 12. **Molecular mass is a number which indicates how many times one molecule of a substance is heavier in comparison to $\frac{1}{16}$th of the mass of oxygen atom or $\frac{1}{12}$th of the mass of one atom of carbon-12.**

$$M = \text{Molecular mass}$$

$$= \frac{\text{Mass of one molecule of the substance}}{\frac{1}{12}\text{th mass of one atom of carbon-12}}$$

The mass of a molecule is equal to sum of the masses of the atoms present in a molecule. One molecule of water consists of 2 atoms of hydrogen and one atom of oxygen. Thus, molecular mass of water $= (2 \times 1.008) + 16.00 = 18.016$ amu. One molecule of H_2SO_4 (sulphuric acid) consists of 2 atoms of hydrogen, one atom of sulphur and four atoms of oxygen. Thus, the molecular mass of sulphuric acid is

$$= (2 \times 1.008) + 32.00 + (4 \times 16.00)$$

$$= 98.016 \text{ or } 98.016 \text{ amu}$$

Gram-molecular Mass or Gram Molecule

A quantity of substance whose mass in grams is numerically equal to its molecular mass is called gram-molecular mass. In other words, molecular mass of a substance expressed in grams is called gram-molecular mass or gram molecule. For example, the molecular mass of chlorine is 71 and, therefore, its gram-molecular mass or gram molecule is 71 g.

Similarly, molecular mass of oxygen (O_2) is 32, *i.e.*, $2 \times 16 = 32$ amu.

$$\text{Gram-molecular mass of oxygen} = 32 \text{ g}$$

Molecular mass of nitric acid (HNO_3) is 63, *i.e.*,

$$= 1 + 14 + 3 \times 16 = 63 \text{ amu}$$

$$\text{Gram-molecular mass of nitric acid} = 63 \text{ g}$$

Gram-molecular mass should not be confused with the mass of one molecule of the substance in grams. The mass of one molecule of a substance is known as its **actual mass.** For example, the actual mass of one molecule of oxygen is equal to $32 \times 1.66 \times 10^{-24}$ g, *i.e.*, 5.32×10^{-23} g.

The number of gram molecules of a substance present in a given mass of a substance can be determined by the application of following formula:

No. of gram molecules

$$= \frac{\text{Mass of a substance in grams}}{\text{Molecular mass of the substance in grams}}$$

$$\text{Mass of single molecule} = \frac{\text{Molar mass in grams}}{6.023 \times 10^{23}}$$

$$= \text{Molar mass in amu} \times 1.66 \times 10^{-24} \text{ grams}$$

Example 13. *Calculate the mass of 2.5 gram atoms of oxygen.*

Solution: We know that,

$$\text{No. of gram atoms} = \frac{\text{Mass of the element in grams}}{\text{Atomic mass of the element in grams}}$$

So, Mass of oxygen $= 2.5 \times 32 = 80.0\,g$

Example 14. *Calculate the gram atoms in 2.3 g of sodium.*

Solution: No. of gram atoms $\frac{2.3}{23} = 0.1$

[Atomic mass of sodium = 23 g]

Example 15. *Calculate the mass of 1.5 gram molecule of sulphuric acid.*

Solution: Molecular mass of

$$H_2SO_4 = 2 \times 1 + 32 + 4 \times 16 = 98.0\,\text{amu}$$

Gram-molecular mass of $H_2SO_4 = 98.0\,g$

Mass of 1.5 gram molecule of $H_2SO_4 = 98.0 \times 1.5 = 147.0\,g$

Example 16. *Calculate the actual mass of one molecule of carbon dioxide (CO_2).*

Solution: Molecular mass of CO_2 = 44 amu

$$1\,\text{amu} = 1.66 \times 10^{-24}\,g$$

So, The actual mass of $CO_2 = 44 \times 1.66 \times 10^{-24}$

$$= 7.304 \times 10^{-23}\,g$$

1.12 AVOGADRO'S HYPOTHESIS

According to Dalton's atomic theory, elements react with each other in the simple ratio of their atoms. Gay-Lussac proposed that gases combine in simple ratio of their volumes. In an attempt to correlate Dalton's atomic theory with Gay-Lussac law of gaseous volumes, **Berzelius** stated that **under similar conditions of temperature and pressure, equal volume of all gases contain the same number of atoms.** This hypothesis was subsequently found to be incorrect as it failed to interpret the experimental results and contradicted the very basic assumption of Dalton's atomic theory, *i.e.*, an atom is indivisible. For example, the formation of hydrogen chloride from hydrogen and chlorine could not be explained on the basis of Berzelius hypothesis.

Hydrogen +	Chlorine =	Hydrogen chloride
1 vol	1 vol	2 vol
n atoms	n atoms	$2n$ compound atoms
1 atom	1 atom	2 compound atoms
$\frac{1}{2}$ atom	$\frac{1}{2}$ atom	1 compound atom

i.e., for the formation of 1 compound atom of hydrogen chloride, $\frac{1}{2}$ atom of hydrogen and $\frac{1}{2}$ atom of chlorine are needed. In other words, each atom of hydrogen and chlorine has been divided which is against Dalton's atomic theory. Thus, the hypothesis of Berzelius was discarded.

The Italian scientist, **Amedeo Avogadro,** in 1811, solved the above problem by proposing two types of particles from which whole of the matter is composed of.

(i) Atom: The smallest particle of an element that can take part in chemical change but generally cannot exist freely as such.

(ii) Molecule: The smallest particle of a substance (element or compound) which has free or independent existence and possesses all characteristic properties of the substance. A molecule of an element is composed of like atoms while a molecule of a compound contains fixed number of atoms of two or more different elements. A molecule may be broken down into its constituent atoms but the atom is indivisible during a chemical change.

Avogadro after making the above differentiation, presented a hypothesis known as **Avogadro hypothesis** which can be stated as follows:

"Under similar conditions of temperature and pressure, equal volumes of all gases contain equal number of molecules."

Avogadro hypothesis explains successfully the formation of hydrogen chloride.

Hydrogen +	Chlorine =	Hydrogen chloride
1 vol	1 vol	2 vol
n molecules	n molecules	$2n$ molecules
1 molecule	1 molecule	2 molecules
$\frac{1}{2}$ molecule	$\frac{1}{2}$ molecule	1 molecule
1 atom	1 atom	1 molecule

(Both hydrogen and chlorine are diatomic in nature.)

Thus, the hypothesis explains that the molecules of reacting gases break up into constituent atoms during chemical change which then combine to form new molecules of the product or products.

Applications of Avogadro's hypothesis

(i) Atomicity*: Atomicity means number of atoms present in one molecule of an elementary gas. Hydrogen, oxygen, nitrogen, chlorine, etc., are diatomic in nature. Noble gases are monoatomic while ozone is triatomic in nature. Avogadro's hypothesis helps in determining the atomicity of elements.

(ii) Relationship between molecular mass and vapour density: The vapour density of any gas is the ratio of the densities of the gas and hydrogen under similar conditions of temperature and pressure.

$$\text{Vapour Density (V.D.)} = \frac{\text{Density of gas}}{\text{Density of hydrogen}}$$

$$= \frac{\text{Mass of a certain volume of the gas}}{\text{Mass of same volume of hydrogen at the same temp. and pressure}}$$

If n molecules are present in the given volume of a gas and hydrogen under similar conditions of temperature and pressure,

$$\text{V.D.} = \frac{\text{Mass of } n \text{ molecules of gas}}{\text{Mass of } n \text{ molecules of hydrogen}}$$

$$= \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 1 molecule of hydrogen}}$$

*Atomicity can be ascertained with the values of ratio of two specific heats of gases $\left(\frac{C_P}{C_V}\right)$

$\frac{C_P}{C_V} = 1.66$ (Monoatomic), $\frac{C_P}{C_V} = 1.40$ (Diatomic), $\frac{C_P}{C_V} = 1.33$ (Polyatomic)

$$= \frac{\text{Molecular mass of gas}}{\text{Molecular mass of hydrogen}}$$
$$= \frac{\text{Mol. mass}}{2}$$
(since, mol. mass of hydrogen $= 2$)

Hence, $2 \times$ V. D. = Mol. mass

This formula can be used for the determination of molecular masses of volatile substances from vapour density. Vapour density is measured mainly by two methods:

(a) Victor Meyer and (b) Duma's methods.

(iii) Gram-molecular volume: 1 g mole of any gas occupies 22.4 litres or 22400 mL of volume at NTP or STP conditions.*

The density of hydrogen at NTP is 0.00009 g mL^{-1}. Thus, 0.00009 g of hydrogen will occupy volume at NTP = 1 mL

$$1 \text{ g of hydrogen occupies volume at NTP} = \frac{1}{0.00009} \text{ mL}$$

1 g mole of hydrogen (2.016 g) occupies volume at NTP

$$= \frac{2.016}{0.00009} = 22400 \text{ mL} = 22.4 \text{ litre}$$

According to Avogadro's hypothesis, equal volumes of different gases contain same number of molecules under similar conditions of temperature and pressure. Thus, 22.4 litre or 22400 mL of any gas at NTP will contain one gram mole or its molecular mass in grams.

Loschmidt number : Number of molecules in 1 cm^3 or 1 mL of a gas at S.T.P. is known as Loschmidt number.

$$\text{Loschmidt number} = \frac{6.023 \times 10^{23}}{22400}$$
$$= 2.68 \times 10^{18} \text{ molecules mL}^{-1}$$

(iv) Molecular formula: Avogadro's hypothesis helps in finding the molecular formulae of gases. Under similar conditions of temperature and pressure, 2 volumes of ozone after decomposition give 3 volumes of oxygen.

$$\underset{2\text{ vol}}{\text{Ozone}} \xrightarrow{\text{Decomposition}} \underset{3\text{ vol}}{\text{Oxygen}}$$

2 molecules	3 molecules
1 molecule	3/2 molecules
1 molecule	3 atoms

Thus, the formula of ozone is O_3.

1.13 MOLE CONCEPT

For the counting of articles, the unit **dozen** or unit **gross** is commonly used irrespective of their nature. For example, one dozen pencils means 12 pencils or one dozen apples means 12 apples or one gross books means 144 books or one gross oranges means 144 oranges. In a similar way, for counting of atoms, molecules, ions, etc., chemists use the unit mole. The term mole was introduced by **Ostwald** in 1896. This is the Latin word 'moles' meaning heap or pile. **A mole (mol) is defined as the number of atoms in 12.00 g of carbon-12.** The number of atoms in 12 g of carbon-12 has been found experimentally to be 6.02×10^{23}. This number is also known as Avogadro's number named in honour of **Amedeo Avogadro** (1776 – 1856).

Thus, a mole contains 6.02×10^{23} units. These units can be atoms, molecules, ions, electrons or anything else.

1 mole of hydrogen atoms means 6.02×10^{23} hydrogen atoms.

1 mole of hydrogen molecules means 6.02×10^{23} hydrogen molecules.

1 mole of potassium ions means 6.02×10^{23} potassium ions.

1 mole of electrons means 6.02×10^{23} electrons.

The type of entity must be specified when the mole designation is used. A mole of oxygen atoms contains 6.02×10^{23} oxygen atoms and a mole of oxygen molecules contains 6.02×10^{23} oxygen molecules. Therefore, a mole of oxygen molecules is equal to two moles of oxygen atoms, *i.e.*, $2 \times 6.02 \times 10^{23}$ oxygen atoms.

How much does one mole weigh? That depends on the nature of particles (units). **The mass of one mole atoms of any element is exactly equal to the atomic mass in grams (gram-atomic mass or gram atom) of that element.**

For example, the atomic mass of aluminium is 27 amu. One amu is equal to 1.66×10^{-24} g. One mole of aluminium contains 6.02×10^{23} aluminium atoms.

$$\text{Mass of one atom aluminium} = 27 \times 1.66 \times 10^{-24} \text{ g}$$
$$\text{Mass of one mole aluminium} = 27 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23}$$
$$= 27 \text{ g}$$

This is the atomic mass of aluminium in grams or it is one gram atomic mass or one gram atom of aluminium.

Similarly, **the mass of 6.02×10^{23} molecules (1 mole) of a substance is equal to its molecular mass in grams or gram-molecular mass or gram molecule.** For example, molecular mass of water is 18 amu. Thus, mass of one mole of water will be $18 \times 1.66 \times 10^{-24} \times 6.02 \times 10^{23}$, *i.e.*, 18 g. This is the molecular mass of water in grams or one gram-molecular mass or one gram molecule.

Mole concept is also applicable to ionic compounds which do not contain molecules. In such cases, the formula of an ionic compound represents the ratio between constituent ions. The mass of 6.02×10^{23} formula units represents one mole of an ionic compound.

* 0°C or 273 K temperature and one atmosphere or 760 mm of Hg or 76 cm of Hg pressure are known as the standard conditions of temperature and pressure (STP) or normal conditions of temperature and pressure (NTP).

One mole of $BaCl_2 = 6.02 \times 10^{23}$ $BaCl_2$ units
$= 208.2$ g $BaCl_2$
= Molecular mass (formula mass) of $BaCl_2$
$= 6.02 \times 10^{23}$ Ba^{2+} ions $+ 2 \times 6.02 \times 10^{23}$ Cl^- ions
$= 137.2 + 71.0 = 208.2$ g

One mole of a substance will have mass equal to formula mass of that substance expressed in grams.

It has been established by Avogadro's hypothesis that one gram-molecular mass of any gaseous substance occupies a volume of 22.4 litres at NTP. One gram-molecular mass is nothing but one mole of substance. Thus, one mole, *i.e.*, 6.02×10^{23} molecules of any gaseous substance occupies 22.4 litres as volume at NTP.

The following formulae satisfy the above discussion.

1 mole of a substance $= 6.02 \times 10^{23}$ particles of the substance

Number of moles of a substance

$$= \frac{\text{Mass of substance in gram}}{\text{Mass of one mole of the substance in gram}}$$

Further, $\text{Number of moles} = \dfrac{\text{No. of particles}}{6.02 \times 10^{23}}$

Thus,

$$\frac{\text{No. of particles}}{6.02 \times 10^{23}} = \frac{\text{Mass of substance in gram}}{\text{Mass of one mole of the substance in gram}}$$

Mass of one atom of an element

$$= \frac{\text{Gram atom of an element}}{6.02 \times 10^{23}}$$

Mass of one molecule of a substance

$$= \frac{\text{Gram-molecular mass of the substance}}{6.02 \times 10^{23}}$$

Number of molecules

$$= \frac{\text{Volume of gas in litres at NTP}}{22.4} \times 6.02 \times 10^{23}$$

SOME SOLVED EXAMPLES

Example 17. *A piece of copper weighs* 0.635 g. *How many atoms of copper does it contain?* **[CEE (Bihar) 1992]**

Solution: Gram-atomic mass of copper = 63.5 g

$$\text{Number of moles in 0.635 g of copper} = \frac{0.635}{63.5} = 0.01$$

Number of copper atoms in one mole $= 6.02 \times 10^{23}$

Number of copper atoms in 0.01 moles $= 0.01 \times 6.02 \times 10^{23}$
$= 6.02 \times 10^{21}$

Example 18. *How many molecules of water and oxygen atoms are present in* 0.9 *g of water?*

Solution: Gram-molecular mass of water = 18 g

$$\text{Number of moles in 0.9 g of water} = \frac{0.9}{18} = 0.05$$

Number of water molecules in one mole of water
$= 6.02 \times 10^{23}$

Number of molecules of water in 0.05 moles
$= 0.05 \times 6.02 \times 10^{23}$
$= 3.010 \times 10^{22}$

As one molecule of water contains one oxygen atom,

So, number of oxygen atoms in 3.010×10^{22} molecule of water $= 3.010 \times 10^{22}$

Example 19. *Calculate the mass of a single atom of sulphur and a single molecule of carbon dioxide.*

Solution:

Gram-atomic mass of sulphur = 32 g

$$\text{Mass of one sulphur atom} = \frac{\text{Gram-atomic mass}}{6.02 \times 10^{23}}$$

$$= \frac{32}{6.02 \times 10^{23}} = 5.33 \times 10^{-23} \text{ g}$$

Formula of carbon dioxide $= CO_2$

Molecular mass of $CO_2 = 12 + 2 \times 16 = 44$

Gram-molecular mass of $CO_2 = 44$ g

$$\text{Mass of one molecule of } CO_2 = \frac{\text{Gram-molecular mass}}{6.02 \times 10^{23}}$$

$$= \frac{44}{6.02 \times 10^{23}} = 7.308 \times 10^{-23} \text{ g}$$

Example 20. *What is the mass of* 3.01×10^{22} *molecules of ammonia?*

Solution: Gram-molecular mass of ammonia = 17 g

Number of molecules in 17 g (one mole) of $NH_3 = 6.02 \times 10^{23}$

Let the mass of 3.01×10^{22} molecules of NH_3 be $= x$ g

So, $$\frac{3.01 \times 10^{22}}{6.02 \times 10^{23}} = \frac{x}{17}$$

or $$x = \frac{17 \times 3.01 \times 10^{22}}{6.02 \times 10^{23}} = 0.85 \text{ g}$$

Example 21. *From* 200 *mg of* CO_2, 10^{21} *molecules are removed. How many moles of* CO_2 *are left?*

Solution:

Gram-molecular mass of $CO_2 = 44$ g

$$\text{Mass of } 10^{21} \text{ molecules of } CO_2 = \frac{44}{6.02 \times 10^{23}} \times 10^{21} = 0.073 \text{ g}$$

Mass of CO_2 left $= (0.2 - 0.073) = 0.127$ g

$$\text{Number of moles of } CO_2 \text{ left} = \frac{0.127}{44} = 2.88 \times 10^{-3}$$

Example 22. *How many molecules and atoms of oxygen are present in 5.6 litres of oxygen (O_2) at NTP?*

Solution: We know that, 22.4 litres of oxygen at NTP contain 6.02×10^{23} molecules of oxygen.

So, 5.6 litres of oxygen at NTP contain

$$= \frac{5.6}{22.4} \times 6.02 \times 10^{23} \text{ molecules}$$

$$= 1.505 \times 10^{23} \text{ molecules}$$

1 molecule of oxygen contains = 2 atoms of oxygen

So, 1.505×10^{23} molecules of oxygen contain

$$= 2 \times 1.505 \times 10^{23} \text{ atoms}$$

$$= 3.01 \times 10^{23} \text{ atoms}$$

Example 23. *How many electrons are present in 1.6 g of methane?*

Solution: Gram-molecular mass of methane,

$$(CH_4) = 12 + 4 = 16 \text{ g}$$

Number of moles in 1.6 g of methane

$$= \frac{1.6}{16} = 0.1$$

Number of molecules of methane in 0.1 mole

$$= 0.1 \times 6.02 \times 10^{23}$$

$$= 6.02 \times 10^{22}$$

One molecule of methane has $= 6 + 4 = 10$ electrons

So, 6.02×10^{22} molecules of methane have

$$= 10 \times 6.02 \times 10^{22} \text{ electrons}$$

$$= 6.02 \times 10^{23} \text{ electrons}$$

Example 24. *The electric charge on the electron is 1.602×10^{-19} coulomb. How much charge is present on 0.1 mole of Cu^{2+} ions?*

Solution: Charge on one mole of electrons

$$= 6.02 \times 10^{23} \times 1.602 \times 10^{-19} \text{ coulomb}$$

$$= 96500 \text{ coulomb} = 1 \text{ faraday}$$

Charge on one mole of Cu^{2+} ions

$$= 2 \times 96500 \text{ coulomb} = 2 \text{ faraday}$$

Charge on 0.1 mole of Cu^{2+} ions

$$= 0.1 \times 2 = 0.2 \text{ faraday}$$

Example 25. *How many years it would take to spend one Avogadro's number of rupees at a rate of 10 lakh of rupees in one second?* **(MLNR 1990)**

Solution: Number of rupees spent in one second $= 10^6$

Number of rupees spent in one year

$$= 10^6 \times 60 \times 60 \times 24 \times 365$$

Avogadro's number of rupees will be spent

$$= \frac{6.02 \times 10^{23}}{10^6 \times 60 \times 60 \times 24 \times 365}$$

$$= 19.089 \times 10^9 \text{ years} = 1.9089 \times 10^{10} \text{ years}$$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

1. 116 mg of a compound on vaporisation in Victor Meyer's apparatus displaces 44.8 mL of air measured at STP. The molecular mass of the compound is: **[CEE (Kerala) 2004]**
(a) 116 (b) 232 (c) 58 (d) 44.8 (e) 46.4
[Ans. (c)]
[Hint: Molar mass of compound

$$= \text{Mass of 22400 mL vapour at STP}$$

$$= \frac{0.116 \times 22400}{44.8} = 58]$$

2. A gas has a vapour density 11.2. The volume occupied by 1 g of the gas at NTP is: **(JCECE 2004)**
(a) 1 L (b) 11.2 L (c) 22.4 L (d) 4 L
[Ans. (a)]
[Hint: Molar mass $= 2 \times 11.2 = 22.4$ g

Volume of 1 g compound at STP $= \frac{22.4}{22.4} = 1$ L]

3. 3 g of hydrocarbon on combustion with 11.2 g of oxygen produce 8.8 g of CO_2 and 5.4 g of H_2O. The data illustrate the law of:
(a) conservation of mass (b) multiple proportions
(c) constant proportions (d) reciprocal proportions
[Ans. (a)]
[Hint: Σ Masses of reactants $= \Sigma$ Masses of products
(3 + 11.2) g (8.8 + 5.4) g
Hence, law of conservation of mass is verified.]

4. The maximum number of molecules is present in:
[CBSE (PMT) 2004; Manipal (Medical) 2007]
(a) 15 L of H_2 gas at STP (b) 5 L of N_2 gas at STP
(c) 0.5 g of H_2 gas (d) 10 g of O_2 gas
[Ans. (a)]
[Hint:

Number of molecules in 15 L $H_2 = \frac{15}{22.4} \times N = 0.669\,N$

Number of molecules in 5 L $N_2 = \frac{5}{22.4} \times N = 0.223\,N$

Number of molecules in 0.5 g $H_2 = \frac{0.5}{2} \times N = 0.25\,N$

Number of molecules in 10 g $O_2 = \frac{10}{32} \times N = 0.312\,N$]

5. Insulin contains 3.4% sulphur. Then, the minimum molecular mass of the insulin is about:
(a) 940 amu (b) 9400 amu
(c) 3600 amu (d) 970 amu
[Ans. (a)]
[Hint: $\because$ 3.4 g sulphur is present in 100 g insulin

$\therefore$ 32 g sulphur will be present in $\frac{100}{3.4} \times 32$ g insulin = 940

$\therefore$ Molar mass of insulin is about 940 amu]

6. 25 g of MCl_4 contains 0.5 mol chlorine then its molecular mass is: **(DPMT 2007)**
(a) $100 g\ mol^{-1}$ (b) $200 g\ mol^{-1}$
(c) $150 g\ mol^{-1}$ (d) $400 g\ mol^{-1}$
[**Ans.** (b)]
[**Hint:** 1 mol of MCl_4 contains 4 mol of chlorine
$\because$ 0.5 mol chlorine is present in 25 g of MCl_4
$\therefore$ 4 mol chlorine will be present in $\frac{25}{0.5} \times 4$, *i.e.*, 200 g of MCl_4.]

1.14 EQUIVALENT MASSES OR CHEMICAL EQUIVALENTS

Equivalent mass of a substance (element or compound) is defined as **the number of parts by mass of the substance which combine or displace directly or indirectly 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of chlorine or 108 parts by mass of silver.**

The equivalent mass is a pure number. When the equivalent mass of a substance is expressed in grams, it is called gram equivalent mass. For example, equivalent mass of sodium is 23, hence, its gram equivalent mass is 23 g.

The equivalent mass of a substance may have different values under different conditions. The equivalent mass of an element may vary with change of valency. For example, copper forms two oxides CuO and Cu_2O. In CuO, 63.5 parts of copper combine with 16 parts of oxygen. Thus, equivalent mass of copper in this oxide is $\frac{63.5}{2} = 31.75$. In Cu_2O, 2×63.5 parts of copper combine with 16 parts of oxygen; thus, the equivalent mass of copper in this oxide is:

$$\frac{2 \times 63.5}{2} = 63.5$$

Relation between atomic mass, equivalent mass and valency: Suppose an element X combines with hydrogen to form a compound, XH_n, where n is the valency of the element X.

n parts by mass of hydrogen combine with atomic mass of element X.

1 part by mass of hydrogen combines with

$$\frac{\text{Atomic mass of element}}{n}$$

By above definition, $\frac{\text{Atomic mass of element}}{n}$ is the equivalent mass of the element.

Thus, $$\text{Equivalent mass} = \frac{\text{Atomic mass}}{n}$$

or $$\text{Atomic mass} = \text{Equivalent mass} \times \text{Valency}$$

Note: Detailed discussion on equivalent masses of compounds (acids, bases, salts, oxidising agents, reducing agents, etc.,) will be taken in chapter on volumetric analysis.

The following methods are used for the determination of equivalent mass of elements.

(i) Hydrogen displacement method: This method is used for those elements which can evolve hydrogen from acids, *i.e.*, active metals. A known mass of the active metal is reacted with dilute mineral acid. Hydrogen gas thus evolved is measured under experimental conditions. The volume of hydrogen is then reduced to NTP conditions. The mass of liberated hydrogen is determined using density of hydrogen (0.00009 at NTP).

$$\text{Equivalent mass} = \frac{\text{Mass of element}}{\text{Mass of hydrogen}} \times 1.008$$

$$= \frac{\text{Mass of element} \times 1.008}{\text{Volume in mL of hydrogen displaced at NTP} \times 0.00009}$$

$$= \frac{\text{Mass of element} \times 11200}{\text{Volume in mL of hydrogen displaced at NTP}}$$

(ii) Oxide formation method: A known mass of the element is changed into oxide directly or indirectly. The mass of oxide is noted.

Mass of oxygen = (Mass of oxide − Mass of element)

Thus, the equivalent mass of the element

$$= \frac{\text{Mass of element}}{(\text{Mass of oxide} - \text{Mass of element})} \times 8$$

$$= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$$

(iii) Chloride formation method: A known mass of the element is changed into chloride directly or indirectly. The mass of the chloride is determined.

Mass of chlorine = (Mass of chloride − Mass of element)

Thus, the equivalent mass of the element

$$= \frac{\text{Mass of element} \times 35.5}{(\text{Mass of chloride} - \text{Mass of element})}$$

$$= \frac{\text{Mass of element} \times 35.5}{\text{Mass of chlorine}}$$

(iv) Metal to metal displacement method: A more active metal can displace less active metal from its salt's solution. For example, when zinc is added to copper sulphate, copper is precipitated. A known mass of active metal is added to the salt's solution of less active metal. The precipitated metal after drying is accurately weighed. The masses of the displacing metal and the displaced metal bear the same ratio as their equivalent masses. If E_1 and E_2 are the equivalent masses of two elements and m_1 and m_2 their respective masses, then,

$$\frac{m_1}{m_2} = \frac{E_1}{E_2}$$

Knowing the equivalent mass of one metal, the equivalent mass of the other metal can be calculated.

(v) Double decomposition method: This method is based on the following points:

(a) The mass of the compound reacted and the mass of product formed are in the ratio of their equivalent masses.

(b) The equivalent mass of the compound (electrovalent) is the sum of equivalent masses of its radicals.

(c) The equivalent mass of a radical is equal to the formula mass of the radical divided by its charge.

$$AB + CD \longrightarrow \underset{\text{ppt.}}{AD} + CB$$

$$\frac{\text{Mass of } AB}{\text{Mass of } AD} = \frac{\text{Equivalent mass of } AB}{\text{Equivalent mass of } AD}$$

$$= \frac{\text{Eq. mass of } A + \text{Eq. mass of } B}{\text{Eq. mass of } A + \text{Eq. mass of } D}$$

Knowing the equivalent masses of B and D, equivalent mass of A can be calculated.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

7. An unknown element forms an oxide. What will be the equivalent mass of the element if the oxygen content is 20% by mass: [JEE (WB) 2008]
(a) 16 (b) 32 (c) 8 (d) 64
[**Ans.** (b)]
[**Hint:** Equivalent mass of element $= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$
$= \frac{80}{20} \times 8 = 32$]

8. A metal M of equivalent mass E forms an oxide of molecular formula M_xO_y. The atomic mass of the metal is given by the correct equation: [PMT (Kerala) 2008]
(a) $2E(y/x)$ (b) xyE
(c) E/y (d) y/E
(e) $\frac{E}{2} \times \frac{x}{y}$
[**Ans.** (a)]
[**Hint:** Let atomic mass of metal M is 'a'.
Mass of metal $= a \times x$
Mass of oxygen $= 16 \times y$
Equivalent mass of element $= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$
$$E = \frac{ax}{16y} \times 8$$
$$a = 2E\left(\frac{y}{x}\right)$$]

9. The percentage of an element M is 53 in its oxide of molecular formula M_2O_3. Its atomic mass is about: [PET (Kerala) 2008]
(a) 45 (b) 9 (c) 18 (d) 38
(e) 27
[**Ans.** (e)]
[**Hint:** Equivalent mass of element $= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$
$= \frac{53}{47} \times 8 \simeq 9$
Atomic mass = Equivalent mass × Valency
$= 9 \times 3 = 27$ amu.]

10. The equivalent weight of a metal is double than that of oxygen. How many times is the weight of its oxide greater than the weight of metal?
(a) 4 (b) 2 (c) 3 (d) 1.5
[**Ans.** (d)]
[**Hint:** Equivalent mass of metal $= 16 = \frac{x}{n}$
Where x = atomic mass of metal
n = valency of metal
Molecular formula of metal oxide $= M_2O_n$
$$\frac{\text{Mass of metal oxide}}{\text{Mass of metal}} = \frac{2(16n) + 16(n)}{2(16n)} = 1.5$$]

1.15 METHODS FOR THE DETERMINATION OF ATOMIC MASS

(i) Dulong and Petit's Law: According to this law, the product of atomic mass and specific heat of a solid element is approximately equal to 6.4. The product of atomic mass and specific heat is called atomic heat. Thus,

$$\text{Atomic mass} \times \text{Specific heat} = 6.4$$

or $$\text{Atomic mass (approximate)} = \frac{6.4}{\text{Specific heat}}$$

In above formula, the specific heat must be in cal/g unit.

The equivalent mass of the element is determined experimentally and the valency, which is always a whole number, can be obtained by dividing approximate atomic mass with the equivalent mass and changing the value so obtained to the nearest whole number. In this way, exact atomic mass can be determined by multiplying equivalent mass with valency.

Example 26. *A chloride of an element contains 49.5% chlorine. The specific heat of the element is 0.056. Calculate the equivalent mass, valency and atomic mass of the element.*

Solution: Mass of chlorine in the metal chloride = 49.5%

Mass of metal $= (100 - 49.5) = 50.5$

$$\text{Equivalent mass of the metal} = \frac{\text{Mass of metal}}{\text{Mass of chlorine}} \times 35.5$$

$$= \frac{50.5}{49.5} \times 35.5 = 36.21$$

According to Dulong and Petit's law,

$$\text{Approximate atomic mass of the metal} = \frac{6.4}{\text{Specific heat}}$$

$$= \frac{6.4}{0.056} = 114.3$$

$$\text{Valency} = \frac{\text{Approximate atomic mass}}{\text{Equivalent mass}} = \frac{114.3}{36.21} = 3.1 \approx 3$$

Hence, exact atomic mass $= 36.21 \times 3 = 108.63$

Example 27. *On dissolving 2.0 g of metal in sulphuric acid, 4.51 g of the metal sulphate was formed. The specific heat of the metal is 0.057 cal g^{-1}. What is the valency of the metal and exact atomic mass?*

Solution: Equivalent mass of SO_4^{2-} radical

$$= \frac{\text{Ionic mass}}{\text{Valency}} = \frac{96}{2} = 48$$

Mass of metal sulphate = 4.51 g

Mass of metal = 2.0 g

Mass of sulphate radical = (4.51 − 2.0) = 2.51 g

2.51 g of sulphate combine with 2.0 g of metal.

So, 48 g of sulphate will combine with

$$= \frac{2}{2.51} \times 48 = 38.24 \text{ g metal}$$

Equivalent mass of metal = 38.24

According to Dulong and Petit's law,

$$\text{Approximate atomic mass} = \frac{6.4}{\text{Specific heat}} = \frac{6.4}{0.057} = 112.5$$

$$\text{Valency} = \frac{\text{Approximate atomic mass}}{\text{Equivalent mass}}$$

$$= \frac{112.5}{38.24} = 2.9 \approx 3$$

Exact atomic mass = 38.24 × 3 = 114.72

(ii) Cannizzaro's method: Atomic mass of an element may be defined as the smallest mass of the element present in the molecular mass of any one of its compounds. For this purpose, the following steps are followed:

(a) Molecular masses of a number of compounds in which the element is present are determined.

(b) Each compound is analysed. Mass of the element is determined in the molecular mass of each compound.

(c) The lowest mass of the element is taken its atomic mass.

The following table shows the application of this method:

Compound	Vapour density (V.D.)	Molecular mass = 2 V.D.	% of carbon by mass in compound	Mass of carbon in one molecular mass of the compound
Methane	8	16	75.0	$\frac{75.0 \times 16}{100} = 12$ g
Ethane	15	30	80.0	$\frac{80.0 \times 30}{100} = 24$ g
Carbon monoxide	14	28	42.9	$\frac{42.9 \times 28}{100} = 12$ g
Carbon dioxide	22	44	27.3	$\frac{27.3 \times 44}{100} = 12$ g
Propane	22	44	81.8	$\frac{81.8 \times 44}{100} = 36$ g

Least mass of carbon is 12 g.

Thus, the atomic mass of carbon is 12.

(iii) The law of isomorphism: Isomorphous substances form crystals which have same shape and size and can grow in the saturated solution of each other. They have a property of forming mixed crystals. Isomorphous substances have same composition, *i.e.*, they have same number of atoms arranged similarly.

Examples of isomorphous compounds are:

(a) K_2SO_4 and K_2CrO_4 (potassium sulphate and potassium chromate)

(b) $ZnSO_4 \cdot 7H_2O$ and $FeSO_4 \cdot 7H_2O$ (zinc sulphate and ferrous sulphate)

(c) $KClO_4$ and $KMnO_4$ (potassium perchlorate and potassium permanganate)

(d) $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$ and $K_2SO_4 \cdot Cr_2(SO_4)_3 \cdot 24H_2O$ (potash alum and chrome alum).

The following conclusions have been deduced from the phenomenon of isomorphism:

(i) Masses of two elements that combine with same mass of other elements in their respective compounds are in the ratio of their atomic masses.

$$\frac{\text{Mass of one element } (A) \text{ that combines with a certain mass of other elements}}{\text{Mass of other element } (B) \text{ that combines with the same mass of other elements}} = \frac{\text{Atomic mass of } A}{\text{Atomic mass of } B}$$

(ii) The valencies of the elements forming isomorphous compounds are the same.

Example 28. *Potassium chromate is isomorphous to potassium sulphate* (K_2SO_4) *and is found to contain 26.78% chromium. Calculate the atomic mass of chromium* ($K = 39.10$).

Solution: Since, the formula of potassium sulphate is K_2SO_4, so the formula of potassium chromate should be K_2CrO_4 as it is isomorphous to K_2SO_4.

If the atomic mass of chromium is A, then

formula mass of potassium chromate should be

$$= 2 \times 39.1 + A + 64 = (142.2 + A)$$

$$\% \text{ of chromium} = \frac{A}{(142.2 + A)} \times 100$$

So, $$\frac{100A}{(142.2 + A)} = 26.78$$

$$100A = 26.78\,(142.2 + A)$$

or $$A = \frac{26.78 \times 142.2}{73.22} = 52.00$$

(iv) Atomic mass from vapour density of a chloride: The following steps are involved in this method:

(a) Vapour density of the chloride of the element is determined.

(b) Equivalent mass of the element is determined.

Let the valency of the element be x. The formula of its chloride will be MCl_x.

$$\text{Molecular mass} = \text{Atomic mass of } M + 35.5x$$

$$= A + 35.5x$$

$$\text{Atomic mass} = \text{Equivalent mass} \times \text{Valency}$$

$$A = E \times x$$

$$\text{Molecular mass} = E \times x + 35.5x$$

$$2\,V.D. = x(E + 35.5)$$

$$x = \frac{2\,V.D.}{E + 35.5}$$

Knowing the value of valency, the atomic mass can be determined.

Example 29. *One gram of a chloride was found to contain* 0.835 *g of chlorine. Its vapour density is* 85. *Calculate its molecular formula.*

Solution: Mass of metal chloride = 1g

$$\text{Mass of chlorine} = 0.835\,g$$

$$\text{Mass of metal} = (1 - 0.835) = 0.165\,g$$

$$\text{Equivalent mass of metal} = \frac{0.165 \times 35.5}{0.835} = 7.01$$

$$\text{Valency of the metal} = \frac{2\,V.D.}{E + 35.5} = \frac{2 \times 85}{7.01 + 35.5} = 4$$

$$\text{Formula of the chloride} = MCl_4$$

Example 30. *The oxide of an element contains* 32.33 *per cent of the element and the vapour density of its chloride is* 79. *Calculate the atomic mass of the element.*

Solution: Mass of the element = 32.33 parts

Mass of oxygen = (100 − 32.33) = 67.67 parts

$$\text{Equivalent mass of the element} = \frac{32.33}{67.67} \times 8 = 3.82$$

$$\text{Valency of the element} = \frac{2\,V.D.}{E + 35.5} = \frac{2 \times 79}{3.82 + 35.5} = 4$$

Hence, the atomic mass of the element = 3.82 × 4

= 15.28

1.16 TYPES OF FORMULAE

As already stated in section 1.10, a formula is a group of symbols of the elements which represents one molecule of the substance. Formula represents chemical composition of the substance. There are three kinds of formulae in the case of compounds.

(i) Empirical formula: It represents the simplest relative whole number ratio of atoms of each element present in the molecule of the substance. For example, CH is the empirical formula of benzene in which ratio of the atoms of carbon and hydrogen is 1 : 1. It also indicates that the ratio of carbon and hydrogen is 12 : 1 by mass.

(ii) Molecular formula: Molecular formula of a compound is one which expresses as the actual number of atoms of each element present in one molecule. C_6H_6 is the molecular formula of benzene indicating that six carbon atoms and six hydrogen atoms are present in a molecule of benzene. Thus,

$$\text{Molecular formula} = n \times \text{Empirical formula}$$

where, $$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

Molecular formula gives the following informations:

(i) Various elements present in the molecule.

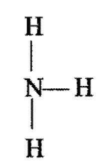
(ii) Number of atoms of various elements in the molecule.

(iii) Mass ratio of the elements present in the molecule. The mass ratio of carbon and oxygen in CO_2 molecule is 12 : 32 or 3 : 8.

(iv) Molecular mass of the substance.

(v) The number written before the formula indicates the number of molecules, *e.g.*, $2CO_2$ means 2 molecules of carbon dioxide.

(iii) Structural formula: It represents the way in which atoms of various elements present in the molecule are linked with one another. For example, ammonia is represented as:



The formula indicates that three hydrogen atoms are linked to one nitrogen atom by three single covalent bonds.

1.17 PERCENTAGE COMPOSITION OF A COMPOUND

Percentage composition of a compound is the relative mass of the each of the constituent element in 100 parts of it. It is readily calculated from the formula of the compound. Molecular mass of a compound is obtained from its formula by adding up the masses of all the atoms of the constituent elements present in the molecule.

Let the molecular mass of a compound be M and X be the mass of an element in the molecule.

$$\text{Percentage of element} = \frac{\text{Mass of element}}{M} \times 100 = \frac{X}{M} \times 100$$

Example 31. *Calculate the percentage composition of calcium nitrate.*

Solution: The formula of calcium nitrate is $Ca(NO_3)_2$.

Thus, the formula mass or molecular mass

= At. mass of Ca + 2 × At. mass of N + 6 × At. mass of oxygen

= 40 + 2 × 14 + 6 × 16

= 164

$$\%\text{ of Ca} = \frac{40}{164} \times 100 = 24$$

$$\%\text{ of N} = \frac{28}{164} \times 100 = 17$$

$$\%\text{ of O} = 100 - (24 + 17) = 59$$

Example 32. *Determine the percentage of water of crystallisation, iron, sulphur and oxygen in pure ferrous sulphate* ($FeSO_4 \cdot 7H_2O$).

Solution: The formula mass of ferrous sulphate

= At. mass of Fe + At. mass of S + 4 × At. mass of oxygen + 7 × Mol. mass of H_2O

$= 56.0 + 32.0 + 4 \times 16.0 + 7 \times 18.0$

$= 278.0$

So, % of water of crystallisation $= \frac{126}{278} \times 100 = 45.32$

$$\% \text{ of iron} = \frac{56}{278} \times 100 = 20.14$$

$$\% \text{ of sulphur} = \frac{32}{278} \times 100 = 11.51$$

$$\% \text{ of oxygen} = \frac{64}{278} \times 100 = 23.02$$

(Oxygen present in water molecules is not taken into account.)

Example 33. *It is found that 16.5 g of metal combine with oxygen to form 35.60 g of metal oxide. Calculate the percentage of metal and oxygen in the compound.*

Solution:

Mass of oxygen in oxide $= (35.60 - 16.50) = 19.10$ g

$$\% \text{ of metal} = \frac{16.50}{35.60} \times 100 = 46.3$$

$$\% \text{ of oxygen} = \frac{19.10}{35.60} \times 100 = 53.7$$

Example 34. *Hydrogen and oxygen are combined in the ratio 1 : 16 by mass in hydrogen peroxide. Calculate the percentage of hydrogen and oxygen in hydrogen peroxide.*

Solution: 17 parts of hydrogen peroxide contain hydrogen = 1 part

100 parts of hydrogen peroxide contain hydrogen

$$= \frac{1}{17} \times 100 = 5.88$$

$$\% \text{ of oxygen} = (100 - 5.88) = 94.12$$

Example 35. *On analysis of an impure sample of sodium chloride, the percentage of chlorine was found to be 45.5. What is the percentage of pure sodium chloride in the given sample?*

Solution: The molecular mass of pure sodium chloride (NaCl)

= At. mass of Na + At. mass of chlorine

$= (23 + 35.5) = 58.5$

% of chlorine in pure NaCl

$$= \frac{35.5}{58.4} \times 100 = 60.6$$

Thus,

% of purity of NaCl in the sample

$$= \frac{45.5}{60.6} \times 100 = 75$$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

11. A gas mixture contains 50% helium and 50% methane by volume. What is the percentage by mass of methane in the mixture? [CEE (Kerala) 2004]

(a) 19.97% (b) 20.05% (c) 50% (d) 75%
(e) 80.03%

[**Ans.** (e)]

[**Hint:** Molar and volume ratio will be same, *i.e.*, 1 : 1.

∴ Mass of 1 mole CH_4 and He will be 16 and 4 g respectively.

$$\text{Percentage by mass of } CH_4 = \frac{\text{Mass of } CH_4}{\text{Total mass}} \times 100$$

$$= \frac{16}{20} \times 100 \approx 80\%]$$

12. The atomic composition of the entire universe is approximately given in the table below:

Atom	% of total no. of atoms
H	93
He	7

Hydrogen atoms constitute what percentage of the universe by mass?

(a) 77% (b) 23% (c) 37% (d) 73%

[**Ans.** (a)]

[**Hint:** Mass of 93 'H' atoms = 93 amu

Mass of 7 'He' atoms = 28 amu

$$\% \text{ Hydrogen by mass} = \frac{93}{(93 + 28)} \times 100 = 77\%]$$

13. Which pair of species has same percentage of carbon?

(a) CH_3COOH and $C_6H_{12}O_6$
(b) CH_3COOH and C_2H_5OH
(c) $HCOOCH_3$ and $C_{12}H_{22}O_{11}$
(d) $C_6H_{12}O_6$ and $C_{12}H_{22}O_{11}$

[**Ans.** (a)]

[**Hint:** Percentage of carbon in acetic acid $= \frac{24}{60} \times 100 = 40\%$

Percentage of carbon in $C_6H_{12}O_6 = \frac{72}{180} \times 100 = 40\%$]

14. Which of the following alkanes has 75% of carbon?

(a) C_2H_6 (b) CH_4 (c) C_3H_8 (d) C_4H_{10}

[**Ans.** (b)]

[**Hint:** Percentage of carbon in methane $= \frac{12}{16} \times 100 = 75\%$]

15. Which of the following two oxides of nitrogen have 30.5% nitrogen?

(a) NO (b) NO_2 (c) N_2O_4 (d) N_2O_5

[**Ans.** (b) and (c)]

[**Hint:** Percentage of nitrogen in $NO_2 = \frac{14}{46} \times 100 = 30.5\%$

Percentage of nitrogen in $N_2O_4 = \frac{28}{92} \times 100 = 30.5\%$]

1.18 DETERMINATION OF EMPIRICAL AND MOLECULAR FORMULAE

The following steps are followed to determine the empirical formula of the compound :

(i) The percentage composition of the compound is determined by quantitative analysis.

(ii) The percentage of each element is divided by its atomic mass. It gives atomic ratio of the elements present in the compound.

(iii) The atomic ratio of each element is the divided by the minimum value of atomic ratio as to get the simplest ratio of the atoms of elements present in the compound.

(iv) If the simplest ratio is fractional, then values of simplest ratio of each element is multiplied by a smallest integer to get a simplest whole number for each of the element.

(v) To get the empirical formula, symbols of various elements present are written side by side with their respective whole number ratio as a subscript to the lower right hand corner of the symbol.

The molecular formula of a substance may be determined from the empirical formula if the molecular mass of the substance is known. The molecular formula is always a simple multiple of empirical formula and the value of simple multiple is obtained by dividing molecular mass with empirical formula mass.

Example 36. *Calculate the empirical formula for a compound that contains* 26.6% *potassium,* 35.4% *chromium and* 38.1% *oxygen.*

[*Given* $K = 39.1; Cr = 52; O = 16$]

Solution:

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio	Simplest whole number ratio
Potassium	26.6	39.1	$\frac{26.6}{39.1} = 0.68$	$\frac{0.68}{0.68} = 1$	$1 \times 2 = 2$
Chromium	35.4	52.0	$\frac{35.4}{52} = 0.68$	$\frac{0.68}{0.68} = 1$	$1 \times 2 = 2$
Oxygen	38.1	16.0	$\frac{38.1}{16} = 2.38$	$\frac{2.38}{0.68} = 3.5$	$3.5 \times 2 = 7$

Therefore, empirical formula is $K_2Cr_2O_7$.

Example 37. *A compound contains* 34.8% *oxygen,* 52.2% *carbon and* 13.0% *hydrogen. What is the empirical formula mass of the compound?*

Solution:

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Oxygen	34.8	16	$\frac{34.8}{16} = 2.175$	$\frac{2.175}{2.175} = 1$
Carbon	52.2	12	$\frac{52.2}{12} = 4.35$	$\frac{4.35}{2.175} = 2$
Hydrogen	13.0	1	$\frac{13.0}{1} = 13.0$	$\frac{13.0}{2.175} = 6$

The empirical formula is C_2H_6O.

Empirical formula mass $= (2 \times 12) + (6 \times 1) + 16 = 46$

Example 38. *A compound of carbon, hydrogen and nitrogen contains these elements in the ratio* 9 : 1 : 3.5. *Calculate the empirical formula. If its molecular mass is* 108, *what is the molecular formula?*

Solution:

Element	Element ratio	Atomic mass	Relative number of atoms	Simplest ratio
Carbon	9	12	$\frac{9}{12} = 0.75$	$\frac{0.75}{0.25} = 3$
Hydrogen	1	1	$\frac{1}{1} = 1$	$\frac{1}{0.25} = 4$
Nitrogen	3.5	14	$\frac{3.5}{14} = 0.25$	$\frac{0.25}{0.25} = 1$

The empirical formula $= C_3H_4N$

Empirical formula mass $= (3 \times 12) + (4 \times 1) + 14 = 54$

$$n = \frac{\text{Mol. mass}}{\text{Emp. mass}} = \frac{108}{54} = 2$$

Thus, molecular formula of the compound

$$= 2 \times \text{Empirical formula}$$
$$= 2 \times C_3H_4N = C_6H_8N_2$$

Example 39. *A carbon compound containing only carbon and oxygen has an approximate molecular mass of* 290. *On analysis, it is found to contain* 50% *by mass of each element. What is the molecular formula of the compound?*

Solution:

Element percentage	Atomic mass	Relative number of atoms	Simplest ratio	Simplest whole number ratio
Carbon 50.0	12	4.166	$\frac{4.166}{3.125} = 1.33$	4
Oxygen 50.0	16	3.125	$\frac{3.125}{3.125} = 1$	3

The empirical formula $= C_4O_3$

Empirical formula mass $= (4 \times 12) + (3 \times 16) = 96$

Molecular mass $= 290$

$$n = \frac{\text{Mol. mass}}{\text{Emp. mass}} = \frac{290}{96} = 3 \text{ approximately}$$

Molecular formula $= n \times$ Empirical formula

$$= 3 \times C_4O_3 = C_{12}O_9$$

Example 40. *A compound on analysis, was found to have the following composition: (i) Sodium* $= 14.31\%$, *(ii) Sulphur* $= 9.97\%$, *(iii) Oxygen* $= 69.50\%$, *(iv) Hydrogen* $= 6.22\%$. *Calculate the molecular formula of the compound assuming that whole of hydrogen in the compound is present as water of crystallisation. Molecular mass of the compound is* 322.

Solution:

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Sodium	14.31	23	0.622	$\frac{0.622}{0.311} = 2$
Sulphur	9.97	32	0.311	$\frac{0.311}{0.311} = 1$
Hydrogen	6.22	1	6.22	$\frac{6.22}{0.311} = 20$
Oxygen	69.50	16	4.34	$\frac{4.34}{0.311} = 14$

The empirical formula $= Na_2SH_{20}O_{14}$

Empirical formula mass $= (2 \times 23) + 32 + (20 \times 1) + (14 \times 16)$
$= 322$

Molecular mass $= 322$

Molecular formula $= Na_2SH_{20}O_{14}$

Whole of the hydrogen is present in the form of water. Thus, 10 water molecules are present in the molecule.

So, molecular formula $= Na_2SO_4 \cdot 10H_2O$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

16. An organic compound contains 49.30% carbon, 6.84% hydrogen and its vapour density is 73. Molecular formula of the compound is: **[CET (Kerala) 2004]**
(a) $C_3H_8O_2$ (b) $C_3H_{10}O_2$ (c) C_6H_9O (d) $C_4H_{10}O_2$
(e) $C_6H_{10}O_4$
[Ans. (e)]
[Hint: Molecular mass $= 2 \times 73 = 146$

$$C = \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}} = \frac{49.30}{100} \times \frac{146}{12} = 6$$

$$H = \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}} = \frac{6.84}{100} \times \frac{146}{1} = 10$$

$$O = \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}} = \frac{43.86}{100} \times \frac{146}{16} = 4$$

Molecular formula $= C_6H_{10}O_4$

Molecular mass $= 12 \times 6 + 10 \times 1 + 16 \times 4 = 146$

'Or'

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Carbon	49.30	12	4.10	$1.5 \times 2 = 3$
Hydrogen	6.84	1	6.84	$2.5 \times 2 = 5$
Oxygen	43.86	16	2.74	$1 \times 2 = 2$

The empirical formula $= C_3H_5O_2$

$$n = \frac{2 \times 73}{73} = 2$$

Molecular formula $= 2 \times C_3H_5O_2 = C_6H_{10}O_4$]

17. A compound has an empirical formula C_2H_4O. An independent analysis gave a value of 132.16 for its molecular mass. What is the correct molecular formula? **[CET (Kerala) 2004]**
(a) $C_4H_4O_5$ (b) $C_{10}H_{12}$
(c) C_7O_3 (d) $C_6H_{12}O_3$
(e) $C_4H_8O_5$
[Ans. (d)]
[Hint: Molecular formula $= (C_2H_4O)_n$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{132.16}{44} = 3$$

Molecular formula $= (C_2H_4O)_3 = C_6H_{12}O_3$]

18. An organic compound containing C and H has 92.30% carbon. Its empirical formula is:
(a) CH (b) CH_3
(c) CH_2 (d) CH_4
[Ans. (a)]
[Hint: Percentage of carbon $= \frac{12}{13} \times 100 = 92.30\%$

'Or'

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
Carbon	92.30	12	7.69	1
Hydrogen	7.70	1	7.70	1

Empirical formula = CH]

19. Two oxides of a metal contain 50% and 40% of metal M respectively. If the formula of first oxide is MO, the formula of 2nd oxide will be:
(a) MO_2 (b) M_2O_3
(c) M_2O (d) M_2O_5
[Ans. (b)]
[Hint:

Compound 1		Compound 2	
M	O	M	O
50%	50%	40%	60%
50 g	50 g	40 g	60 g
1 g	$\frac{50}{50} = 1$ g	1 g	$\frac{60}{40} = 1.5$ g
2 g	2 g	2 g	3 g

Formula: MO M_2O_3]

20. Two elements X and Y have atomic mass 75 and 16 respectively. They combine to give a compound having 75.8% X. The formula of the compound is:
(a) XY (b) X_2Y
(c) X_2Y_2 (d) X_2Y_3
[Ans. (d)]
[Hint: Molecular mass of $X_2Y_3 = 2 \times 75 + 3 \times 16 = 198$

$$\text{Percentage of } X = \frac{150}{198} \times 100 = 75.80\%$$

'Or'

Element	Percentage	Atomic mass	Relative number of atoms	Simplest ratio
X	75.80	75	1.01	$1 \times 2 = 2$
Y	24.20	16	1.51	$1.5 \times 2 = 3$

Formula = X_2Y_3]

21. The crystalline salt $Na_2SO_4 \cdot xH_2O$ on heating loses 55.9% of its mass. The formula of crystalline salt is:
(a) $Na_2SO_4 \cdot 5H_2O$ (b) $Na_2SO_4 \cdot 7H_2O$
(c) $Na_2SO_4 \cdot 2H_2O$ (d) $Na_2SO_4 \cdot 10H_2O$
(e) $Na_2SO_4 \cdot 6H_2O$ [PMT (Kerala) 2007]
[Ans. (d)]
[Hint: Molecular mass of $Na_2SO_4 \cdot 10H_2O$
$= 46 + 96 + 180 = 322$ amu

$$\% \text{ by mass of } H_2O = \frac{180}{322} \times 100 = 55.9\%]$$

1.19 CHEMICAL EQUATION

A chemical equation is a symbolic representation of a chemical change.

The substances, in which the chemical change is brought, are called reactants and the substances which come into existence as the result of chemical change are called products. The relationship between reactants and products is represented in the form of a chemical equation. The symbols or formulae of the reactants are written on left hand side of equality (=) or → sign and the symbols or formulae of products on right hand side. The symbols or formulae on both the sides are added by + sign. Such an equation is known as skeleton equation. The equation becomes balanced when total number of atoms of various elements are made equal on both the sides. Gases are always written in molecular form.

$$KClO_3 \longrightarrow KCl + O_2$$

This is the skeleton equation as it only represents reactant and products involved in the chemical change but the following equation is a balanced equation as the number of atoms of various elements is equal on both sides.

$$\underset{\text{Reactant}}{2KClO_3} = \underbrace{2KCl + 3O_2}_{\text{Products}}$$

The following notations are also used in chemical equations as to provide more information about chemical change:

(i) Upper arrow (↑) is written immediately after the gaseous product.
(ii) Lower arrow (↓) is written immediately after the insoluble substance (solid) which deposits from a solution.
(iii) Symbols, (s) for solid, (l) for liquid and (g) for gas are also written to represent the physical state of the reactants and products.
(iv) Symbol ($aq.$) is written for substances dissolved in water.
(v) Symbol (Δ) is written over an arrow or over an equality sign to represent heating.

Information Obtained from Chemical Equation

A balanced chemical equation provides the following informations:

(i) What are the reactants and products involved in the chemical change?
(ii) The relative number of molecules of reactants and products.
(iii) The relative number by parts of mass of reactants and products.
(iv) Relative volumes of gaseous reactants and products.

For example, consider the following reaction:

$$CH_4(g) + 2O_2(g) = CO_2(g) + 2H_2O(g)$$

This equation tells us that methane and oxygen are reactants and carbon dioxide and water are products. One molecule of methane reacts with two molecules of oxygen to produce one molecule of CO_2 and two molecules of water or one mole of methane reacts with two moles of oxygen to produce one mole of carbon dioxide and two moles of water or 16 g of methane reacts with 64 g of oxygen to produce 44 g of CO_2 and 36 g of water. This equation also tells that 1 vol. of methane reacts with 2 vol. of oxygen to produce 1 vol. of CO_2 and 2 vol. of steam under similar conditions of temperature and pressure.

Limitations of Chemical Equation

A chemical equation fails to provide the following informations:

(i) Actual concentration of the reactants taken and the actual concentration of the products obtained.
(ii) Time taken for the completion of the chemical change.
(iii) Conditions applied for bringing the chemical change.
(iv) Whether the reaction is reversible or irreversible.

The following efforts have been made to make the chemical equations more informative by introducing:

(i) Experimental conditions: If a particular chemical change occurs under certain temperature and pressure conditions, these are mentioned above and below the (→) or (=) sign.

$$N_2 + 3H_2 \xrightarrow[450°C]{200\text{ atm}} 2NH_3$$

If the reaction occurs in presence of a catalyst, it is written above the (→) or (=) sign.

$$2SO_2 + O_2 \xrightarrow{Pt} 2SO_3$$

(ii) Heat evolved or absorbed: Heat evolved or absorbed in a chemical change can be represented by adding or subtracting the amount of heat on right hand side.

$$N_2 + O_2 \longrightarrow 2NO - 43.2 \text{ kcals}$$
$$C + O_2 \longrightarrow CO_2 + 94.3 \text{ kcals}$$

(iii) Reversible or irreversible nature: Reversible reactions are shown by changing the sign of equality (=) or arrow (→) with sign of double arrow ($\rightleftharpoons$).

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Types of Chemical Equations

Chemical equations are of two types:

(i) Molecular equations
(ii) Ionic equations.

Molecular equations are those in which reactants and products are represented in the form of molecules.

$$BaCl_2 + Na_2SO_4 = BaSO_4 \downarrow + 2NaCl$$
$$2NaOH + H_2SO_4 = Na_2SO_4 \downarrow + 2H_2O$$

Ionic equations are those in which reactants and products are written in ionic form. The molecular equation

$$BaCl_2 + Na_2SO_4 = BaSO_4 \downarrow + 2NaCl$$

can be written in ionic form as:

$$Ba^{2+} + 2Cl^- + 2Na^+ + SO_4^{2-} = BaSO_4 \downarrow + 2Na^+ + 2Cl^-$$
$$Ba^{2+} + SO_4^{2-} = BaSO_4 \downarrow$$

Note: Calculations based on chemical equations have been dealt in the chapter 'Stoichiometry' in 'Inorganic Chemistry'.

1.20 MEASUREMENT IN CHEMISTRY: FUNDAMENTAL AND DERIVED UNITS

Chemistry is an experimental science. An experiment always involves observation of a phenomenon under certain set of conditions. The quantitative scientific observation generally requires the measurement of one or more physical quantities such as mass, length, density, volume, pressure, temperature, etc.

A physical quantity is expressed in terms of a number and a unit. Without mentioning the unit, the number has no meaning. For example, the distance between two points is "four" has no meaning unless a specific unit (inch, centimetre, metre, etc.,) is associated with the number. The units of physical quantities depend on three basic units, *i.e.*, units of mass, length and time. Since, these are independent units and cannot be derived from any other units, they are called **fundamental units.** It was soon realised that the three fundamental units cannot describe all the physical quantities such as temperature, intensity of luminosity, electric current and the amount of the substance. Thus, seven units of measurement, namely mass, length, time, temperature, electric current, luminous intensity and amount of substance are taken as **basic units.** All other units can be derived from them and are, therefore, called **derived units.** The units of area, volume, force, work, density, velocity, energy, etc., are all derived units.

SI Units of Measurement

Various systems of units were in use prior to 1960. The common ones are the following:

(i) The English or FPS system: The system uses the **foot,** the **pound** and the **second** for length, mass and time measurements respectively. It is not used now-a-days.

(ii) MKS system: Here M stands for **metre** (a unit of length), K for **kilogram** (a unit of mass) and S for **second** (a unit of time). This is a decimal system.

(iii) CGS system: Here the unit of length is **centimetre,** the unit of mass is **gram** and the unit of time is **second.** It is also a decimal system.

MKS system often known as metric system was very popular throughout the world, but the drawback with this system was that a number of different metric units for the same quantity were used in different parts of the world. In 1964, the National Bureau of Standards adopted a slightly modified version of the metric system, which had been officially recommended in 1960 by an international body, **General Conference of Weights and Measures.** This revised set of units is known as the **International System of Units** (abbreviated SI). Now the SI units have been accepted by the scientists all over the world in all branches of science, engineering and technology.

The SI system have seven basic units. The various fundamental quantities that are expressed by these units along with their symbols are tabulated below:

Basic physical quantity	Unit	Symbol
Length	Metre	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Electric current	Ampere	amp or A
Luminous intensity	Candela	cd
Amount of substance	Mole	mol

Sometimes, submultiples and multiples are used to reduce or enlarge the size of the different units. The names and symbols of sub-multiples and multiples are listed in the table given below.

The name for the base unit for mass, the kilogram, already contains a prefix. The names of other units of mass are obtained by substituting other prefixes for prefix kilo. The names of no other base units contain prefixes.

The use of SI system is slowly growing, however, older systems are still in use. Furthermore, the existence of older units in scientific literature demands that one must be familiar with both old and new systems.

Submultiples			Multiples		
Prefix	**Symbol**	**Sub-multiple**	**Prefix**	**Symbol**	**Multiple**
deci	d	10^{-1}	deca	da	10
centi	c	10^{-2}	hecto	h	10^2
milli	m	10^{-3}	kilo	k	10^3
micro	μ	10^{-6}	mega	M	10^6
nano	n	10^{-9}	giga	G	10^9
pico	p	10^{-12}	tera	T	10^{12}
femto	f	10^{-15}	peta	P	10^{15}
atto	a	10^{-18}	exa	E	10^{18}
zepto	z	10^{-21}	zeta	Z	10^{21}
yocto	y	10^{-24}	yotta	Y	10^{24}

Greek Alphabets

Alpha	A	α	Nu	N	ν
Beta	B	β	Xi	Ξ	ξ
Gamma	Γ	γ	Omicron	O	o
Delta	Δ	δ	Pi	Π	π
Epsilon	E	ε	Rho	P	ρ
Zeta	Z	ζ	Sigma	Σ	σ
Eta	H	η	Tau	τ	τ

Theta	Θ	θ	Upsilon	Υ	υ
Iota	I	ι	Phi	Φ	ϕ
Kappa	K	κ	Chi	X	χ
Lambda	Λ	λ	Psi	Ψ	ψ
Mu	M	μ	Omega	Ω	ω

Numerical Prefix

Prefix	Value	Prefix	Value
Hemi	(1/2)	Deca	10
Mono	1	Undeca	11
Sesqui	$1\frac{1}{2}$	Dodeca	12
Di or Bi	2	Trideca	13
Tri	3	Tetradeca	14
Tetra	4	Pentadeca	15
Penta	5	Hexadeca	16
Hexa	6	Heptadeca	17
Hepta	7	Octadeca	18
Octa	8	Nonadeca	19
Nona	9	Eicosa	20

SI Units for Some Common Derived Quantities

(a) $$\text{Area} = \text{length} \times \text{breadth}$$
$$= \text{m} \times \text{m} = \text{m}^2 \text{ [square metre]}$$

(b) $$\text{Volume} = \text{length} \times \text{breadth} \times \text{height}$$
$$= \text{m} \times \text{m} \times \text{m} = \text{m}^3 \text{ [cubic metre]}$$

(c) $$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{\text{kg}}{\text{m}^3} = \text{kg m}^{-3}$$

(d) $$\text{Speed} = \frac{\text{distance covered}}{\text{time}} = \frac{\text{metre}}{\text{time}} = \text{m s}^{-1}$$

(e) $$\text{Acceleration} = \frac{\text{change in velocity}}{\text{time taken}} = \frac{\text{m s}^{-1}}{\text{s}} = \text{m s}^{-2}$$

(f) $$\text{Force} = \text{mass} \times \text{acceleration}$$
$$= \text{kg} \times \text{m s}^{-2}$$
$$= \text{kg m s}^{-2} \text{ (Newton, abbreviated as N)}$$

(g) $$\text{Pressure} = \text{force per unit area}$$
$$= \frac{\text{kg m s}^{-2}}{\text{m}^2} = \text{kg m}^{-1}\text{ s}^{-2} \text{ or Nm}^{-2}$$
(Pascal—Pa)

(h) $$\text{Energy} = \text{force} \times \text{distance travelled}$$
$$= \text{kg m s}^{-2} \times \text{m}$$
$$= \text{kg m}^2 \text{ s}^{-2}$$
(joule—J)

Some Old Units Still in Use

The use of some of the old units is still permitted. The 'litre', for example, which is defined as 1 cubic decimetre is used frequently by chemists. Certain other units which are not a part of SI units are still retained for a limited period of time. The term atmosphere (atm), the unit of pressure, falls into this category. Few of the old units along with conversion factors are given below:

Length: The interatomic distances are reported in units of angstrom (Å), nanometre (nm) or picometre (pm).

$$1\,\text{Å} = 10^{-8} \text{ cm} = 10^{-10} \text{ m}$$
$$1\,\text{nm} = 10^{-7} \text{ cm} = 10^{-9} \text{ m} = 10\,\text{Å}$$
$$1\,\text{pm} = 10^{-10} \text{ cm} = 10^{-12} \text{ m} = 10^{-2}\,\text{Å}$$
$$1\,\text{nm} = 10^{3} \text{ pm}$$

Mass: The basic unit of mass is generally taken as gram (g). The gram is 10^{-3} kg.

$$1 \text{ kilogram (kg)} = 10^3 \text{ g}$$
$$1 \text{ milligram (mg)} = 10^{-3} \text{ g}$$
$$1 \text{ microgram } (\mu\text{g}) = 10^{-6} \text{ g}$$

While dealing with atoms and molecules, the term atomic mass unit (amu) is used. One amu is taken exactly as $\frac{1}{12}$ of the mass of an atom of the carbon isotope, C^{12}.

$$1\,\text{amu} = 1.6605 \times 10^{-24} \text{ g} = 1.6605 \times 10^{-27} \text{ kg}$$

Volume: The units of volume are reported as cubic centimetre (cm^3) and cubic decimetre (dm^3). Cubic decimetre is termed **litre** while cubic centimetre is termed **millilitre.**

$$1 \text{ litre (lit or L)} = (10\,\text{cm})^3 = 1000\,\text{cm}^3 = 10^{-3} \text{ m}^3$$
$$1 \text{ millilitre (mL)} = (1\,\text{cm})^3 = 1\,\text{cm}^3 \text{ (cc)} = 10^{-6} \text{ m}^3$$

So, $$1 \text{ litre} = 1000 \text{ mL}$$

Temperature: The **celsius temperature scale** which is not a part of SI system, is employed in scientific studies. This scale is based on the assignment of 0°C to the normal freezing point of water and 100°C to the normal boiling point of water. The celsius scale was formerly called the **centigrade scale.**

The unit of temperature in SI system is **Kelvin.** A degree on the kelvin scale has the same magnitude as the degree on the celsius scale but zero on the kelvin scale is equal to –273.15°C. The temperature (0 K) is often referred to as **absolute zero.**

So, $$K = (°C + 273.15)$$
or $$°C = (K - 273.15)$$

There is another important temperature scale known as **fahrenheit scale.** In this scale, the normal freezing point of water is 32°F and normal boiling point is 212°F. Thus, 100°C equals 180°F. Both the scales are related by the following equations:

$$°C = \frac{5}{9} \times (°F - 32)$$ [since, 100 parts on celsius scale = 180 parts on fahrenheit scale]

$$°F = \frac{9}{5} \times (°C) + 32$$

Pressure: There are three non-SI units for pressure which are commonly used.

(a) Atmosphere (atm) is defined as the pressure exerted by a column of mercury of 760 mm or 76 cm height at 0°C.
(b) Torr is defined as the pressure exerted by a 1 mm column of mercury at 0°C.
(c) Millimetre of mercury (mm Hg).

These three units are related as:

$$1\,\text{atm} = 760\,\text{torr} = 760\,\text{mm Hg} = 76\,\text{cm Hg} = 1.013 \times 10^5\ \textbf{Pa}$$

Energy: Calorie has been used in the past as a unit of energy measurement. The calorie was defined as the amount of heat required to raise the temperature of one gram of water from 14.5°C to 15.5 °C. One calorie is defined as exactly equal to 4.184 joules.

$$1\,\text{cal} = 4.184\ \text{J} \quad \text{or} \quad 1\,\text{J} = 0.2390\,\text{cal}$$

$$1\,\text{kcal} = 1000\,\text{cal} = 4.184\ \text{kJ}$$

Conversion factors

$$1\,\text{angstrom (Å)} = 10^{-8}\ \text{cm} = 10^{-10}\ \text{m} = 10^{-1}\ \text{nm} = 10^{2}\ \text{pm}$$

$$1\ \text{inch} = 2.54\,\text{cm} \quad \text{or} \quad 1\,\text{cm} = 0.394\ \text{inch}$$

$$39.37\ \text{inch} = 1\,\text{metre} \qquad 1\,\text{km} = 0.621\,\text{mile}$$

$$1\,\text{kg} = 2.20\ \text{pounds (lb)} \qquad 1\,\text{g} = 0.0353\,\text{ounce (o)}$$

$$1\ \text{pound (lb)} = 453.6\,\text{g}$$

$$\begin{aligned} 1\ \text{atomic mass unit (amu)} &= 1.6605 \times 10^{-24}\,\text{g} \\ &= 1.6605 \times 10^{-27}\,\text{kg} \\ &= 1.492 \times 10^{-3}\ \text{erg} = 1.492 \times 10^{-10}\ \text{J} \\ &= 3.564 \times 10^{11}\ \text{cal} = 9.310 \times 10^{8}\ \text{eV} \\ &= 931.48\ \text{MeV} \end{aligned}$$

$$\begin{aligned} 1\,\text{atmosphere (atm)} &= 760\ \text{torr} = 760\ \text{mm Hg} = 76\,\text{cm Hg} \\ &= 1.01325 \times 10^{5}\ \text{Pa} \end{aligned}$$

$$\begin{aligned} 1\,\text{calorie (cal)} &= 4.1840 \times 10^{7}\ \text{erg} = 4.184\ \text{J} \\ &= 2.613 \times 10^{19}\ \text{eV} \end{aligned}$$

$$1\,\text{coulomb (coul)} = 2.9979 \times 10^{9}\ \text{esu}$$

$$1\,\text{curie (Ci)} = 3.7 \times 10^{10}\ \text{disintegrations sec}^{-1}$$

$$\begin{aligned} 1\,\text{electron volt (eV)} &= 1.6021 \times 10^{-12}\,\text{erg} = 1.6021 \times 10^{-19}\ \text{J} \\ &= 3.827 \times 10^{-20}\,\text{cal} \\ &= 23.06\ \text{kcal mol}^{-1} \end{aligned}$$

$$1\,\text{erg} = 10^{-7}\ \text{J} = 2.389 \times 10^{-8}\,\text{cal} = 6.242 \times 10^{11}\,\text{eV}$$

$$1\,\text{electrostatic unit (esu)} = 3.33564 \times 10^{-10}\ \text{coul}$$

$$1\ \text{faraday}\ (F) = 9.6487 \times 10^{4}\ \text{coul}$$

$$1\,\text{dyne (dyne)} = 10^{-5}\ \text{N}$$

$$1\ \text{joule} = 10^{7}\ \text{erg} = 0.2390\,\text{cal}$$

$$\begin{aligned} 1\ \text{litre} &= 1000\,\text{cc} = 1000\ \text{mL} = 1\,\text{dm}^3 \\ &= 10^{-3}\ \text{m}^3 \end{aligned}$$

Values of Some Useful Constants

Fundamental constant	Value in old units	Value in SI units
Avogadro's number (N)	6.023×10^{23} mol^{-1}	6.023×10^{23} mol^{-1}
Atomic mass unit (amu)	1.6605×10^{-24} g	1.6605×10^{-27} g
Bohr radius (a_0)	0.52918 Å $= 0.52918 \times 10^{-8}$ cm	5.2918×10^{-11} m
Boltzmann constant (k)	1.3807×10^{-16} erg deg^{-1}	1.3807×10^{-23} JK^{-1}
Charge on electron (e)	(–) 4.8029×10^{-10} esu	(–) 1.6021×10^{-19} coul
Charge to mass ratio e/m of electron	1.7588×10^{8} coul g^{-1}	1.7588×10^{11} coul kg^{-1}
Electron rest mass (m_e)	9.1091×10^{-28} g	9.1091×10^{-31} kg
Gas constant (R)	0.0821 lit atm deg^{-1}mol^{-1} 8.314×10^{7} erg deg^{-1}mol^{-1} $1.987 \simeq 2.0$ cal deg^{-1}mol^{-1}	8.314 J K^{-1} mol^{-1}
Molar volume at NTP (V_m)	22.4 L mol^{-1}	22.4×10^{-3} m^3 mol^{-1}
Planck's constant (h)	6.6252×10^{-27} erg sec	6.6252×10^{-34} J sec
Proton mass (m_p)	1.6726×10^{-24} g	1.6726×10^{-27} kg
Neutron mass (m_n)	1.67495×10^{-24} g	1.67495×10^{-27} kg
Rydberg constant (R_z)	109678 cm^{-1}	1.09678×10^{7} m^{-1}
Velocity of light (c) in vacuum	2.9979×10^{10} cm sec^{-1} or 186281 miles sec^{-1}	2.9979×10^{8} m sec^{-1}
Faraday (F)	9.6487×10^{4} C / equiv. or 96500 C/equiv.	9.6487×10^{4} C / equiv.
$\dfrac{1}{4\pi\varepsilon_0}$	1	0.8988×10^{10} N m^{2}C^{-2} or 9×10^{9} N m^{2}C^{-2}

Derived SI Units

Quantity with Symbol	Unit (SI)	Symbol
Velocity (v)	metre per sec	m s^{-1}
Area (A)	square metre	m^2
Volume (V)	cubic metre	m^3
Density (ρ)	kilogram m^{-3}	kg m^{-3}
Acceleration (a)	metre per sec^2	m s^{-2}
Energy (E)	joule (J)	kg m^2 s^{-2}
Force (F)	newton (N)	kg m s^{-2}

Power (W)	watt (W)	$J\ s^{-1}$; $kg\ m^2\ s^{-3}$
Pressure (P)	pascal (Pa)	$N\ m^{-2}$
Resistance (R)	ohm (Ω)	$V\ A^{-1}$
Conduction (C)	ohm^{-1}, mho, siemens	$m^{-2}\ kg^{-1}\ s^3\ A^2$ or Ω^{-1}
Potential difference	volt (V)	$kg\ m^2\ s^{-3}\ A^{-1}$
Electrical charge	coulomb (C)	A-s (ampere-second)
Frequency (ν)	hertz (Hz)	cycle per sec
Magnetic flux × density	tesla (T)	$kg\ s^{-2}\ A^{-1} = N\ A^{-1}\ m^{-1}$

Popular Units and their SI Equivalents

Physical quantity	Unit with symbol	Equivalent in SI unit
Mass	1 amu	1 amu = 1.6605×10^{-27} kg
Energy	1 electron volt (eV)	1.602×10^{-19} joule
Length	1 Å	10^{-10} m (10^{-1} nm)
Volume	litre	$10^{-3}\ m^3 = dm^3$
Force	dyne	10^{-5} N
Pressure	1 atmosphere	760 torr (760 mm Hg)
		101325 pa or 10^5 pa
	1 bar	101325 pa or 10^5 pa
	1 torr	133.322 $N\ m^{-2}$
Dipole moment	debye, 10^{-18} esu-cm	3.324×10^{-30} cm
Magnetic flux density	gauss (G)	10^{-4} T
Area of nuclear cross section	1 barn	$10^{-28}\ m^2$
Nuclear Diameter	1 fermi (1 femto)	10^{-15} m

Significant Figures

There is always some degree of uncertainty in every scientific measurement except in counting. The uncertainty in measurement mainly depends upon two factors:

(i) Skill and accuracy of the observer,

(ii) Limitation of the measuring scale.

To indicate the precision of a measurement, scientists use the term **significant figures.** The significant figures in a number are all certain digits plus one doubtful digit. The number of significant figures gives the information that except the digit at extreme right, all other digits are precise or reproducible. For example, mass of an object is 11.24 g. This value indicates that actual mass of the object lies between 11.23 g and 11.25 g. Thus, one is sure of first three figures (1, 1 and 2) but the fourth figure is somewhat inexact. The total significant figures in this number are four.

The following rules are observed in counting the number of significant figures in a given measured quantity:

(i) All non-zero digits are significant. For example,

42.3 has three significant figures.

243.4 has four significant figures.

24.123 has five significant figures.

(ii) A zero becomes significant figure if it appears between two non-zero digits. For example,

5.03 has three significant figures.

5.604 has four significant figures.

4.004 has four significant figures.

(iii) Leading zeros or the zeros placed to the left of the number are never significant. For example,

0.543 has three significant figures.

0.045 has two significant figures.

0.006 has one significant figure.

(iv) Trailing zeros or the zeros placed to the right of the number are significant. For example,

433.0 has four significant figures.

433.00 has five significant figures.

343.000 has six significant figures.

(v) In exponential notation, the numerical portion gives the number of significant figures. For example,

1.32×10^{-2} has three significant figures.

1.32×10^{4} has three significant figures.

(vi) The non-significant figures in the measurements are rounded off.

(a) If the figure following the last number to be retained is less than 5, all the unwanted figures are discarded and the last number is left unchanged, *e.g.*,

5.6724 is 5.67 to three significant figures.

(b) If the figure following the last number to be retained is greater than 5, the last figure to be retained is increased by 1 unit and the unwanted figures are discarded, *e.g.*,

8.6526 is 8.653 to four significant figures.

(c) If the figure following the last number to be retained is 5, the last figure is increased by 1 only in case it happens to be odd. In case of even number the last figure remains unchanged.

2.3524 is 2.4 to two significant figures.

7.4511 is 7.4 to two significant figures.

Calculations Involving Significant Figures

In most of the experiments, the observations of various measurements are to be combined mathematically, *i.e.*, added, subtracted, multiplied or divided as to achieve the final result. Since, all the observations in measurements do not have the same precision, it is natural that the final result cannot be more precise than the least precise measurement. The following two rules should be followed to obtain the proper number of significant figures in any calculation.

Rule 1: The result of an addition or subtraction in the numbers having different precisions should be reported to the same number of decimal places as are present in the number having the least number of decimal places. The rule is illustrated by the following examples:

(a)

$$\begin{array}{rl} 33.3 & \leftarrow \text{(has only one decimal place)} \\ 3.11 & \\ 0.313 & \\ \hline \text{Sum} \quad 36.723 & \leftarrow \text{(answer should be reported to one decimal place)} \end{array}$$

Correct answer = 36.7

(b)

$$\begin{array}{rl} 3.1421 & \\ 0.241 & \\ 0.09 & \leftarrow \text{(has 2 decimal places)} \\ \hline \text{Sum} \quad 3.4731 & \leftarrow \text{(answer should be reported to 2 decimal places)} \end{array}$$

Correct answer = 3.47

(c)

$$\begin{array}{rl} 62.831 & \leftarrow \text{(has 3 decimal places)} \\ -\,24.5492 & \\ \hline \text{Difference} \quad 38.2818 & \leftarrow \text{(answer should be reported to 3 decimal places after rounding off)} \end{array}$$

Correct answer = 38.282

Rule 2: The answer to a multiplication or division is rounded off to the same number of significant figures as is possessed by the least precise term used in the calculation. Examples are:

(a)

$$\begin{array}{rl} 142.06 & \\ \times\, 0.23 & \leftarrow \text{(two significant figures)} \\ \hline 32.6738 & \leftarrow \text{(answer should have two significant figures)} \end{array}$$

Correct answer = 33

(b)

$$\begin{array}{rl} 51.028 & \\ \times\, 1.31 & \leftarrow \text{(three significant figures)} \\ \hline 66.84668 & \end{array}$$

Correct answer = 66.8

(c) $\dfrac{0.90}{4.26} = 0.2112676$

Correct answer = 0.21

Note: (i) Same procedure is followed if an expression involves multiplication as well as division.

(ii) The presence of exact numbers in an expression does not affect the number of significant figures in the answer.

Examples are:

(a) $\dfrac{3.24 \times 0.0866}{5.046} = 0.055643$ (b) $\dfrac{4.28 \times 0.146 \times 3}{0.0418} = 44.84784$

Correct answer = 0.0556 Correct answer = 44.8

MISCELLANEOUS NUMERICAL EXAMPLES

Example 1. *0.44 g of a hydrocarbon on complete combustion with oxygen gave 1.8 g water and 0.88 g carbon dioxide. Show that these results are in accordance with the law of conservation of mass.*

Solution: A hydrocarbon is a compound which consists of carbon and hydrogen only. It undergoes combustion forming carbon dioxide and water as products.

Formula of carbon dioxide = CO_2;

Molecular mass = $12 + 32 = 44$ g

Formula of water = H_2O;

Molecular mass = $2 + 16 = 18$ g

$$\text{Mass of carbon in 0.88 g of } CO_2 = \frac{12}{44} \times 0.88 = 0.24 \text{ g}$$

$$\text{Mass of hydrogen in 1.8 g of } H_2O = \frac{2}{18} \times 1.8 = 0.20 \text{ g}$$

Total masses of carbon and hydrogen in the products

$= 0.24 + 0.20 = 0.44$ g

This is equal to the mass of hydrocarbon before combustion.

Thus, the results are in accordance with the law of conservation of mass.

Example 2. *Calcium carbonate decomposes completely, on heating, into lime (CaO) and carbon dioxide (CO_2). 1 kg of calcium carbonate is completely decomposed by heat, when 560 g of lime are obtained. How much quantity of carbon dioxide in grams, moles and litres at NTP is produced in the process?*

Solution: According to law of conservation of mass,

Mass of lime + Mass of carbon dioxide = Mass of calcium carbonate

$$560 \text{ g} + \text{Mass of } CO_2 = 1000 \text{ g}$$

$$\text{Mass of } CO_2 = 1000 - 560 = 440 \text{ g}$$

$$\text{Molecular mass of } CO_2 = 12 + 32 = 44 \text{ g (1 mole)}$$

$$\text{No. of moles in 440 g of } CO_2 = \frac{440}{44} = 10$$

1 mole of CO_2 occupies volume at NTP = 22.4 litre

10 moles of CO_2 will occupy volume at NTP

$= 22.4 \times 10 = 224$ litre

Example 3. *10 mL of hydrogen combine with 5 mL of oxygen to yield water. When 200 mL of hydrogen at NTP are passed over heated CuO, the CuO loses 0.144 g of its mass. Do these results correspond to the law of constant proportions?*

Solution: 1st Case:

$$\text{Mass of 10 mL hydrogen at NTP} = \frac{2}{22400} \times 10 = 0.00089 \text{ g}$$

$$\text{Mass of 5 mL of oxygen at NTP} = \frac{32}{22400} \times 5 = 0.00714 \text{ g}$$

$$\text{Mass of 22400 mL of Hg vapour at NTP} = \frac{8.923}{1000} \times 22400$$
$$= 199.87 \text{ g}$$

Hence, molecular mass of Hg $= 199.87$ g

(c) $\text{Approximate atomic mass} = \frac{6.4}{\text{Sp. heat}} = \frac{6.4}{0.033} \doteq 193.93 \text{ g}$

$$\text{Valency of Hg} = \frac{193.93}{100} \simeq 2 \text{ (nearest whole number)}$$

So, accurate atomic mass = Eq. mass × Valency

$$= 100 \times 2 = 200 \text{ g}$$

$$\text{Atomicity} = \frac{\text{Mol. mass}}{\text{At. mass}} = \frac{199.88}{200} \simeq 1$$

Hence, mercury molecules are monoatomic.

Example 23. *How many grams of CaO are required to neutralise* $852\,g$ *of* P_4O_{10}? **(IIT 2005)**

Solution: The reaction will be:

$$6CaO + P_4O_{10} \longrightarrow 2Ca_3(PO_4)_2$$

$$852 \text{ g } P_4O_{10} \equiv 3 \text{ mol } P_4O_{10}$$

1 mole of P_4O_{10} neutralises 6 moles of CaO.

$\therefore$ 3 moles of P_4O_{10} will neutralise 18 moles of CaO.

$\therefore$ Mass of CaO $= 18 \times 56 = 1008$ g

Example 24. *If 1 grain is equal to 64.8 mg, how many moles of aspirin (mol. wt. = 169) are present in a 5 grain aspirin tablet?*

Solution: Mass of aspirin in the tablet $= 64.8 \times 5 = 324$ mg

$$= 0.324 \text{ g}$$

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{0.324}{169}$$

$$= 1.92 \times 10^{-3}$$

Example 25. *If the volume occupied in a crystal by a molecule of NaCl is* 47×10^{-24} *mL, calculate the volume of the crystal weighing* 1 g.

Solution: Number of molecules of NaCl

$$= \frac{\text{Mass}}{\text{Molar mass}} \times 6.023 \times 10^{23}$$

$$= \frac{1}{58.5} \times 6.023 \times 10^{23} = 1.03 \times 10^{22}$$

Volume of crystal $= 1.03 \times 10^{22} \times 47 \times 10^{-24} = 0.484$ mL

Example 26. *A plant virus is found to consist of uniform cylindrical particles of* 150 Å *in diameter and* 5000 Å *long. The specific volume of the virus is* $0.75\,cm^3/g$. *If the virus is considered to be a single particle, find its molecular mass.* **(IIT 1999)**

Solution: Volume of cylindrical virus $= \pi r^2 l$

$$= 3.14 \times \left(\frac{150}{2} \times 10^{-8}\right)^2 \times 5000 \times 10^{-8}$$

$$= 0.884 \times 10^{-16} \text{ cm}^3$$

$$\text{Mass of virus} = \frac{\text{Volume}}{\text{Specific volume}} = \frac{0.884 \times 10^{-16}}{0.75}$$

$$= 1.178 \times 10^{-16} \text{ g}$$

$$\text{Molar mass of virus} = \text{Mass of single virus} \times 6.023 \times 10^{23}$$

$$= 1.178 \times 10^{-16} \times 6.023 \times 10^{23}$$

$$= 7.095 \times 10^{7}$$

Example 27. *Weighing* 3104 *carats* (1 *carat* = 200 *mg*), *the Cullinan diamond was the largest natural diamond ever found. How many carbon atoms were present in the stone?*

Solution: Mass of the stone

$$= 3104 \times 200 = 620800 \text{ mg} = 620.8 \text{ g}$$

Number of atoms of carbon

$$= \frac{\text{Mass in gram}}{\text{Gram-atomic mass}} \times 6.023 \times 10^{23}$$

$$= \frac{620.8}{12} \times 6.023 \times 10^{23} = 3.12 \times 10^{25}$$

Example 28. *A cylinder of compressed gas contains nitrogen and oxygen in the ratio* 3 : 1 *by mole. If the cylinder is known to contain* 2.5×10^4 *g of oxygen, what is the total mass of the gas mixture?*

Solution: Number of moles of oxygen in the cylinder

$$= \frac{\text{Mass in gram}}{\text{Molecular mass in gram}} = \frac{2.5 \times 10^4}{32}$$

$$= 781.25$$

$\therefore$ Number of moles of $N_2 = 3 \times 781.25 = 2343.75$

$$\text{Mass of nitrogen in the cylinder} = 2343.75 \times 28$$
$$= 65625 \text{ g}$$
$$= 6.5625 \times 10^4 \text{ g}$$

Total mass of the gas in the cylinder

$$= 2.5 \times 10^4 + 6.5625 \times 10^4 = 9.0625 \times 10^4 \text{ g}$$

Example 29. *Atmospheric air has* 78% N_2; 21% O_2; 0.9 % *Ar and* 0.1 % CO_2 *by volume. What is the molecular mass of air in the atmosphere?*

Solution: Molecular mass of mixture

$$= \frac{\Sigma \text{ \% of each}}{100} \times \text{Molar mass}$$

$$= \frac{78}{100} \times 28 + \frac{21}{100} \times 32 + \frac{0.9}{100} \times 40 + \frac{0.1}{100} \times 44 = 28.964$$

Example 30. *The famous toothpaste Forhans contains* 0.76 *g of sodium per gram of sodium monofluoroorthophosphate* (Na_3PO_4F) *in* 100 *mL.*

(*a*) *How many fluorine atoms are present?*

(*b*) *How much fluorine in milligrams is present?*

Solution:

Molar mass of $Na_3PO_4F = 3 \times 23 + 31 + 16 \times 4 + 19 = 183$

183 g Na_3PO_4F contains = 19 g fluorine

$\therefore$ 0.76 g Na_3PO_4F contains $= \frac{19}{183} \times 0.76$ g fluorine

$= 0.0789$ g $= 78.9$ mg fluorine

Number of fluorine atoms

$$= \frac{\text{Mass in gram}}{\text{Gram-atomic mass}} \times 6.023 \times 10^{23}$$

$$= \frac{0.0789}{19} \times 6.023 \times 10^{23}$$

$$= 2.5 \times 10^{21} \text{ atoms}$$

Example 31. *An alloy of iron (54.7%), nickel (45%) and manganese (0.3%) has a density of 8.17 g/cm^3. How many iron atoms are there in a block of alloy measuring $10\,cm \times 20\,cm \times 15\,cm$?*

Solution:

Volume of the block of alloy $= 10 \times 20 \times 15\,\text{cm}^3$

$= 3000\,\text{cm}^3$

Mass of the block $= 3000 \times 8.17\,\text{g} = 24510\,\text{g}$

Mass of iron in the block $= \frac{54.7}{100} \times 24510 = 13406.97$ g

Number of iron atoms in the block $= \frac{\text{Mass}}{\text{Atomic mass}} \times 6.023 \times 10^{23}$

$$= \frac{13406.97}{56} \times 6.023 \times 10^{23}$$

$$= 1.442 \times 10^{26}$$

Example 32. *An analysis of pyrex glass showed 12.9% B_2O_3, 2.2% Al_2O_3, 3.8% Na_2O, 0.4% K_2O and remaining is SiO_2. What is the ratio of silicon to boron atoms in the glass?* **(BCECE 2007)**

Solution:

Percentage composition of $B_2O_3 = 12.9\%$

Percentage composition of

$SiO_2 = 100 - [12.9 + 2.2 + 3.8 + 0.4]$

$= 80.7\%$

Number of moles of $B_2O_3 = \frac{\text{Mass}}{\text{Molar mass}} = \frac{12.9}{70} = 0.184$

Number of moles of boron atoms $= 2 \times 0.184$

Number of moles of $SiO_2 = \frac{\text{Mass}}{\text{Molar mass}} = \frac{80.7}{60} = 1.345$

Number of moles of silicon atoms $= 1.345$

$$\frac{\text{Number of atoms of silicon}}{\text{Number of atoms of boron}} = \frac{N_A \times 1.345}{N_A \times 0.184} = \frac{7.3}{1}$$

Where, N_A = Avogadro's number

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

22. x gram of $CaCO_3$ was completely burnt in air. The mass of the solid residue formed is 28 g. What is the value of 'x' in gram? **(EAMCET 2005)**

(a) 44 (b) 200 (c) 150 (d) 50

[Ans. (d)]

[Hint: $\underset{100\,g}{CaCO_3(s)} \longrightarrow \underset{56\,g}{CaO(s)} + CO_2(g)$

56 g residue $\equiv$ 100 g $CaCO_3$

$\therefore$ 28 g residue $\equiv$ 50 g $CaCO_3$]

23. The mass of carbon anode consumed (giving only carbon dioxide) in the production of 270 kg of Al metal from bauxite by Hall process is:

(a) 270 kg (b) 540 kg (c) 90 kg (d) 180 kg

[Ans. (c)]

[Hint: $\underset{3 \times 12g}{3C} + 2Al_2O_3 \longrightarrow \underset{4 \times 27 = 108\,g}{4Al} + 3CO_2$

$\because$ 108 g Al is produced by consuming = 36 g carbon

$\therefore 270 \times 10^3$ g Al will be produced by consuming

$= \frac{36}{108} \times 270 \times 10^3$ g carbon

$= 90 \times 10^3$ g $= 90$ kg carbon]

24. The equivalent mass of an element is 4. Its chloride has vapour density 59.25. Then the valency of the element is:

(a) 4 (b) 3 (c) 2 (d) 1

[Ans. (b)]

[Hint: Molecular mass of $MCl_n = 59.25 \times 2 = 118.5$

$a + 35.5 \times n = 118.5$ ... (i)

Equivalent mass $\times n + 35.5 \times n = 118.5$

$4n + 35.5n = 118.5$ ... (ii)

$\therefore$ $n = 3$]

25. Sulphur trioxide is prepared by the following two reactions:

$$S_8(s) + 8O_2(g) \longrightarrow 8SO_2(g)$$

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

How many grams of SO_3 are produced from 1 mole S_8?

(a) 1280 (b) 640

(c) 960 (d) 320

[Ans. (b)]

[Hint: From the given reaction, it is clear that, 1 mole S_8 will give 8 moles of SO_3.

$\therefore$ Mass of SO_3 formed will be $= 80 \times 8 = 640$ g.]

26. Calculate the number of millilitres at STP of H_2S gas needed to precipitate cupric sulphide completely from 100 mL of a solution containing 0.75 g of $CuCl_2$ in 1 L.

(a) 21.4 (b) 14.2

(c) 41.2 (d) 124

[Ans. (d)]

[**Hint:** $CuCl_2 + H_2S \longrightarrow CuS + 2HCl$

Number of moles of H_2S = Number of moles of $CuCl_2$

$$= \frac{0.75}{134.5} = 0.00557$$

Volume of $H_2S = 0.00557 \times 22400 = 124.8$ mL]

27. In the reaction,

$$As_2S_5 + xHNO_3 \longrightarrow 5H_2SO_4 + yNO_2 + 2H_3AsO_4 + 12H_2O$$

the values of x and y are: **[JEE (Orissa) 2006]**

(a) 40, 40 (b) 10, 10

(c) 30, 30 (d) 20, 20

[**Ans.** (a)]

[**Hint:** In RHS, there are 40 hydrogen atoms, hence only option (a) will be suitable.]

SUMMARY AND IMPORTANT POINTS TO REMEMBER

1. **Chemistry:** Branch of physical science which deals with the properties, composition and changes of matter. It has several branches. Main branches are (i) organic (ii) inorganic (iii) physical and (iv) analytical. It is wide in its scope and touches almost every aspect of our lives.
2. **Matter:** It is anything which has mass and occupies space. Matter exists in three physical states (i) solid (ii) liquid and (iii) gas. It is chemically classified into (a) elements (b) compounds and (c) mixtures.
3. **Energy:** The capacity of doing work. It is of various forms. One form can be converted into another but cannot be created or destroyed. The total amount of matter and energy available in the universe is constant. The relationship between mass and energy is given by Einstein equation, $E = mc^2$ (where, E = energy, m = mass, c = velocity of light).
4. **Intensive properties:** Do not depend on the quantity of matter, *e.g.*, colour, density, melting point, boiling point, etc.
5. **Extensive properties:** Depend on the quantity of matter, *e.g.*, volume, mass, weight, etc.
6. **Substance:** A variety of matter, all samples of which have the same composition and properties. Pure substances are divided into (i) elements and (ii) compounds.
7. **Element:** A substance which cannot be decomposed into anything more simpler by ordinary physical or chemical means. 117 elements are known. 88 elements have been isolated from natural sources and remaining 29 have been prepared by artificial means. Every element is represented by a **symbol** which is a small abbreviation of its full and lengthy name. Oxygen is the most abundant element. Silicon, aluminium, iron are second, third and fourth most abundant elements. Elements are classified as (i) metals (ii) non-metals and (iii) metalloids.
8. **Metals:** Generally solids (Hg—exception). They have properties such as lustre, hardness, malleable, ductile, good conductors of heat and electricity. Copper, zinc, iron, aluminium are metals.
9. **Non-metals:** Usually non-lustrous, brittle and poor conductors of electricity. Oxygen, carbon, nitrogen, chlorine, helium, etc., are non-metals.
10. **Metalloids:** Possess mixed properties of metals and non-metals both (*e.g.*, As, Sb, Sn).
11. **Compound:** Pure substance composed of two or more different elements in a fixed proportion of mass. The properties of a compound are altogether different from the properties of elements from which it has been constituted.
12. **Mixture:** A material containing two or more substances (elements or compounds) in any proportion, in which components do not lose their identity. Homogeneous mixture has a single phase while heterogeneous has more than one phase. Mixture can be separated into components by physical methods.
13. **Alloy:** A homogeneous mixture of two or more elements—metal and metal, metal and non-metal or non-metal and non-metal. They have unique properties.
14. **Physical change:** A temporary change, no change in chemical composition and mass. Physical properties alter. It can be reversed easily.
15. **Chemical change:** A permanent change, new substance is formed which possesses different composition and properties. It cannot be reversed easily. Chemical changes are of various types. The important ones are decomposition, synthesis, substitution, addition, internal rearrangement, polymerisation, double decomposition, etc.
16. **Law of conservation of mass:** (Lavoisier—1774) In a chemical change, mass is neither created nor destroyed. In chemical reactions:

    Total masses of reactants = Total masses of products.
17. **Law of constant proportions:** (Proust—1799) A chemical compound always contains the same element combined together in fixed proportion by mass.
18. **Law of multiple proportions:** (Dalton—1808) When two elements combine to form two or more compounds, the different masses of one element which combine with a fixed mass of the other element, bear a simple ratio to one another.
19. **Law of reciprocal proportions:** (Richter—1794) When two different elements combine with the same mass of a third element, the ratio in which they do so will be the same or simple multiple if both directly combine with each other. In all chemical reactions, substances react in the ratio of their equivalent masses.
20. **Law of gaseous volumes:** (Gay-Lussac—1808) Gases react with each other in simple ratio of their volumes and if product is also in gaseous state, its volume also bears a simple ratio with the volumes of gaseous reactants under similar conditions of temperature and pressure.

21. **Dalton's atomic theory:** Every element is composed of small indivisible, indestructible particles called atoms. Atoms of the same element are identical but differ in properties, mass and size of atoms of other elements. Atoms of different elements combine in simple ratio to form compounds. The relative number and kind of atoms are always the same in a given compound. Atoms cannot be created or destroyed.
22. **Atom:** The smallest particle of an element that takes part in a chemical reaction.
23. **Molecule:** The smallest particle of an element or compound that can have a stable existence.
24. **Formula:** Group of symbols of elements which represents one molecule of a substance. It represents also the chemical composition.
25. **Atomic mass:** Atomic mass of an element is the ratio of mass of one atom of an element to $\frac{1}{12}$th part of the mass of carbon-12.

Atomic mass of an element

$$= \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of carbon-12}} \times 12$$

26. **Atomic mass unit (amu):** $\frac{1}{12}$th mass of carbon-12. It is equal to 1.66×10^{-24} g.

Atomic mass of an element

$$= \frac{\text{Mass of one atom of the element}}{1\,\text{amu}}$$

The actual mass of an atom of element = Atomic mass in amu $\times 1.66 \times 10^{-24}$ g.

The atomic masses of elements are actually average relative masses because elements occur as mixture of isotopes.

27. **Gram-atomic mass or Gram atom:** Atomic mass expressed in grams. It is the absolute mass in grams of 6.02×10^{23} atoms of any element.

$$\text{No. of gram atoms} = \frac{\text{Mass of element in grams}}{\text{Atomic mass of the element in grams}}$$

28. **Molecular mass:** It indicates how many times one molecule of a substance is heavier in comparison to $\frac{1}{12}$th of mass of one atom of carbon-12. Mass of a molecule is equal to sum of masses of the atoms present in a molecule.
29. **Gram-molecular mass or Gram molecule:** Molecular mass expressed in gram. It is the absolute mass in gram of 6.02×10^{23} molecules of any substance.

No. of gram molecules

$$= \frac{\text{Mass of a substance in gram}}{\text{Molecular mass of the substance in gram}}$$

30. **Avogadro's hypothesis:** Under similar conditions of temperature and pressure, equal volumes of all gases contain same number of molecules.
31. **Gram molar volume:** The volume occupied by one gram-molecular mass of any gas at NTP (0°C or 273 K and one atm or 76 cm of Hg as pressure). Its value is 22.4 litre.
32. **Vapour density:**

$$\text{V.D.} = \frac{\text{Density of a gas}}{\text{Density of hydrogen}}$$

$$= \frac{\text{Mass of a certain volume of a gas}}{\text{Mass of same volume of hydrogen under same temperature and pressure}}$$

$$2\,\text{V.D.} = \text{Molecular mass}$$

33. **Mole:** A mole (mol) is defined as the number of atoms in 12.0 g of carbon-12. The number of atoms is 6.02×10^{23}. This number is called Avogadro's number.

$$\text{No. of moles} = \frac{\text{Mass of substance in gram}}{\text{Mass of one mole of the substance in gram}}$$

$$= \frac{\text{No. of particles}}{6.02 \times 10^{23}}$$

$$= \frac{\text{Volume of gas in litres at NTP}}{22.4}$$

Mass of one atom of an element

$$= \frac{\text{Gram atom of an element}}{6.02 \times 10^{23}}$$

Mass of one molecule of a substance

$$= \frac{\text{Gram-molecular mass of a substance}}{6.02 \times 10^{23}}$$

34. **Equivalent mass:** The number of parts by mass of the substance which combine or displace directly or indirectly 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of chlorine or 108 parts by mass of silver.

The equivalent mass of an element may vary with change of valency.

Eq. mass of an element

$$= \frac{\text{Mass of element}}{\text{Mass of hydrogen}} \times 1.008$$

$$= \frac{\text{Mass of element} \times 11200}{\text{Volume in mL of hydrogen displaced at NTP}}$$

$$= \frac{\text{Mass of element}}{\text{Mass of oxygen}} \times 8$$

$$= \frac{\text{Mass of element}}{\text{Mass of chlorine}} \times 35.5$$

35. **Metal to metal displacement:** $\frac{m_1}{m_2} = \frac{E_1}{E_2}$
36. **Double decomposition:** $AB + CD \rightarrow \underset{\text{ppt.}}{AD} + CB$

$$\frac{\text{Mass of } AB}{\text{Mass of } AD} = \frac{\text{Eq. mass of } A + \text{Eq. mass of } B}{\text{Eq. mass of } A + \text{Eq. mass of } D}$$

Atomic mass of an element

$$= \text{Eq. mass of the element} \times \text{Valency}$$

37. Dulong and Petit's law:

$$\text{Atomic mass (approximate)} = \frac{6.4}{\text{Specific heat}}$$

38. Cannizzaro's method: Atomic mass of an element is the smallest mass of the element present in the molecular mass of any one of its compounds.

39. Law of isomorphism: Isomorphous compounds form crystals which have same size and shape and can grow in the saturated solution of each other.

Masses of two elements that combine with same mass of other elements in their respective compounds are in the ratio of their atomic masses.

40. Atomic mass from vapour density of a chloride:

$$\text{Valency of an element} = \frac{2\,\text{V.D. of a volatile chloride}}{\text{Eq. mass} + 35.5}$$

41. Types of formulae:

(i) **Empirical:** It represents the simplest relative whole number ratio of atoms of each element present in the molecule of a substance.

(ii) **Molecular:** It represents the actual number of atoms of each element present in one molecule of a substance.

Molecular formula = $n \times$ Empirical formula

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

(iii) **Structural:** It represents the way in which atoms of various elements are linked with each other.

42. Percentage of element:

$$\text{Percentage of element} = \frac{\text{Mass of element}}{\text{Molecular mass}} \times 100$$

43. Chemical equation: It is a symbolic representation of a chemical change. The equation becomes balanced when total number of atoms of various elements are made equal on both the sides of equation. Chemical equations are of two types (i) molecular and (ii) ionic. Chemical equation is based on law of conservation of mass.

44. Unit: It is the primary standard chosen to measure any physical quantity.

The seven units of measurement, namely mass, length, time, temperature, electric current, luminous intensity and amount of substance are taken as basic units. All other units can be derived from them and are, therefore, called derived units. SI units are used these days in all branches of science.

45. Significant figure: It is the total number of certain digits plus one doubtful digit.

Questions

1. Match the following, choosing one item from Column-X and the appropriate related item from Column-Y.

[A]

Column-X	Column-Y
(a) Efflorescence	(i) Homogeneous mixture
(b) Malleability	(ii) Heterogeneous mixture
(c) Alloy	(iii) Mole
(d) 1 amu	(iv) (1/12)th mass of carbon-12
(e) Sulphur and sand	(v) Tendency to lose water of crystallisation
(f) Amount of substance	(vi) Property of metal being hammered into thin sheets

[B]

Column-X	Column-Y
(a) Equal volumes of all gases contain equal number of molecules at NTP.	(i) Dalton's atomic theory
(b) The atom is indestructible.	(ii) Law of conservation of mass
(c) All pure samples of the same compound contain the same elements combined in the same proportion by mass.	(iii) Avogadro's law
(d) Total mass before and after the chemical reaction is same.	(iv) Dulong and Petit's law
(e) Atomic mass $= \dfrac{6.4}{\text{Specific heat}}$	(v) Gay-Lussac's law
(f) Gases react in simple ratio of their volumes.	(vi) Law of constant proportions

[C]

Column-X	Column-Y
(a) Most abundant element	(i) Platinum
(b) Most abundant metal	(ii) Diamond
(c) Liquid at room temp.	(iii) Aluminium
(d) Hardest substance	(iv) Plutonium
(e) Most ductile metal	(v) Mercury
(f) Transuranic element	(vi) Oxygen

2. **Matrix Matching Problems:**
(According to the new pattern of IIT Screening)

[A] Match the Column-X and Column-Y:

Column-X	Column-Y
(a) Vapour density	(i) Unitless
(b) Mole	(ii) 1 mol electrons
(c) 12 g carbon	(iii) Collection of 6.023×10^{23} atoms
(d) 96500 C	(iv) Molecular mass $\times \frac{1}{2}$

[B] Match the Column-X and Column-Y:

Column-X	Column-Y
(a) 1.6 g CH_4	(i) 0.1 mol
(b) 1.7 g NH_3	(ii) 6.023×10^{23} electrons
(c) HCHO	(iii) 40% carbon
(d) $C_6H_{12}O_6$	(iv) Vapour density = 15

[C] Match the Column-X and Column-Y:

Column-X	Column-Y
(a) 1 amu	(i) Heaviest particle of atom
(b) Proton	(ii) 1.66×10^{-27} kg
(c) Neutron	(iii) 931.5 MeV
(d) α-particle	(iv) Positively charged

[D] Match the Column-X with Column-Y for the reaction:

$$A + B_2 \rightarrow AB_2$$

Column-X	Column-Y
(a) 300 atoms of A + 200 molecules of B_2	(p) B_2 is limiting reagent
(b) 100 atoms of A + 100 molecules of B_2	(q) A is limiting reagent
(c) 5 mol of A + 2.5 mol of B_2	(r) None of the reactant is in excess
(d) 2.5 mol of A + 5 mol of B_2	(s) 200 molecules of AB_2 will be formed

Answers

1. [A] (a–v); (b–vi); (c–i); (d–iv); (e–ii); (f–iii)
[B] (a–iii); (b–i); (c–vi); (d–ii); (e–iv); (f–v)
[C] (a–vi); (b–iii); (c–v); (d–ii); (e–i); (f–iv).

2. [A] (a–i, iv); (b–iii); (c–iii); (d–ii)
[B] (a–i, ii); (b–i, ii); (c–iii, iv); (d–iii)
[C] (a–ii, iii); (b–ii, iv); (c–i); (d–iv)
[D] (a–p, s); (b–r); (c–p); (d-q).

• PRACTICE PROBLEMS •

1. The density of mercury is 13.6 g/mL. Calculate the diameter of an atom of mercury assuming that each atom of mercury is occupying a cube of edge-length equal to the diameter of mercury atom.
(Atomic mass of mercury = 200)
[**Ans.** 2.9×10^{-8} cm]

2. A metal M of atomic mass 54.94 has a density of 7.42 g/cc. Calculate the apparent volume occupied by one atom of the metal.
[**Ans.** 1.23×10^{-23} cc]

3. Find the charge of 1 g ion of N^{3-} in coulomb.
[**Ans.** 2.894×10^5 coulomb]

4. Calculate the volume at NTP occupied by 6.25 g of nitrogen.
[**Ans.** 5.0 litre]

5. 10 mL of hydrogen contains 2×10^3 molecules of hydrogen at certain pressure and temperature. Calculate the number of molecules of oxygen whose volume is 200 mL at the same temperature and pressure.
[**Ans.** 4×10^4 molecules]

6. The masses of equal volumes of a gas and hydrogen are 25.6 g and 0.8 g respectively under same conditions of temperature and pressure. Find the molecular mass of the gas.
[**Hint:** V.D. of the gas $= \frac{25.6}{0.8} = 32.0$
Molecular mass $= 2$ V.D. $= 2 \times 32.0 = 64.0$]

7. One litre of a gas at NTP weighs 1.97 g. Find the molecular mass of gas.
[**Ans.** 44.128]

8. How many moles of water are present in one litre of water?
[**Ans.** 55.5 moles]

9. Calculate the mass of 6.02×10^{21} molecules of nitrogen.
[**Ans.** 0.28 g]

10. 1.5276 g of $CdCl_2$ was found to contain 0.9367 g of cadmium. Calculate the atomic mass of cadmium.
[**Ans.** 112.54]
[**Hint:** Equivalent mass of cadmium $= \frac{\text{Mass of Cd}}{\text{Mass of Cl}} \times 35.5$
$= \frac{0.9367}{0.5909} \times 35.5 = 56.27$
Atomic mass = Equivalent mass × Valency]

11. Calculate how many methane molecules and how many hydrogen and carbon atoms are there in 25.0 g of methane? **(MLNR 1990; Dhanbad 1992)**
[**Ans.** 9.41×10^{23} CH_4 molecules, 9.41×10^{23} carbon atoms and 37.64×10^{23} hydrogen atoms.]
[**Hint:** No. of moles of methane $= \frac{25.0}{16.0}$
One molecule of methane contains one carbon atom and four hydrogen atoms.]

12. How much sugar ($C_{12}H_{22}O_{11}$) will be required if each person on the earth is given 100 molecules of sugar? The population of the earth is 3×10^{10}.
[**Ans.** 170.43×10^{-11} g]

13. A mixture of hydrogen and oxygen contains 20% by mass of hydrogen. What is the total number of molecules present per gram of the mixture?
[**Ans.** 7.528×10^{22}]
[**Hint:** In 1 gram of the mixture, 0.2 g of hydrogen and 0.8 g of oxygen are present. Moles of $H_2 = \frac{0.2}{2} = 0.1$, moles of oxygen $= \frac{0.8}{32} = 0.025$. Calculate the number of molecules of hydrogen and oxygen and then add.]

14. How many electrons are present in 18 mL of water? **(MLNR 1995)**
[**Hint:** 18 mL water = 18 g water = 1 mole water $= 6.02 \times 10^{23}$ molecules, each molecule consists 10 electrons (8 electrons per oxygen atom, 2 electrons for two hydrogen atoms). Total electrons $= 10 \times 6.02 \times 10^{23} = 6.02 \times 10^{24}$]

15. Sulphur molecule is known to be composed of 8 atoms of the element. In a sample of 192 g of pure sulphur, calculate (i) number of g-atoms of sulphur; (ii) number of atoms of sulphur; (iii) number of moles of sulphur; (iv) number of molecules of sulphur.
[**Ans.** g-atoms = 6; No. of atoms $= 6 \times 6.02 \times 10^{23}$; No. of moles = 0.75; No. of molecules $= 4.52 \times 10^{23}$]
[**Hint:** The atomic mass of sulphur is 32.]

16. The vapour density of a mixture containing NO_2 and N_2O_4 is 38.3 at 27°C. Calculate the moles of NO_2 in 100 g of the mixture. **(MLNR 1993)**
[**Hint:** Mol. mass of mixture $= 2 \times 38.3 = 76.6$
No. of moles in 100 g of mixture $= \frac{100}{76.6}$
Let a g of NO_2 is present in mixture.

Moles of NO_2 + Moles of N_2O_4 = Moles of mixture

$$\frac{a}{46}+\frac{100-a}{92}=\frac{100}{76.6} \quad \text{or} \quad a = 20.10 \text{ g}$$

Moles of NO_2 in mixture $=\frac{20.10}{46}=0.437$]

17. Calculate the number of oxygen atoms in 88 g CO_2. What would be the mass of CO having the same number of oxygen atoms? **(BITS 1990)**

[Hint: 88 g CO_2 = 2 moles of CO_2. One molecule consists of 2 oxygen atoms.

No. of oxygen atoms $= 2 \times 2 \times 6.02 \times 10^{23} = 24.08 \times 10^{23}$

CO molecule has one oxygen atom.

Mass of CO containing 24.08×10^{23} oxygen atoms

$$=\frac{28}{6.02 \times 10^{23}} \times 24.08 \times 10^{23} = 112 \text{ g }]$$

18. Density of water at room temperature is 1.0 g cm^{-3}. How many molecules are there in one drop of water if its volume is 0.1 cm^3?

[Ans. 3.34×10^{21} molecules]

[Hint: Mass of one drop = Vol. $\times d = 0.1 \times 1 = 0.1$ g

No. of moles $=\frac{0.1}{18}$; No. of molecules $= 6.02 \times 10^{23} \times \frac{0.1}{18}$]

19. Naturally occurring boron consists of two isotopes, whose atomic masses are 10.01 and 11.01. The atomic mass of natural boron is 10.81. Calculate the percentage of each isotope in natural boron. **(MLNR 1994)**

[Ans. % of isotope with atomic mass 10.01 = 20; % of isotope with atomic mass 11.01 = 80]

[Hint: Let x be the percentage of the isotope with atomic mass 10.01.

$$\frac{10.01 \times x}{100}+\frac{11.01(100-x)}{100}=10.81 \quad \text{or} \quad x = 20]$$

20. Chlorine has isotopes ^{35}Cl and ^{37}Cl. There are three ^{35}Cl isotopes for every ^{37}Cl isotope in a sample of chlorine. Calculate the atomic mass of chlorine.

[Ans. $A=\frac{3 \times 35 + 37 \times 1}{4}=35.5$]

21. Natural hydrogen gas is a mixture of ^1H and ^2H in the ratio of 5000 : 1. Calculate the atomic mass of the hydrogen.

[Ans. 1.000199]

22. Chromium has the following isotopic composition:

Mass number	Isotopic mass	Fractional abundance
50	49.9461	x
52	51.9405	0.8379
53	52.9407	0.0950
54	53.9389	0.0236

Calculate the value of x.

[Ans. 0.0435]

23. Use the data given in the following table to calculate the molar mass of naturally occurring argon:

Isotope	Isotopic molar mass	Abundance
^{36}Ar	35.96755 g mol^{-1}	0.337%
^{38}Ar	37.96272 g mol^{-1}	0.063%
^{40}Ar	39.9624 g mol^{-1}	99.6%

[Ans. 39.947]

24. Density of oxygen at NTP is 1.429 g/litre. Calculate the standard molar volume of the gas.

[Ans. 22.39 litre mol^{-1}]

25. How many iron atoms are present in a stainless steel ball bearing having a radius of 0.254 cm? The stainless steel contains 85.6% Fe by weight and has density of 7.75 g/cm^3.

[Ans. 4.91×10^{21}]

26. The nucleus of an atom X is supposed to be a sphere with a radius of 5×10^{-13} cm. Find the density of the matter in the atomic nucleus if the atomic weight of X is 19.

[Ans. 6.02×10^{13} g/mL]

27. Calculate the number of atoms of each element present in 122.5 g of $KClO_3$.

[Ans. Number of atoms of 'K' $= 1 \times 6.023 \times 10^{23}$

Number of atoms of 'Cl' $= 1 \times 6.023 \times 10^{23}$

Number of atoms of 'O' $= 3 \times 6.023 \times 10^{23}$]

28. In an experiment, 1.0 g $CaCO_3$ on heating evolved 224 mL of CO_2 at NTP. What mass of CaO (calcium oxide) is formed?

[Ans. Mass of CaO = 0.56 g]

[Hint: Mass of 224 mL of $CO_2 = \frac{44}{22400} \times 224 = 0.44$ g]

29. What mass of potassium chlorate ($KClO_3$) on heating gives 1.491 g of potassium chloride (KCl) and 0.672 litres of oxygen at NTP?

[Ans. Mass of $KClO_3$ = 2.451 g]

[Hint: Mass of 22.4 litre of oxygen at NTP = 32 g]

30. A compound AB completely decomposes into A and B on heating. 50 g of AB, on strong heating, gave 40 g of A. How much quantity of AB should be decomposed by heating to obtain 2.5 g of B? How much quantity of A will be produced in the process?

[Ans. 12.5 g AB is to be decomposed, 10.0 g of A will be produced.]

[Hint: $\underset{50\text{ g}}{AB} \rightarrow \underset{40\text{ g}}{A} + \underset{10\text{ g}}{B}$]

31. If 12.6 g of $NaHCO_3$ is added to 20.0 g of HCl solution, the residue solution is found to weigh 24.0 g. What is the mass and volume of CO_2 released at NTP in the reaction?

[Ans. 8.6 g CO_2 released. Volume at NTP $=\frac{22.4}{44} \times 8.6 = 4.378$ litre]

32. (i) 5.06 g of pure cupric oxide (CuO), on complete reduction by heating in a current of hydrogen, gave 4.04 g of metallic copper.

(ii) 1.3 g of pure metallic copper was completely dissolved in nitric acid and the resultant solution was carefully dried and ignited. 1.63 g CuO was produced in the process. Show that these results illustrate the law of constant proportions.

[**Ans.** In both cases, the ratio of copper and oxygen is 1 : 0.25. Hence, the law of constant proportions is illustrated.]

33. Metal M and chlorine combine in different proportions to form two compounds A and B. The mass ratio M : Cl is 0.895 : 1 in A and 1.791 : 1 in B. What law of chemical combination is illustrated?

[**Ans.** Masses of metal which combine with 1 part of chlorine are in the ratio of 1 : 2, which is a simple ratio. Hence, law of multiple proportions is illustrated.]

34. 2.8 g of calcium oxide (CaO) prepared by heating limestone were found to contain 0.8 g of oxygen. When one gram of oxygen was treated with calcium, 3.5 g of calcium oxide was obtained. Show that the results illustrate the law of definite proportions.

35. By means of the given analytical results show that law of multiple proportions is true:

Mercurous chloride	**Mercuric chloride**
Mercury = 84.92 %	Mercury = 73.80%
Chlorine = 15.08 %	Chlorine = 26.20%

[**Ans.** The masses of mercury which combine with 1 part of chlorine are in the ratio of 2 : 1, which is a simple ratio. Hence, law of multiple proportions is illustrated.]

36. 1 g of a metal, having no variable valency, produces 1.67 g of its oxide when heated in air. Its carbonate contains 28.57% of the metal. How much oxide will be obtained by heating 1 g of the carbonate?

[**Ans.** 0.477 g]

[**Hint:** $\dfrac{\text{Mass of metal}}{\text{Mass of oxygen}} = \dfrac{\text{Mass of metal in 1 g of carbonate}}{x}$

i.e., $x = 0.1914$ g of oxygen

Mass of oxide $= 0.2857 + 0.1914 = 0.4771$ g]

37. 0.36 g of Mg combines with chlorine to produce 1.425 g of magnesium chloride. 9.50 g of another sample of anhydrous magnesium chloride gave, on electrolysis 2.24 litre of chlorine at NTP. Show that these data agree with the law of constant proportions.

[**Hint:** Mass of 2.24 litre of chlorine at NTP $= \dfrac{71}{22.4} \times 2.24$ $= 7.1$ g. In both cases, the ratio of masses of Mg and Cl is 1 : 3. Hence, law of constant proportions is followed.]

38. Carbon dioxide contains 27.27% carbon, carbon disulphide contains 15.97% carbon and sulphur dioxide contains 50% sulphur. Show that these figures illustrate the law of reciprocal proportions.

[**Hint:** The masses of oxygen and sulphur which combine with 1 part of carbon are in the ratio of 2.667 : 5.25, *i.e.*, 1 : 2. In sulphur dioxide, the masses of sulphur and oxygen are in the ratio of 1 : 1 which is a simple multiple of first. Hence, law of reciprocal proportions is illustrated.]

39. Phosphorus and chlorine form two compounds. The first contains 22.54% by mass of phosphorus and the second 14.88% of phosphorus. Show that these data are consistent with law of multiple proportions.

[**Hint:** The ratio of the masses of chlorine which combines with a fixed mass of phosphorus in two compounds is 3 : 5 which is a simple whole number ratio. Thus, the data illustrate law of multiple proportions.]

40. A and B are two hydrocarbons. A and B are heated separately in excess of oxygen when 0.028 g of A gave 44.8 mL CO_2 and 0.044 g of B gave 67.2 mL CO_2 at NTP. Show that the results are in agreement with law of multiple proportions.

[**Hint:** Determine the masses of CO_2 at NTP and then masses of carbon.

(A) Mass of $CO_2 = \dfrac{44}{22400} \times 44.8 = 0.088$ g,

Mass of carbon = 0.024 g, mass of hydrogen = 0.004 g.

(B) Mass of $CO_2 = \dfrac{44}{22400} \times 67.2 = 0.132$ g,

Mass of carbon = 0.036 g, mass of hydrogen = 0.008 g.

Thus, the masses of carbon combining with same mass of hydrogen are in the ratio of 4 : 3 which is a simple ratio. Hence, law of multiple proportions is followed.]

41. Aluminium oxide contains 52.9% aluminium and carbon dioxide contains 27.27% carbon. Assuming the validity of the law of reciprocal proportions, calculate the percentage of aluminium in aluminium carbide.

[**Hint:** From the data, it is observed that the ratio of masses of aluminium and carbon in aluminium carbide should be 3 : 1 or its simple multiple. Hence, percentage of aluminium in aluminium carbide $= \dfrac{3}{4} \times 100 = 75.0$]

42. Two volumes of ammonia, on dissociation gave one volume of nitrogen and three volumes of hydrogen. How much hydrogen will be obtained from dissociation of 40 mL of NH_3?

[**Ans.** 60 mL]

43. The following results were obtained by heating different oxides of lead in a current of hydrogen:

(a) 1.393 g of litharge gave 1.293 g of lead.
(b) 2.173 g of lead peroxide gave 1.882 g of lead.
(c) 1.721 g of red lead gave 1.552 g of lead.

Show that these results are in accordance with the law of multiple proportions.

[**Ans.** Masses of lead that combine with same mass of oxygen are in the ratio of 4 : 2 : 3 which is a simple ratio. So, the results are in accordance with the law of multiple proportions.]

44. Calculate the number of g-moles of CaO that could be obtained from 42.54 g of $CaCO_3$ and convert the number of g-moles to grams.

[No. of g-moles $= \dfrac{42.54}{100} = 0.4254$,

Mass of CaO $= 0.4254 \times 56 = 23.8$ g]

45. 1 g of a metal M which has specific heat of 0.06 combines with oxygen to form 1.08 g of oxide. What is the atomic mass of M?

[**Hint:** Approximate atomic mass $= \frac{6.4}{0.06} = 106.6$

$$\text{Equivalent mass of } M = \frac{1}{0.08} \times 8 = 100$$

$$\text{Valency} = \frac{106.6}{100} \approx 1$$

$$\text{Exact atomic mass} = 100 \times 1 = 100]$$

46. A compound contains 28% of nitrogen and 72% metal by mass. 3 atoms of the metal combine with 2 atoms of the nitrogen. Find the atomic mass of the metal.

[**Hint:** Valency of metal = 2 and valency of nitrogen = 3

$$\text{Equivalent mass of nitrogen} = \frac{14}{3}; \frac{\text{Eq. mass of metal}}{14/3} = \frac{72}{28}$$

Equivalent mass of metal = 12

Atomic mass of metal = 12 × 2 = 24]

47. The chloride of a solid metallic element contains 57.89% by mass of the element. The specific heat of the element is 0.0324 cal deg^{-1} g^{-1}. Calculate the exact atomic mass of the element.

[**Hint:** Equivalent mass of the element $= \frac{57.89}{42.11} \times 35.5 = 48.8$

$$\text{Approximate atomic mass} = \frac{6.4}{0.0324} = 200$$

$$\text{Valency} = \frac{200}{48.8} \approx 4$$

$$\text{Exact atomic mass} = 48.8 \times 4 = 195.2\ ]$$

48. Two oxides of a metal contain 63.2% and 69.62% of the metal. The specific heat of the metal is 0.117. What are the formulae of the two oxides?

[**Ans.** MO_2 and M_2O_3]

49. White vitriol (hydrated zinc sulphate) is isomorphous with $MgSO_4 \cdot 7H_2O$. White vitriol contains 22.95% zinc and 43.9% of water of crystallisation. Find the atomic mass of zinc.

[**Hint:** The formula of white vitriol should be $ZnSO_4 \cdot 7H_2O$ as it is isomorphous to $MgSO_4 \cdot 7H_2O$, *i.e.*, 7 water molecules are associated with one zinc atom. $7H_2O = 7 \times 18 = 126$. Mass of Zn with which 126 parts of water by mass are associated $= \frac{22.95}{43.90} \times 126 = 65.87$. Atomic mass of zinc.]

50. A solid element burns in oxygen without any change in volume (of gas) under similar conditions of temperature and pressure. If the vapour density of pure gaseous product is 32, what is the equivalent mass of the element?

[**Hint:** One vol. of oxide contains 1 vol. of O_2.

One mole of oxide contains one mole of O_2.

$$\text{Mol. mass of oxide} = A + 32 = 2\ \text{V.D.} = 64$$

So, $A = 32$

32 parts of element combine with 32 parts of oxygen.

So, Equivalent mass of element $= \frac{32}{32} \times 8 = 8$]

51. If the equivalent mass of a metal (M) is x and the formula of its oxide is $M_m\ O_n$, then show that the atomic mass of M is $\frac{2xn}{m}$.

[**Hint:** m atoms of M combine with n atoms of oxygen.

1 atom of M combines with $\frac{n}{m}$ atoms of oxygen.

Hence, $\text{Valency} = \frac{2n}{m}$

$$\text{Atomic mass} = \text{Equivalent mass} \times \text{Valency}$$
$$= x \times \frac{2n}{m} = \frac{2nx}{m}\ ]$$

52. Two oxides of metals A and B are isomorphous. The metal A whose atomic mass is 52, forms a chloride whose vapour density is 79. The oxide of the metal B contains 47.1% oxygen. Calculate the atomic mass of B.

[**Hint:** Let the valency of A be x. The formula of chloride $= ACl_x$

$$2\ \text{V.D.} = A + x \times 35.5 \quad \text{or} \quad x \approx 3$$

As the two oxides are isomorphous, the valency of B is also 3. Equivalent mass of $B = \frac{52.9}{47.1} \times 8 = 8.99$, atomic mass of $B = 8.99 \times 3 = 26.97$]

53. A mixture of 1.65×10^{21} molecules of X and 1.85×10^{21} molecules of Y weighs 0.688 g. If molecular mass of Y is 187, what is the molecular mass of X ?

[**Hint:** $\frac{A \times 1.65 \times 10^{21}}{6.02 \times 10^{23}} + \frac{187 \times 1.85 \times 10^{21}}{6.02 \times 10^{23}} = 0.688, A = 41.35$]

54. The equivalent mass of a metal is 29.73 and the vapour density of its chloride is 130.4. Find out the atomic mass of the metal.

[**Ans.** Atomic mass = 118.92]

[**Hint:** Valency $= \frac{2 \times \text{V.D.}}{\text{Eq. mass} + 35.5} = \frac{2 \times (130.4)}{(29.73 + 35.5)} \approx 4$]

55. Calculate the percentage of aluminium, sulphate radical and water in potash alum.

[**Ans.** Al = 5.69%; $SO_4^{2-} = 40.51\%$; Water = 45.57%]

56. Carbohydrates are represented by the general formula $C_m(H_2O)_n$. On heating, in absence of air, they decompose into steam (H_2O) and carbon. 3.1 g of a carbohydrate, on complete decomposition by heating in absence of air, leave a residue of 1.24 g of carbon. If the molecular mass of the carbohydrate be 180, find its molecular formula.

[**Hint:** Determine % of carbon in carbohydrate. It is 40%. Water is 60%. Empirical formula = CH_2O. Molecular formula $= 6 \times CH_2O = C_6H_{12}O_6$.]

57. A gaseous hydrocarbon contains 85.7% carbon and 14.3% hydrogen. 1 litre of the hydrocarbon weighs 1.26 g at NTP. Determine the molecular formula of the hydrocarbon.

[**Ans.** C_2H_4]

58. Equal masses of oxygen, hydrogen and methane are taken in a container under identical conditions. Find the ratio of their volumes.

[**Ans.** 1 : 16 : 2]

59. How many moles are there in 1 m^3 of any gas at NTP?
[**Ans.** 44.6 moles]

60. A hydrated chloride of metal contains 18.26% metal and 32.42% chloride ion by mass. The specific heat of metal is 0.16. What is hydrated chloride?
[**Ans.** $CaCl_2 \cdot 6H_2O$]

61. 1.878 g of MBr_x when heated in a stream of HCl gas was completely converted to chloride MCl_x which weighed 1.0 g. The specific heat of metal is 0.14 cal g^{-1}. Calculate the molecular masses of metal bromide and metal chloride.
[**Ans.** Mol. mass of metal bromide = 285.54;
Mol. mass of metal chloride = 152.2]

62. An automobile antifreeze consists of 38.7% C; 9.7% H and remaining oxygen by weight. When 0.93 g of it are vaporised at 200°C and 1 atm pressure, 582 mL of vapour are formed. Find the molecular formula of the antifreeze.
[**Ans.** $C_2H_6O_2$]

63. A mineral contained MgO = 31.88%; SiO_2 = 63.37% and H_2O = 4.75%. Show that the simplest formula for the mineral is $H_2Mg_3Si_4O_{12}$.
(H = 1; Mg = 24; Si = 28; O = 16)

64. How many moles of NH_3 are there in 250 cm^3 of a 30% solution, the specific gravity of which is 0.90?
[**Ans.** 3.97 moles]

65. Haemoglobin contains 0.25% iron by mass. The molecular mass of haemoglobin is 89600. Calculate the number of iron atoms per molecule of haemoglobin.
[Atomic mass of Fe = 56]
[**Ans.** 4 atoms]

66. A sample of potato-starch was ground to give a starch like molecule. The product analysed 0.086% phosphorus. If each molecule is assumed to contain one atom of phosphorus, what is the average molecular mass of the material?
[**Ans.** 36000 amu]

67. Insulin contains 3.4% sulphur. Calculate minimum molecular mass of the insulin.
[**Ans.** 941.176 amu]
[**Hint:** For minimum molecular mass, one molecule of insulin must have atleast one sulphur atom.]

68. Calculate the number of carbon, hydrogen and oxygen atoms in 18 g of glucose.
[**Ans.** 3.61×10^{23} carbon atoms, 7.22×10^{23} hydrogen atoms, 3.61×10^{23} oxygen atoms]

69. Hydrated sulphate of a divalent metal of atomic weight 65.4 loses 43.85% of its weight on dehydration. Find the number of molecules of water of crystallisation in the formula of hydrated salt. **[JEE (West Bengal) 2005]**
[**Hint:** Formula of divalent hydrated metal sulphate will be

$$MSO_4 \cdot xH_2O$$

$$\text{Molecular mass of salt} = 65.4 + 96 + 18x = (161.4 + 18x)$$

$$\%\text{ of water} = \frac{18x}{161.4 + 18x} \times 100 = 43.85$$

On solving, $x = 7$

$\therefore$ Molecular formula of hydrated salt = $MSO_4 \cdot 7H_2O$]

70. A person with fever has a temperature of 102.5°F. What is the temperature in degree celsius?
[**Hint:** Use $C = \frac{5(F-32)}{9}$]

71. An ornamental ring contains 275 carats of diamond. How many grams diamond does it have?
[**Hint:** 1 carat = 200 mg
$\therefore$ Mass of diamond = $275 \times 200 \times 10^{-3}$ g]

72. 1 volume of a gaseous compound consisting C, H, O on complete combustion in presence of 2.5 volume of O_2 gives 2 vol. of steam and 2 vol. of CO_2. What is the formula of the compound if all measurements are made at NTP?
[**Ans.** C_2H_4O]

73. 60 mL of a mixture of nitrous oxide and nitric oxide was exploded with excess of hydrogen. If 38 mL of N_2 was formed, calculate the volume of each gas in the mixture.
[**Ans.** NO = 44 mL and N_2O = 16 mL]

74. For a precious stone, 'carat' is used for specifying its mass. If 1 carat = 3.168 grains (a unit of mass) and 1 gram = 15.4 grains, find the total mass in kilogram of the ring that contains 0.5 carat diamond and 7 gram gold.
[**Ans.** 7.1×10^{-3} kg]

75. The density of a gaseous element is 5 times that of oxygen under similar conditions. If the molecule of the element is triatomic, what will be its atomic mass?
[**Ans.** 53.33]

76. Calculate the number of electrons, protons and neutrons in 1 mole of $^{18}O^{2-}$ ions.
[**Ans.** Electrons = $10 \times 6.023 \times 10^{23}$
Protons = $8 \times 6.023 \times 10^{23}$
Neutrons = $8 \times 6.023 \times 10^{23}$]

77. 600 mL of a mixture of O_2 and O_3 weighs 1 gm at NTP. Calculate the volume of ozone in the mixture.
[**Ans.** 200 mL]

OBJECTIVE QUESTIONS

Set-1: Questions with single correct answer

1. The father of modern chemistry is:
(a) Priestley (b) Lavoisier
(c) Dalton (d) Mendeleev
2. A pure substance can only be:
(a) a compound
(b) an element
(c) an element or a compound
(d) a heterogeneous mixture
3. A pure substance which contains only one type of atoms is called:
(a) an element (b) a compound
(c) a solid (d) a liquid
4. Which one of the phrases would be incorrect to use?
(a) A mole of an element (b) A mole of a compound
(c) An atom of an element (d) An atom of a compound
5. A symbol not only represents the name of the element but also its:
(a) atomic mass (b) atomic number
(c) atomicity (d) atomic volume
6. The credit for the discovery of transuranic elements goes to:
(a) Hahn (b) Rutherford
(c) Seaborg (d) Curie
7. The most abundant metal in earth's crust is:
(a) iron (b) magnesium
(c) calcium (d) aluminium
8. The most abundant element in earth's crust is:
(a) hydrogen (b) oxygen (c) nitrogen (d) silicon
9. Which one of the elements is not found in nature?
(a) Radium (b) Technetium
(c) Polonium (d) Helium
10. Which one of the following is not a compound?
(a) Marble (b) Ozone
(c) Carborundum (d) Quicklime
11. Which one of the following is not an element?
(a) Diamond (b) Ozone (c) Silica (d) Graphite
12. The direct change from solid to gaseous state is referred to as:
(a) dissociation (b) decomposition
(c) sublimation (d) deliquescence
13. Sulphur burns in oxygen to form sulphur dioxide. The properties of sulphur dioxide are:
(a) totally different from sulphur and oxygen
(b) similar to sulphur
(c) similar to oxygen
(d) more similar to sulphur than oxygen
14. Which one of the following statements is incorrect?
(a) All elements are homogeneous
(b) Compounds always contain two or more different elements
(c) A mixture is not always heterogeneous
(d) Air is a heterogeneous mixture
15. One sample of air is found to have 0.03% carbon dioxide and another sample 0.02%. This illustrates that:
(a) air is a compound
(b) air is an element
(c) air does not follow the law of constant proportions
(d) air is a mixture
16. Which one of the following is not a mixture?
(a) Distilled water
(b) Sugar dissolved in water
(c) Liquefied Petroleum Gas (LPG)
(d) Gasoline
17. Which of the following is a characteristic property of both mixtures and compounds?
(a) Their properties are same as those of their components
(b) Energy is released when they are formed
(c) Their masses are equal to the sum of the masses of their components
(d) They contain the components in fixed proportions
18. Name the scientist who stated that matter can be converted into energy:
(a) Boyle (b) Lavoisier (c) Avogadro (d) Einstein
19. Which one of the following is not an intensive property?
(a) Weight (b) Density
(c) Refractive index (d) Melting point
20. The metalloid among the following group of elements is: **(CPMT 1993)**
(a) P (b) As (c) Al (d) Po
21. Which of the following alloys does contain Cu and Zn? **(IIT 1993)**
(a) Bronze (b) Brass
(c) Type metal (d) Rolled gold
22. Which metal is present in german silver?
(a) Copper (b) Iron (c) Silver (d) Zinc
23. Which of the following processes results in the formation of a new compound?
(a) Dissolving common salt in water
(b) Heating water
(c) Heating platinum rod
(d) Heating iron rod
24. Which one of the following is not a chemical change?
(a) Sublimation (b) Combustion
(c) Electrolysis (d) Rusting
25. The law of multiple proportions is illustrated by the pair of compounds:
(a) sodium chloride and sodium bromide
(b) water and heavy water
(c) sulphur dioxide and sulphur trioxide
(d) magnesium hydroxide and magnesium oxide
26. In compound A, 1.0 g nitrogen combines with 0.57 g oxygen. In compound B, 2.0 g nitrogen unite with 2.24 g oxygen and in compound C, 3.0 g nitrogen combine with 5.11 g oxygen. These results obey the law of:

(a) multiple proportions (b) constant proportions
(c) reciprocal proportions (d) none of these

27. Which one is the best example of law of conservation of mass?
(a) 6 g of carbon is heated in vacuum, there is no change in mass
(b) 6 g of carbon combines with 16 g of oxygen to form 22 g of CO_2
(c) 6 g water is completely converted into steam
(d) A sample of air is heated at constant pressure when its volume increases but there is no change in mass

28. A chemical equation is balanced according to the law of:
(a) multiple proportions (b) constant proportions
(c) reciprocal proportions (d) conservation of mass

29. SO_2 gas was prepared by (i) burning sulphur in oxygen, (ii) reacting sodium sulphite with dilute H_2SO_4 and (iii) heating copper with conc. H_2SO_4. It was found that in each case sulphur and oxygen combined in the ratio of 1 : 1. The data illustrates the law of:
(a) conservation of mass (b) multiple proportions
(c) constant proportions (d) reciprocal proportions

30. A sample of $CaCO_3$ has Ca = 40%, C = 12% and O = 48%. If the law of constant proportions is true, then the mass of Ca in 5 g of $CaCO_3$ from another source will be:
(a) 2.0 g (b) 0.2 g (c) 0.02 g (d) 20.0 g

31. Potassium combines with two isotopes of chlorine (^{35}Cl and ^{37}Cl) respectively to form two samples of KCl. Their formation follows the law of:
(a) constant proportions (b) multiple proportions
(c) reciprocal proportions (d) none of these

32. Different proportions of oxygen in the various oxides of nitrogen, prove the law of:
(a) reciprocal proportions (b) multiple proportions
(c) constant proportions (d) conservation of mass

33. One part of an element *A* combines with two parts of *B* (another element). Six parts of element *C* combine with four parts of element *B*. If *A* and *C* combine together, the ratio of their masses will be governed by:
(a) law of definite proportions
(b) law of multiple proportions
(c) law of reciprocal proportions
(d) law of conservation of mass

34. H_2S contains 5.88% hydrogen, H_2O contains 11.11% hydrogen while SO_2 contains 50% sulphur. These figures illustrate the law of:
(a) conservation of mass (b) constant proportions
(c) multiple proportions (d) reciprocal proportions

35. Number of atoms in 4.25 g of NH_3 is: **(AFMC 2010)**
(a) 6.023×10^{23} (b) $4\times6.023\times10^{23}$
(c) 1.7×10^{24} (d) $4.5\times6.023\times10^{23}$

[**Hint:** Number of molecules of $NH_3 = \frac{w}{M}\times6.023\times10^{23}$

$$= \frac{4.25}{17}\times6.023\times10^{23}$$

Number of atom $= 4\times\frac{4.25}{17}\times6.023\times10^{23}$

$$= 6.023\times10^{23}]$$

36. Hydrogen combines with chlorine to form HCl. It also combines with sodium to form NaH. If sodium and chlorine also combine with each other, they will do so in the ratio of their masses as:
(a) 23 : 35.5 (b) 35.5 : 23
(c) 1 : 1 (d) 23 : 1

37. Zinc sulphate contains 22.65% Zn and 43.9% H_2O. If the law of constant proportions is true, then the mass of zinc required to give 40 g crystals will be:
(a) 90.6 g (b) 9.06 g (c) 0.906 g (d) 906 g

38. 3 g of a hydrocarbon on combustion in excess of oxygen produces 8.8 g of CO_2 and 5.4 g of H_2O. The data illustrates the law of:
(a) conservation of mass (b) multiple proportions
(c) constant proportions (d) reciprocal proportions

[**Hint:** Mass of carbon in 8.8 g $CO_2 = \frac{12}{44}\times8.8 = 2.4$g;

Mass of hydrogen in 5.4 g $H_2O = \frac{2}{18}\times5.4 = 0.6$ g

Total mass of (C + H) = 2.4 + 0.6 = 3.0 g]

39. In the reaction, $N_2 + 3H_2 \longrightarrow 2NH_3$, the ratio of volumes of nitrogen, hydrogen and ammonia is 1 : 3 : 2. These figures illustrate the law of:
(a) constant proportions (b) Gay-Lussac
(c) multiple proportions (d) reciprocal proportions

40. Two volumes of ammonia, on dissociation gave one volume of nitrogen and three volumes of hydrogen. How much hydrogen will be obtained from the dissociation of 10 litre of NH_3?
(a) 30 litre (b) 10 litre (c) 15 litre (d) 20 litre

41. If 6 litre of H_2 and 5.6 litre of Cl_2 are mixed and exploded in an eudiometer, the volume of HCl formed is:
(a) 6.0 litre (b) 5.6 litre (c) 11.2 litre (d) 11.6 litre

42. The law of constant proportions was enunciated by:
(a) Dalton (b) Berthelot (c) Avogadro (d) Proust

43. An important postulate of Dalton's atomic theory is:
(a) an atom contains electrons, protons and neutrons
(b) atom can neither be created nor destroyed nor divisible
(c) all the atoms of an element are not identical
(d) all the elements are available in nature in the form of atoms

44. The atomic masses of the elements are usually fractional because:
(a) elements consist of impurities
(b) these are mixtures of allotropes
(c) these are mixtures of isobars
(d) these are mixtures of isotopes

45. The chemical formula of a particular compound represents:
(a) the size of its molecule
(b) the shape of its molecule
(c) the total number of atoms in a molecule
(d) the number of different types of atoms in a molecule

46. Which one of the following properties of an element is not variable?
(a) Valency (b) Atomic mass
(c) Equivalent mass (d) All of these

47. 1 amu is equal to:
(a) 1.00758 g (b) 0.000549 g
(c) 1.66×10^{-24} g (d) 6.02×10^{-23} g

48. Which one of the following relationships is correct?
(a) At. mass = 6.4 × Sp. heat.
(b) At. mass × Sp. heat = 6.4
(c) At. mass × 6.4 = Sp. heat
(d) At. mass × Sp. heat × 6.4 = 1

49. A_1 g of an element gives A_2 g of its oxide. The equivalent mass of the element is:
(a) $\frac{A_2 - A_1}{A_1} \times 8$ (b) $\frac{A_2 - A_1}{A_2} \times 8$
(c) $\frac{A_1}{A_2 - A_1} \times 8$ (d) $(A_2 - A_1) \times 8$

50. A_1 g of an element gives A_2 g of its chloride; the equivalent mass of the element is:
(a) $\frac{A_1}{A_2 - A_1} \times 35.5$ (b) $\frac{A_2}{A_2 - A_1} \times 35.5$
(c) $\frac{A_2 - A_1}{A_1} \times 35.5$ (d) $\frac{A_2 - A_1}{A_2} \times 35.5$

51. Which one of the relationship is wrong?
(a) 2 V.D. = Mol. mass
(b) At. mass = Eq. mass × Valency
(c) At. mass = $\frac{6.4}{\text{Sp. heat}}$
(d) Valency = $\frac{\text{Mol. mass}}{\text{Eq. mass}}$

52. In m_1 g of a metal A displaces m_2 g of another metal B from its salt solution and if their equivalent masses are E_1 and E_2 respectively, then the equivalent mass of A can be expressed as:
(a) $E_1 = \frac{m_2 \times E_2}{m_1}$ (b) $E_1 = \frac{m_1}{m_2} \times E_2$
(c) $E_1 = \frac{m_1 \times m_2}{E_2}$ (d) $E_1 = \sqrt{\frac{m_1}{m_2} \times E_2}$

53. When the specific heat of a metallic element is 0.214 cal g^{-1}, the atomic mass will be closest to which one of the following?
(a) 1 (b) 12 (c) 30 (d) 66

54. Approximate atomic mass of an element is 26.89. If its equivalent mass is 8.9, the exact atomic mass of the element is:
(a) 26.89 (b) 8.9 (c) 17.8 (d) 26.7

55. When an element forms an oxide in which oxygen is 20% of the oxide by mass, the equivalent mass of the element will be:
(a) 32 (b) 40 (c) 60 (d) 128

56. 0.32 g of a metal gave on treatment with an acid 112 mL of hydrogen at NTP. Calculate equivalent mass of the metal: **[AMU (Engg.) 2010]**
(a) 58 (b 32 (c) 11.2 (d) 24
[**Hint:** Mass of metal that displaces 11200 mL hydrogen at STP will be its equivalent mass.
∴ Equivalent mass of metal = $\frac{0.32}{112} \times 11200 = 32$]

57. 74.5 g of a metallic chloride contains 35.5 g of chlorine. The equivalent mass of the metal is:
(a) 19.5 (b) 35.5 (c) 39.0 (d) 78.0

58. The product of atomic mass and specific heat of any element is a constant, approximately 6.4. This is known as:
(a) Dalton's law (b) Avogadro's law
(c) Gay-Lussac law (d) Dulong Petit's law

59. The molecular mass of chloride, MCl, is 74.5. The equivalent mass of the metal M will be:
(a) 39.0 (b) 74.5
(c) 110.0 (d) 35.5
[**Hint:** Mol. mass = At. mass + 35.5 = Eq. mass × valency + 35.5, Valency of M from the formula MCl is 1]

60. 1 g of hydrogen is found to combine with 80 g of bromine. 1 g of calcium combines with 4 g of bromine. The equivalent mass of calcium is:
(a) 10 (b) 20 (c) 40 (d) 80

61. 2.8 g of iron displaces 3.2 g of copper from a solution of copper sulphate. If the equivalent mass of iron is 28, the equivalent mass of copper will be:
(a) 16 (b) 32 (c) 48 (d) 64

62. The specific heat of a metal of atomic mass 32 is likely to be:
(a) 0.25 (b) 0.24 (c) 0.20 (d) 0.15

63. The equivalent mass of an element is 4. Its chloride has a vapour density 59.25. The valency of the element will be:
(a) 4 (b) 3 (c) 2 (d) 1

64. The equivalent mass of iron in the reaction,
$3Fe + 4H_2O \rightleftharpoons Fe_3O_4 + 4H_2$ would be:
(a) 21 (b) 56 (c) 42 (d) 10

65. The specific heat of a bivalent metal is 0.16. The approximate equivalent mass of the metal will be:
(a) 40 (b) 20 (c) 80 (d) 10

66. A sample of pure calcium weighing 1.35 g was quantitatively converted to 1.88 g of pure calcium oxide. Atomic mass of calcium would be:
(a) 20 (b) 40 (c) 16 (d) 35.5

67. A metal oxide is reduced by heating it in a stream of hydrogen. It is found that after complete reduction, 3.15 g of the oxide have yielded 1.05 g of the metal. We may conclude that: **(MLNR 1991)**
(a) atomic mass of the metal is 4
(b) atomic mass of the metal is 8
(c) equivalent mass of the metal is 4
(d) equivalent mass of the metal is 8

68. Compounds with identical crystal structure and analogous chemical formula are called:
(a) isomers (b) isotones
(c) allotropes (d) isomorphous

69. Which pair of the following substances is said to be isomorphous?
(a) White vitriol and blue vitriol
(b) Epsom salt and Glauber's salt
(c) Blue vitriol and Glauber's salt
(d) White vitriol and epsom salt

c mass of chlorine is 35.5. It has two isotopes of atomic
5 and 37. The percentage of heavier isotope is:
(b) 15 (c) 20 (d) 25

c mass of boron is 10.81. It has two isotopes with 80%
0% abundance respectively. The atomic mass of the
e having 80% abundance is 11.01. The atomic mass of
ier isotope is:
.81 (b) 11.01 (c) 10.01 (d) 21.82

of chlorine combines with a metal giving 111 g of its
le. The chloride is isomorphous with $MgCl_2 \cdot 6H_2O$. The
mass of the metal is:
(b) 30 (c) 40 (d) 69

apour density of a volatile chloride of a metal is 59.5 and
uivalent mass of the metal is 24. The atomic mass of the
nt will be:
(b) 48 (c) 24 (d) 12

xide of an element possesses the molecular formula,
. If the equivalent mass of the metal is 9, the atomic
of the metal will be:
(b) 18 (c) 9 (d) 4.5

olecular mass of a compound having formula MO and
lent mass 20 is:
(b) 40 (c) 28 (d) 20

ensity of air is 0.001293 g mL^{-1}. Its vapour density is:
3 (b) 14.3 (c) 1.43 (d) 0.143
Divide with the density of hydrogen, *i.e.*, 0.00009 g

of a substance when vaporised occupy a volume of 5.6
NTP. The molecular mass of the substance will be:
(b) $2M$ (c) $3M$ (d) $4M$

apour densities of two gases are in the ratio of 1 : 3.
molecular masses are in the ratio of:
3 (b) 1 : 2 (c) 2 : 3 (d) 3 : 1

ganic compound on analysis was found to contain
% of sulphur. The molecular mass of the compound, if its
ule contains two sulphur atoms, is:
0 (b) 2000
000 (d) 200000

tomic mass of an element is 27. If valency is 3, the
r density of the volatile chloride will be:
.75 (b) 6.675 (c) 667.5 (d) 81

ensity of a gas 'A' is three times that of a gas 'B'. If the
ular mass of A is M, the molecular mass of B is:
(b) $M/3$ (c) $\sqrt{3}M$ (d) $\frac{M}{\sqrt{3}}$

r density of a volatile substance is 4 in comparison to
ne ($CH_4 = 1$). Its molecular mass will be:
(b) 2 (c) 64 (d) 128

e the wrong statement:
mole means 6.02×10^{23} particles
olar mass is mass of one molecule
olar mass is mass of one mole of a substance
olar mass is molecular mass expressed in grams

le of CO_2 contains: **(MLNR 1990; CBSE 1993)**
2×10^{23} atoms of C

(c) 18.1×10^{23} molecules of CO_2
(d) 3 g-atoms of CO_2

85. Which among the following is the heaviest? **[PMT (Kerala) 2006]**
(a) One mole of oxygen
(b) One molecule of sulphur trioxide
(c) 100 amu of uranium
(d) 10 moles of hydrogen
(e) 44 g of carbon dioxide

86. The largest number of molecules is in:
(a) 28 g of CO (b) 46 g of C_2H_5OH
(c) 36 g of H_2O (d) 54 g of N_2O_5

87. Which of the following has the smallest number of molecules?
(a) 22.4×10^3 mL of CO_2 gas
(b) 22 g of CO_2 gas
(c) 11.2 litre of CO_2 gas
(d) 0.1 mole of CO_2 gas

88. The number of grams of H_2SO_4 present in 0.25 mole of H_2SO_4 is:
(a) 0.245 (b) 2.45 (c) 24.5 (d) 49.0

89. Number of molecules in 1 litre of oxygen at NTP is:
(a) $\frac{6.02 \times 10^{23}}{32}$ (b) $\frac{6.02 \times 10^{23}}{22.4}$
(c) 32×22.4 (d) $\frac{32}{22.4}$

90. 4.6×10^{22} atoms of an element weigh 13.8 g. The atomic mass of the element is:
(a) 290 (b) 180 (c) 34.4 (d) 10.4

91. The number of molecules in 89.6 litre of a gas at NTP are: **(BHU 1992)**
(a) 6.02×10^{23} (b) $2 \times 6.02 \times 10^{23}$
(c) $3 \times 6.02 \times 10^{23}$ (d) $4 \times 6.02 \times 10^{23}$

92. The total number of protons in 10 g of calcium carbonate is: **(CPMT 1992)**
(a) 3.0115×10^{24} (b) 1.5057×10^{24}
(c) 2.0478×10^{24} (d) 4.0956×10^{24}

93. 19.7 kg of gold was recovered from a smuggler. The atoms of gold recovered are: (Au = 197)
(a) 100 (b) 6.02×10^{23}
(c) 6.02×10^{24} (d) 6.02×10^{25}

94. The molecular mass of CO_2 is 44 amu and Avogadro's number is 6.02×10^{23}. Therefore, the mass of one molecule of CO_2 is:
(a) 7.31×10^{-23} (b) 3.65×10^{-23}
(c) 1.01×10^{-23} (d) 2.01×10^{-23}

95. Equal volumes of different gases at any definite temperature and pressure have:
(a) equal weights (b) equal masses
(c) equal densities (d) equal number of moles

96. A gaseous mixture contains oxygen and nitrogen in the ratio of 1 : 4 by mass. Therefore, the ratio of their number of molecules is:

m will react with ethyl alcohol to give:
of hydrogen (b) one mole of oxygen
of NaOH (d) 1/2 mole of hydrogen
an atom of carbon is:
(b) 1.99×10^{-23} g
(d) 1.99×10^{23} g
of moles of H_2 in 0.224 litre of hydrogen gas at
(MLNR 1994)
(b) 0.1 (c) 0.01 (d) 0.001
of 1 g helium at NTP in litres is:
(b) 0.56 (c) 2.8 (d) 0.28
s to contain 1.0×10^{24} particles, the mass of one
gen is:
(b) 5.32 g
(d) 16.0 g
of 1 litre capacity each are separately filled with
He, O_2 and O_3. At the same temperature and
ratio of the number of atoms of these gases
fferent flasks would be:
1 (b) 2 : 1 : 2 : 3
3 (d) 3 : 2 : 2 : 1
mperature and pressure, two flasks of equal
filled with H_2 and SO_2 separately. Particles which
number, in the two flaskes are:
(b) electrons
es (d) neutrons
a mixture of 6.02×10^{23} oxygen atoms and
ydrogen molecules at NTP is:
(b) 33.6 litre
(d) 22.4 litre
of molecules present in a drop of water, if its
05 mL, are:
^{21} (b) 1.66×10^{22}
^{23} (d) 1.66×10^{24}
was found to contain nitrogen and oxygen in the
n 28 g and oxygen 80 g. The formula of the
:
(b) N_2O_3
(d) N_2O_4
t formula of a compound containing 50% of
At. mass = 10) and 50% of the element Y (At.
s:
(IIT 1994)
(b) X_2Y
(d) X_2Y_3
ts X (At. mass = 75) and Y (At. mass = 16)
ive a compound having 75.8% of X. The formula
und is: **(MLNR 1991)**
(b) X_2Y
(d) X_2Y_3
a certain compound was found to contain iodine
in the ratio of 254 : 80. The formula of the
: (At. mass of I = 127, O = 16)
(b) I_2O
(d) I_2O_5

110. A compound of aluminium and chlorine is composed of 9.0 g Al for every 35.5 g of chlorine. The empirical formula of the compound is:
(a) AlCl (b) $AlCl_3$ (c) $AlCl_2$ (d) $AlCl_4$

111. The haemoglobin from red corpuscles of most mammels contain approximately 0.33% of iron by mass. The molecular mass of haemoglobin is 67200. The number of iron atoms in each molecule of haemoglobin is:
(a) 4 (b) 3 (c) 2 (d) 1

112. The percentage of P_2O_5 in diammonium hydrogen phosphate, $[(NH_4)_2HPO_4]$ is: **(CPMT 1992)**
(a) 23.48 (b) 46.96
(c) 53.78 (d) 71.00

113. The percentage of nitrogen in urea (NH_2CONH_2), is:
(a) 38.4 (b) 46.6 (c) 59.1 (d) 61.3

114. The chloride of a metal has the formula MCl_3. The formula of its phosphate is:
(a) M_2PO_4 (b) MPO_4 (c) M_3PO_4 (d) $M(PO_4)_2$

115. 10 g of hydrofluoric acid gas occupies 5.6 litre of volume at NTP. If the empirical formula of the gas is HF, then its molecular formula will be: (At. mass of F = 19)
(a) HF (b) H_3F_3
(c) H_2F_2 (d) H_4F_4
[**Hint:** Molecular mass $= \dfrac{10}{5.6} \times 22.4 = 40$]

116. Calcium pyrophosphate is represented by the formula $Ca_2P_2O_7$. The molecular formula of ferric pyrophosphate is:
(a) $Fe_2P_2O_7$ (b) FeP_2O_7
(c) $Fe(P_2O_7)_3$ (d) $Fe_4(P_2O_7)_3$

117. The percentage of available chlorine in a sample of bleaching powder, $CaOCl_2 \cdot 2H_2O$, is:
(a) 30 (b) 50 (c) 43.5 (d) 59.9

118. SI unit of energy is:
(a) kg m^2 s^{-2} (b) kg m^{-1} s^2
(c) kg m^2 s^{-1} (d) kg m^2 s^2

119. One micro gram is equal to:
(a) 10^{-3} g (b) 10^3 g (c) 10^6 g (d) 10^{-6} g

120. Significant figures in 0.00051 are:
(a) 5 (b) 3 (c) 2 (d) 4

121. The number of significant figures in 6.02×10^{23} is:
(a) 23 (b) 3 (c) 4 (d) 26

122. Express 0.006006 into scientific notation in three significant digits:
(a) 6.01×10^{-3} (b) 6.006×10^{-3}
(c) 6.00×10^{-3} (d) 6.0×10^{-3}

123. The proper value of significant figures in 38.0 + 0.0035 + 0.00003 is:
(a) 38 (b) 38.0035
(c) 38.00353 (d) 38.0

124. Which of the following is the correct unit for measuring nuclear radii?
(a) Micron (b) Millimetre

er to prepare 1 litre normal solution of $KMnO_4$, how grams of $KMnO_4$ are required if the solution is to be acid medium for oxidation? **[PET (MP) 2002]**

8 g (b) 31.6 g (c) 62 g (d) 790 g

an oxide of a metal is converted to chloride completely yields 5 g chloride. The equivalent weight of metal is: **(KCET 2002)**

.25 (b) 3.325 (c) 12 (d) 20

er of atoms 558.5 g Fe (At. wt. of Fe = 55.85 g mol^{-1}) is: **(AIEEE 2002)**

ice that in 60 g carbon (b) 6.023×10^{23}

lf that of 8 g He (d) $558.5 \times 6.023 \times 10^{23}$

efix 10^{18} is: **[MEE (Kerala) 2002]**

ga (b) exa (c) kilo (d) nano

ega

ence in density is the basis of: **[MEE (Kerala) 2002]**

ra filtration (b) molecular sieving

avity separation (d) molecular attraction

omic absorption

ective of the source, pure sample of water always yields % mass of oxygen and 11.11% mass of h... This is ned by the law of: **(Kerala) 2002]**

nservation of mass (b) constant composition

ltiple proportion (d) constant volume

y-Lussac's l...

...any ... of electron weigh one kilogram? **[IIT (Screening) 2002]**

3×10^{23} (b) $\frac{1}{9.108} \times 10^{31}$

$\frac{3 \times 10^{54}}{9.108}$ (d) $\frac{1}{9.108 \times 6.023} \times 10^{8}$

he numbers: 161 cm; 0.161 cm; 0.0161 cm. The of significant figure for three numbers is: **[AFMC (Pune) 2002]**

, 5 (b) 3, 3, 3 (c) 3, 3, 4 (d) 3, 4, 4

one of the following laws directly explains the law of ation of mass? **(AFMC 2002)**

d's rule (b) Dalton's law

gadro's law (d) Berzelius hypothesis

as maximum number of atoms? **[IIT (Screening) 2003]**

C(12) (b) 56 g Fe(56)

Al(27) (d) 108 g Ag(108)

ich of sulphur is present in an organic compound, if ompound gave 1.158 g of $BaSO_4$ on analysis? **[PET (Kerala) 2005]**

(b) 15% (c) 20% (d) 25%

of H_2 and 20 mL of O_2 react to form water, what is e end of the reaction? **(AFMC 2005)**

137. The density of a liquid is 1.2 g/mL. There are 35 drops in 2 mL. The number of molecules in one drop (molar mass of liquid = 70) is:

(a) $\left(\frac{1.2}{35}\right) N_A$ (b) $\left(\frac{1}{35}\right)^2 N_A$

(c) $\frac{1.2}{(35)^2} N_A$ (d) $1.2 N_A$

138. A sample of PCl_3 contains 1.4 mole of the substance. How many atoms are there in the sample? **[CEE (Kerala) 2004]**

(a) 4 (b) 5.6

(c) 8.431×10^{23} (d) 3.372×10^{24}

(e) 2.409×10^{24}

139. The equivalent weight of phosphoric acid H_3PO_4 in the reaction, $NaOH + H_3PO_4 \longrightarrow NaH_2PO_4 + H_2O$ is: **[BHU (Pre.) 2005]**

(a) 59 (b) 49 (c) 25 (d) 98

[Hint: Only one hydrogen of H_3PO_4 is replaced, *i.e.*, its basicity = 1

$$\text{Equivalent mass} = \frac{\text{Molecular mass}}{\text{Basicity}} = \frac{98}{1} = 98]$$

140. 5.6 g of an organic compound on burning with excess of oxygen gave 17.6 g of CO_2 and 7.2 g H_2O. The organic compound is: **[PET (Kerala) 2006]**

(a) C_6H_6 (b) C_4H_8 (c) C_3H_8 (d) CH_3COOH

(e) CH_3CHO

141. The decomposition of a certain mass of $CaCO_3$ gave 11.2 dm^3 of CO_2 gas at STP. The mass of KOH required to completely neutralise the gas is: **(KCET 2006)**

(a) 56 g (b) 28 g (c) 42 g (d) 20 g

[**Hint:** 11.2 dm^3 of CO_2 at STP = $\frac{1}{2}$ mole CO_2

$$KOH + CO_2 \longrightarrow KHCO_3$$

$\frac{1}{2}$ mole CO_2 will be neutralised by $\frac{1}{2}$ mole KOH, *i.e.*, 28 g KOH.]

142. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$, will contain 0.25 mole of oxygen atoms? **(AIEEE 2006)**

(a) 0.02 (b) 3.125×10^{-2}

(c) 1.25×10^{-2} (d) 2.5×10^{-2}

[**Hint:**

∵ 8 mole oxygen atoms are present in 1 mole $Mg_3(PO_4)_2$

∴ 0.25 mole oxygen atoms will be present in $\frac{1}{8} \times 0.25$ mole

$Mg_3(PO_4)_2$, *i.e.*, 3.125×10^{-2} mole $Mg_3(PO_4)_2$]

143. An element, X has the following isotopic composition,

$^{200}X : 90\%$

$^{199}X : 8\%$

$^{202}X : 2\%$

the weighted average atomic mass of the naturally- occurring element 'X' is closest to: **[CBSE (Med.) 2007]**

(a) 201 amu (b) 202 amu (c) 199 amu (d) 200 amu age atomic mass of

$$\times 200\Big] + \Big[\frac{8}{100} \times 199\Big] + \Big[\frac{2}{100} \times 202\Big]$$

5 amu $\approx$ 200 amu]

statement for 14 g CO is : **(VMMC 2007)**

s 2.24 L at NTP

ponds to $\frac{1}{2}$ mole of CO

pond to same mole of CO and N_2

ponds 3.01×10^{23} molecules of CO

r of hydrogen atoms present in 25.6 g of sucrose

) which has molar mass of 342.3 g, is:

(VITEEE 2008)

3 (b) 9.91×10^{23}

3 (d) 44 ... 23

mber of moles of sucrose

$$= \frac{\text{Mass}}{\text{Molar mass}} = \frac{25.6}{342.3}$$

$$= 0.075$$

noles of hydrogen atom $= 0.075 \times 22$

toms of hydrogen $= 0.075 \times 22 \times 6.023 \times 10^{23}$

$$= 9.9 \times 10^{23}]$$

upied by one molecule of water (density = 1 g/cm^3)

[CBSE-PMT (Pre.) 2008]

^{3} cm^3 (b) 5.5×10^{-23} cm^3

^{3} cm^3 (d) 6.023×10^{-23} cm^3

ass of one molecule $= \frac{18}{6.023 \times 10^{23}}$ g

$$= 2.98 \times 10^{-23} \text{ g}$$

ne molecule $= \frac{M}{\text{Density}} = \frac{2.98 \times 10^{-23}}{1}$ cm^3

$$\approx 3 \times 10^{-23} \text{ cm}^3]$$

gen contains as many atoms as in: **(KCET 2008)**

hydrogen (b) 5 g of hydrogen

hydrogen (d) 1 g of hydrogen

s consists of uniform cylindrical particles of 150Å and 5000Å long. The specific volume of virus is If the virus is considered to be a single particle, its mass is:

)^7 g mol^{-1} (b) 7.90×10^7 g mol^{-1}

)^7 g mol^{-1} (d) 9.70×10^7 g mol^{-1}

ume of single virus $= \pi r^2 h$

$$= 3.14 \times (75 \times 10^{-8})^2 \times (5000 \times 10^{-8})$$

$$= 8.836 \times 10^{-17} \text{ cm}^3$$

ingle virus $= \frac{\text{Volume}}{\text{Specific volume}} = \frac{8.836 \times 10^{-17} \text{ cm}^3}{0.75 \text{ cm}^3/\text{g}}$

$$= 1.178 \times 10^{-16} \text{ g}$$

ss of virus $= 1.178 \times 10^{-16} \times 6.023 \times 10^{23}$

$$= 7.09 \times 10^7 \text{ g mol}^{-1}]$$

149. Common salt obtained from sea-water contains 95% NaCl by mass. The approximate number of molecules present in 10 g salt is: **(DPMT 2009)**

(a) 10^{21} (b) 10^{22}

(c) 10^{23} (d) 10^{24}

[**Hint :** Mass of NaCl in 10 g salt $= 10 \times \frac{95}{100} = 9.5$ g

Number of molecules of NaCl $= \frac{9.5}{58.5} \times 6.023 \times 10^{23}$

$$= 9.78 \times 10^{22} \approx 10^{23}]$$

150. 10 g hydrogen and 64 g oxygen were filled in a steel vessel and exploded. Amount of water produced in this reaction will be: **[CBSE (PMT) 2009]**

(a) 3 mol (b) 4 mol

(c) 1 mol (d) 2 mol

[**Hint :** $n_{H_2} = \frac{10}{2} = 5$ $n_{O_2} = \frac{64}{32} = 2$

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

Case I : If H_2 is completely consumed then :

$$n_{H_2O} = 5 \text{ mol}$$

Case II : If O_2 is completely consumed then

$$n_{H_2O} = \frac{2}{1} \times 2 = 4 \text{ mol}$$

Since, O_2 gives le... hence, it is limiting amount of product on complete consumption ... number of moles of water formed will be equal to 4.]

151. An organic compound made of C, H and N contains 20% nitrogen. What will be its molecular mass if it contains only one nitrogen atom in it? **[J... (WB) 2009]**

(a) 70 (b) 140

(c) 100 (d) 65

[**Hint :** $\% \text{ N} = \frac{\text{Mass of nitrogen}}{\text{Molecular mass}} \times 100$

$$20 = \frac{14}{m} \times 100$$

$$m = 70]$$

152. Given that the abundances of isotopes ^{54}Fe, ^{56}Fe and ^{57}Fe are 5%, 90% and 5% respectively, the atomic mass of Fe is: **(IIT 2009)**

(a) 55.85 u (b) 55.95 u

(c) 55.75 u (d) 56.05 u

[**Hint :** Atomic mass of Fe $= \frac{5}{100} \times 54 + \frac{90}{100} \times 56 + \frac{5}{100} \times 57$

$$= 55.95 \text{ amu}]$$

153. The number of atoms in 0.1 mol triatomic gas is: ($N_A = 6.02 \times 10^{23}$ mol^{-1}) **[CBSE (PMT) 2010]**

(a) 6.026×10^{22} (b) 1.806×10^{23}

(c) 3.6×10^{23} (d) 1.8×10^{22}

[**Hint :** No. of atoms $= 0.1 \times 3 \times 6.02 \times 10^{23}$

$$= 1.806 \times 10^{23}]$$

questions given below may have more than one correct answers

'the following relationships are wrong?

ı ≈ 0.1 bar (b) 1 litre = 1 dm^3

0.239 cal (d) 1 eV = 9.11 × 10^{-4} J

ɔf the following numbers have same significant

ɔ (b) 0.60 (c) 6.0 (d) 60

f the following have the same mass?

nole of O_2 gas

nole of SO_2 gas

3 × 10^{22} molecules of SO_2 gas

4 × 10^{23} molecules of O_2 gas

4. Select the numbers with same significant figures:
   (a) 6.02×10^{23} (b) 0.25
   (c) 6.60×10^{-34} (d) 1.50
5. Which are isomorphic to each other?
   (a) $CuSO_4 \cdot 5H_2O$ (b) $ZnSO_4 \cdot 7H_2O$
   (c) $FeSO_4 \cdot 7H_2O$ (d) $FeSO_4 \cdot 8H_2O$
6. 11.2 L of a gas at STP weighs 14 g. The gas could be:
   (a) N_2 (b) CO
   (c) NO_2 (d) N_2O
7. 8 g O_2 has same number of molecules as that in:
   (a) 14 g CO (b) 7 g CO
   (c) 11 g CO_2 (d) 22 g CO_2

Assertion-Reason TYPE QUESTIONS

e following questions, two statements are given as A) and '**Reason**' (**R**). Answer the questions by er codes given below:

ɔth (A) and (R) are correct and (R) is the correct anation of (A).

ɔth (A) and (R) are correct but (R) is not the correct anation of (A).

) is correct and (R) is wrong.

) is wrong but (R) is correct.

ɔth (A) and (R) are wrong.

O_2 and 1 g O_3 have equal number of atoms.

ss of 1 mole atom is equal to its gram-atomic mass.

our density of sulphur vapour relative to oxygen is 2 ause sulphur atom is twice as heavy as that of oxygen n.

our density depends upon the molecular state of the stance in vapour state.

vogram is equal to 1 amu.

ɔgram is reciprocal of Avogadro's number.

ole H_2SO_4 contains same mass of oxygen and sulphur.

ole H_2SO_4 represents 98 g mass.

nole oxygen and N_2 have same volume at same perature and pressure.

ole gas at NTP occupies 22.4 litre volume at STP.

6. (A) Empirical formula of glucose is HCHO.
   (R) Molecular formula of glucose will also be equal to HCHO.
7. (A) The volume of 1 mole of an ideal gas at 1 bar pressure at 25°C is 24.78 litre.
   (R) 1 bar = 0.987 atm.
8. (A) Atomic weight = Specific heat (cal/mol) × 6.4
   (R) The formula is valid for metals only.
9. (A) Number of moles of H_2 in 0.224 L of H_2 is 0.01 mol.
   (R) 22.4 litres of H_2 at STP contains 6.023×10^{23} mol. **(AIIMS 1996)**
10. (A) The equivalent weight of an element is variable.
    (R) The valency of an element is variable. **(AIIMS 1995)**
11. (A) The number of significant figures in 507000 is three.
    (R) In 507000, all the zeros are significant.
12. (A) Law of conservation of mass is invalid for nuclear fission, fusion and disintegration.
    (R) The law proposes that mass is neither created nor destroyed in a reaction.
13. (A) Mass spectrometer is used for determination of atomic mass of isotopes.
    (R) Isotopes are the atoms of same element having same atomic number but different mass numbers.

E QUESTIONS

3. (a)	4. (d)	5. (a)	6. (c)	7. (d)	8. (b)
1. (c)	12. (c)	13. (a)	14. (d)	15. (d)	16. (a)
9. (a)	20. (b)	21. (b)	22. (a)	23. (d)	24. (a)
7. (b)	28. (d)	29. (c)	30. (a)	31. (d)	32. (b)
5. (a)	36. (a)	37. (b)	38. (a)	39. (b)	40. (c)
3. (b)	44. (d)	45. (d)	46. (b)	47. (c)	48. (b)
1. (d)	52. (b)	53. (c)	54. (d)	55. (a)	56. (b)
9. (a)	60. (b)	61. (b)	62. (c)	63. (b)	64. (a)
7. (c)	68. (d)	69. (d)	70. (d)	71. (c)	72. (c)
5. (a)	76. (b)	77. (d)	78. (a)	79. (d)	80. (a)
3. (b)	84. (a)	85. (e)	86. (c)	87. (d)	88. (c)
1. (d)	92. (a)	93. (d)	94. (a)	95. (d)	96. (c)
9. (c)	100. (a)	101. (a)	102. (b)	103. (c)	104. (d)
7. (b)	108. (d)	109. (d)	110. (b)	111. (a)	112. (c)
5. (c)	116. (d)	117. (c)	118. (a)	119. (d)	120. (c)
3. (d)	124. (d)	125. (b)	126. (a)	127. (a)	128. (b)
1. (d)	132. (b)	133. (c)	134. (a)	135. (e)	136. (d)
). (d)	140. (b)	141. (b)	142. (b)	143. (d)	144. (a)
7. (b)	148. (a)	149. (c)	150. (b)	151. (a)	152. (b)

3. (b, c)	4. (a, c, d)	5. (b, c)	6. (a, b)	7. (b, c)

I-REASON TYPE QUESTIONS

(c)	4. (d)	5. (b)	6. (c)	7. (b)	8. (b)
(e)	12. (b)	13. (a)			

BRAIN STORMING PROBLEMS

OBJECTIVE QUESTIONS for IIT ASPIRANTS

ng questions contain single correct option:

he following table:

npound l. mass)	Mass of the compound (in grams) taken
CO_2 (44)	4.4
NO_2 (46)	2.3
H_2O_2 (34)	6.8
SO_2 (64)	1.6

two compounds have least mass of oxygen? cular masses of compounds are given in brackets.)

(EAMCET 2004)

nd IV (b) I and III (c) I and II (d) III and IV

I. Mass of oxygen present $= \frac{4.4}{44} \times 32 = 3.2$ g

II. Mass of oxygen present $= \frac{2.3}{46} \times 32 = 1.6$ g

II. Mass of oxygen present $= \frac{6.8}{34} \times 32 = 6.4$ g

V. Mass of oxygen present $= \frac{1.6}{64} \times 32 = 0.8$ g

d IV have least mass of oxygen.]

osphate of a certain metal M is $M_3(PO_4)_2$. The correct a of metal sulphate would be:

$(SO_4)_3$ (b) MSO_4

$(SO_4)_2$ (d) M_2SO_4

rcentage of Se in peroxidase enzyme is 0.5% by mass mass of Se = 78.4 amu). Then, the minimum lar mass of enzyme which contains not more than one n is:

68×10^4 amu (b) 1.568×10^7 amu

68×10^3 amu (d) 1.568×10^6 amu

$\because$ 0.5 g Se is present in 100 g enzyme.

Se will be present in $\frac{100}{0.5} \times 78.4$ g enzyme

$= 15680$ amu

$= 1.568 \times 10^4$ amu]

mber of moles of a gas in 1 m^3 of volume at NTP is:

(b) 0.446 (c) 1.464 (d) 44.6

1 m^3 = 1000 L

nber of moles $= \frac{1000}{22.4} = 44.6$]

al number of electrons present in 18 mL water (density) is:

3×10^{23} (b) 6.023×10^{24}

3×10^{25} (d) 6.023×10^{21}

[**Hint:** Mass = 18 g

Number of molecules of H_2O in 18 g mass $= 6.023 \times 10^{23}$

Number of electrons in 18 g water $= 6.023 \times 10^{23} \times 10$

$= 6.023 \times 10^{24}$

$\because$ Each molecule of water contains 10 electrons.]

6. What is the empirical formula of vanadium oxide if 2.74 g of metal oxide contains 1.53 g of metal?

(a) V_2O_3 (b) VO (c) V_2O_5 (d) V_2O_7

[Hint: % of V $= \frac{1.53}{2.74} \times 100 = 55.83$

$\therefore$ % of O = 44.17

Element	%	Atomic ratio	Simplest ratio
V	55.83	$\frac{55.83}{52} = 1.1$	$\frac{1.1}{1.1} = 1$
O	44.17	$\frac{44.17}{16} = 2.76$	$\frac{2.76}{1.1} = 2.5$

V : O = 2 : 5

Thus, empirical formula = V_2O_5]

7. Number of moles of electrons in 4.2 g of N^{3-} ion (nitride ion) is:

(a) 3 (b) 2 (c) 1.5 (d) 4.2

8. The ratio of volumes occupied by 1 mole O_2 and 1 mole CO_2 under identical conditions of temperature and pressure is:

(a) 1 : 1 (b) 1 : 2 (c) 1 : 3 (d) 2 : 1

9. The maximum amount of $BaSO_4$ that can be obtained on mixing 0.5 mole $BaCl_2$ with 1 mole H_2SO_4 is:

(a) 0.5 mol (b) 0.1 mol (c) 0.15 mol (d) 0.2 mol

[**Hint:** $H_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2HCl$

0.5 mole $BaCl_2$ will react with 0.5 mole H_2SO_4 to give 0.5 mole $BaSO_4$]

10. If 10^{21} molecules are removed from 100 mg CO_2, then number of moles of CO_2 left are:

(a) 6.10×10^{-4} (b) 2.8×10^{-3}

(c) 2.28×10^{-3} (d) 1.36×10^{-2}

[**Hint:** Number of molecules in 100 mg CO_2

$$= \frac{\text{Mass}}{\text{Molar mass}} \times 6.023 \times 10^{23}$$

$$= \frac{0.1}{44} \times 6.023 \times 10^{23}$$

$$= 1.368 \times 10^{21}$$

Molecules remaining $= 1.368 \times 10^{21} - 10^{21} = 0.368 \times 10^{21}$

Number of moles remaining $= \frac{0.368 \times 10^{21}}{6.023 \times 10^{23}} = 6.1 \times 10^{-4}$]

on one gram ion of Al^{3+} ion is:

: e coulomb (b) $\frac{1}{3} \times N_A \times e$ coulomb

e coulomb (d) $3 \times N_A \times e$ coulomb

: gram ion of Al^{3+} means one mole ion of Al^{3+}.
mole $Al^{3+} = 3 \times e \times N_A$ coulomb.]

nass of N_2O as well as CO_2 is 44 g mol^{-1}. At 25°C
essure, 1 L N_2O contains n molecules of gas. The
CO_2 molecules in 2 L under same conditions will

(b) $2n$ (c) $\frac{n}{2}$ (d) $\frac{n}{4}$

is dissolved in 1 L water. The number of ions of
$^-$ in 1 mL of this solution will be:

^{19} (b) 1.2×10^{22} (c) 1.2×10^{20} (d) 6.02×10^{20}

nber of moles of NaCl

$$= \frac{\text{Mass}}{\text{Molar mass}} = \frac{5.85}{58.5} = 0.1$$

ons ($Na^+ + Cl^-$) in 1 L

$$= 2 \times 0.1 \times 6.023 \times 10^{23}$$

$$= 12.046 \times 10^{22}$$

ons in 1 mL $= \frac{12.046 \times 10^{22}}{1000} = 1.2 \times 10^{20}$]

de has the formula X_2O_3. It can be reduced by
give free metal and water. 0.1596 g of metal oxide
g of hydrogen for complete reduction. The atomic
al in amu is:

(b) 155.8 (c) 5.58 (d) 55.8

$_2O_3$ + $3H_2 \longrightarrow 2X + 3H_2O$
nol 3 mol
48) g 6g

H_2 is required by 0.1596 g oxide
ill be required by 159.6 g oxide

$$2a + 48 = 159.6$$

$$a = 55.8$$

tomic mass of metal M.]

(PH_3) decomposes to produce vapours of
and H_2 gas. What will be the change in volume
L of phosphine is decomposed?

(b) 500 mL
(d) – 500 mL

$PH_3(g) \longrightarrow P_4(g) + 6H_2(g)$

4 mL $\longrightarrow$ 1 mL 6 mL

0 mL $\longrightarrow \frac{100}{4}$ $\frac{6}{4} \times 100$

0 mL $\longrightarrow$ 25 mL 150 mL

eases by 75 mL.]

neutron is assumed to half of its original value,
t of proton is assumed to be twice of its original
he atomic mass of $^{14}_{6}C$ will be:

[**Hint:** In the isotope $^{14}_{6}C$: Number of protons = 6
Number of neutrons = 8

New atomic mass will be $= 2 \times 6 + \frac{1}{2} \times 8 = 16$

% Increase in mass $= \frac{16 - 14}{14} \times 100 = 14.28\%$]

17. The mass and charge of 1 mole electrons will be:
(a) 1 kg; 96500 C (b) 0.55 mg; 96500 C
(c) 1.55 mg; 96500 C (d) 5.5 mg; 96500 C

18. The simplest formula of the compound containing 50% X (atomic mass 10 amu) and 50% Y (atomic mass 20 amu) is:
(a) XY_2 (b) X_2Y (c) X_2Y_3 (d) XY_3

[**Hint:**

Element	%	Atomic ratio	Simplest ratio
X	50	$\frac{50}{10} = 5$	$\frac{5}{2.5} = 2$
Y	50	$\frac{50}{20} = 2.5$	$\frac{2.5}{2.5} = 1$

Formula = X_2Y]

19. Rest mass of 1 mole neutrons ($m_n = 1.675 \times 10^{-27}$ kg) is:
(a) 1.8×10^{-3} kg (b) 1.008×10^{-4} kg
(c) 1.08×10^{-3} kg (d) 1.008×10^{-3} kg

[**Hint:** Mass of 1 mole neutrons

$$= 1.675 \times 10^{-27} \times 6.023 \times 10^{23}$$
$$= 1.008 \times 10^{-3} \text{ kg}]$$

20. Loschmidt number is the number of:
(a) molecules present in 1 mL of a gas at STP
(b) molecules present in 1 gram mole of a gas at STP
(c) atoms present in 1 mL of a gas at STP
(d) atoms present in 1 gram mole of a gas at STP

21. Which of the following statements is incorrect?
(a) One gram mole of silver equals $\frac{108}{6.023} \times 10^{-23}$ g
(b) One mole of CH_4 and 17 g of NH_3 at NTP occupy same volume
(c) One mole Ag weighs more than that of two moles of Ca
(d) One gram mole of CO_2 is 6.023×10^{23} times heavier than one molecule of CO_2

22. One atom of an element 'X' weighs 6.664×10^{-23} gm. The number of gram atoms in 40 kg of it is:
(a) 10 (b) 100 (c) 10000 (d) 1000

23. The density of a liquid is 1.2 g/mL. There are 35 drops in 2 mL. The number of molecules in 1 drop is (molecular weight of liquid = 70):
(a) $\frac{1.2}{35} N_A$ (b) $\left(\frac{1}{35}\right)^2 N_A$
(c) $\frac{1.2}{(35)^2} N_A$ (d) $1.2 N_A$

[**Hint:** Volume of one drop $= \frac{2}{35}$ mL

Number of moles in one drop $= \frac{2 \times 1.2}{35 \times 70} = \frac{1.2}{(35)^2}$

nolecules in one drop $= \frac{1.2}{(35)^2} \times N_A$]

ne of a liquid will contain 4 mole? Molar mass of 0 and its density is 1.4 g/mL:

(b) 1.6 L (c) 0.8 L (d) 4.8×10^{23} L

t x L liquid contain 4 mole of it.

ber of moles $= \frac{\text{Mass}}{\text{Molar mass}}$

$$4 = \frac{x \times 1000 \times 1.4}{280}$$

$$x = \frac{4 \times 280}{1.4 \times 1000} = 0.8 \text{ L}]$$

ratio of Fe^{2+} to Fe^{3+} in a mixture of $FeSO_4$ and having equal number of sulphate ions in both l ferric sulphates is:

(b) 3 : 2

(d) none of these

$SO_4 \longrightarrow Fe^{2+} + SO_4^{2-}$

$O_4)_3 \longrightarrow 2Fe^{3+} + 3SO_4^{2-}$

mole SO_4^{2-} ions are furnished by both $FeSO_4$ and

moles of $Fe^{2+} = x$

moles of $Fe^{3+} = \frac{2}{3}x$

$$Fe^{2+} : Fe^{3+} :: x : \frac{2}{3}x$$

$$= 3 : 2]$$

f electrons present in 3.6 mg of NH_4^+ are:

0^{21} (b) 1.2×10^{20} (c) 1.2×10^{22} (d) 2×10^{-3}

umber of electrons in one ion of $NH_4^+ = 10$

ions in 3.6 mg NH_4^+

$$= \frac{3.6 \times 10^{-3}}{18} \times 6.023 \times 10^{23} = 1.2 \times 10^{20}$$

ber o. electrons in 3.6 mg $NH_4^+ = 1.2 \times 10^{20} \times 10$

$$= 1.2 \times 10^{21}]$$

ction $4A + 2B + 3C \longrightarrow A_4B_2C_3$, what will be the moles of product formed, starting from one mole of le of B and 0.72 mole of C?

(b) 0.3 (c) 0.24 (d) 2.32

$\underset{\text{nol}}{A} + \underset{0.6\text{ mol}}{2B} + \underset{0.72\text{ mol}}{3C} \longrightarrow A_4B_2C_3$

sent case, reactant 'C' will be the limiting reactant will give least amount of product on being completely

gives 1 mol product,

'C' will give 0.24 mol of product.]

f $Na_2SO_4 \cdot nH_2O$ contains 12.6 gm of water. The ' is:

[**Hint:** $Na_2SO_4 \cdot nH_2O$

Molar mass = (142 + 18n)

Mass of water $= \frac{12.6}{26.8} \times (142 + 18n)$

$$18n = \frac{12.6}{26.8} \times (142 + 18n)$$

$$n = 7]$$

29. Consider the following data:

Element	Atomic weight
A	12.01
B	35.5

A and B combine to form a new substance X. If 4 moles of B combine with 1 mole of A to give 1 mole of X, then the weight of 1 mole of X is:

(a) 154 g (b) 74 g (c) 47.5 g (d) 160 g

30. How many moles of Na^+ ions are in 20 mL of 0.4 M Na_3PO_4?

(a) 0.008 (b) 0.024 (c) 0.05 (d) 0.20

[**Hint:** No. of moles of $Na_3PO_4 = \frac{MV}{1000} = \frac{0.4 \times 20}{1000}$

$= 0.008$

Number of moles of $Na^+ = 3 \times$ Number of moles of Na_3PO_4

$= 3 \times 0.008 = 0.024$]

31. The element whose one atom has mass of 10.86×10^{-26} kg is:

(a) boron (b) calcium (c) silver (d) zinc

32. An electric discharge is passed through a mixture containing 50 cc of O_2 and 50 cc of H_2. The volume of the gases formed (i) at room temperature, (ii) at 110°C will be:

(a) (i) 25 cc (ii) 50 cc (b) (i) 50 cc (ii) 75 cc

(c) (i) 25 cc (ii) 75 cc (d) (i) 75 cc (ii) 75 cc

[**Hint:** $2H_2(g) + O_2(g) \longrightarrow 2H_2O$

50 cc H_2 will combine with 25 cc O_2 to form 50 cc H_2O

$\therefore$ O_2 left = 25 cc

At room temperature, H_2O will be in liquid state but at 110°C, it will be gaseous. Thus, volume of gases at 25°C and 110°C will be 25 cc and 75 cc respectively.]

33. The mass of carbon present in 0.5 mole of $K_4[Fe(CN)_6]$ is:

(a) 1.8 g (b) 18 g (c) 3.6 g (d) 36 g

[**Hint:** 1 mole of $K_4[Fe(CN)_6]$ contains 6 mole carbon, *i.e.*, 72 g carbon.]

34. Caffeine has a molecular weight of 194. If it contains 28.9% by mass of nitrogen, number of atoms of nitrogen in one molecule of caffeine is:

(a) 4 (b) 6 (c) 2 (d) 3

[**Hint:** Mass of nitrogen in 194 amu caffeine $= \frac{28.9}{100} \times 194$

$= 56$ amu

$\therefore$ One molecule of caffeine will contain 4 atoms of nitrogen.]

35. Chlorine can be prepared by reacting HCl with MnO_2. The reaction is represented by the equation,

$$MnO_2(s) + 4HCl(aq.) \longrightarrow Cl_2(g) + MnCl_2(aq.) + 2H_2O(l)$$

Assuming that the reaction goes to completion, what mass of conc. HCl solution (36% HCl by mass) is needed to produce 2.5 g

36. What is the mass per cent of oxygen in $Al_2(SO_4)_3 \cdot 18H_2O$? The molar mass of this substance is 666.43 g/mol :
(a) 9.60 (b) 28.8 (c) 43.2 (d) 72

37. 0.25 g of an element 'M' reacts with excess fluorine to produce 0.547 g of the hexafluoride MF_6. What is the element?
(a) Cr (b) Mo (c) S (d) Te

38. How many electrons are present in 2×10^{-3} moles of $^{18}_{8}O^{2-}$?
(a) 1.2×10^{21} (b) 9.6×10^{21} (c) 1.2×10^{22} (d) 1.9×10^{22}

39. Fluorine reacts with uranium to form UF_6.
$$U(s) + 3F_2(g) \longrightarrow UF_6(g)$$
How many fluorine molecules are required to produce 2 mg of UF_6 from an excess of uranium? The molar mass of UF_6 is 352 g mol^{-1}.
(a) 3.4×10^{18} (b) 1×10^{19} (c) 2×10^{19} (d) 3.4×10^{21}

40. What is the formula of a substance with mass percentages of 35.79% for S, 62.92% for O and 1.13% for H?
(a) H_2SO_3 (b) H_2SO_4 (c) $H_2S_2O_7$ (d) $H_2S_2O_8$

41. In 1811, Avogadro calculated the formula of camphor by means of elemental chemical analysis and by measuring the density of its vapour. Avogadro found the density to be 3.84 g/L when he made the measurement at 210°C at 1 atm pressure. Which of the following is the correct formula of camphor?
(a) $C_{10}H_{14}O$ (b) $C_{10}H_{16}O$ (c) $C_{10}H_{16}O_2$ (d) $C_{10}H_{18}O$
(e) None of these
[**Hint:** $Pm = dRT$
$$m = \frac{dRT}{P} = \frac{3.84 \times 0.0821 \times 483}{1} = 152.27$$
$\therefore C_{10}H_{16}O$ will be the correct formula.]

42. A quantity of aluminium has a mass of 54 g. What is the mass of same number of magnesium atoms?
(a) 12.1 g (b) 24.3 g (c) 48.6 g (d) 97.2 g

43. When 1 L of CO_2 is heated with graphite, the volume of the gases collected is 1.5 L. Calculate the number of moles of CO produced at STP:
(a) $\frac{1}{11.2}$ (b) $\frac{28}{22.4}$ (c) $\frac{1}{22.4}$ (d) $\frac{14}{22.4}$
[**Hint:** $\underset{1-x}{CO_2(g)} + C(s) \longrightarrow \underset{2x}{2CO(g)}$
$$\text{Total volume} = 1 - x + 2x = 1 + x = 1.5$$
$$x = 0.5 \text{ L}$$
$\therefore$ Volume of CO $= 2 \times 0.5 = 1$ L
Number of moles of CO $= \frac{1}{22.4}$]

44. Which of the following has greatest number of atoms?
(a) 1 g of butane (C_4H_{10}) (b) 1 g of nitrogen (N_2)
(c) 1 g of silver (Ag) (d) 1 g of water (H_2O)

45. A metal oxide has the formula M_2O_3. It can be reduced by H_2 to give free metal and water. 0.1596 g of M_2O_3 required 6 mg of H_2 for complete reduction. The atomic mass of the metal is:
(a) 27.9 (b) 79.8
(c) 55.8 (d) 159.8
[**Hint:** $\underset{(2x+48)\text{ g}}{M_2O_3} + \underset{6\text{ g}}{3H_2} \longrightarrow 2M + 3H_2O$
x = Atomic mass of metal
$\because$ 0.006 g H_2 reduces 0.1596 g M_2O_3
$\therefore$ 6 g H_2 will reduce $\frac{0.1596}{0.006} \times 6 \text{ g } M_2O_3 = 159.6 \; M_2O_3$
$$2x + 48 = 159.6$$
$$2x = 111.6$$
$$x = 55.8]$$

46. In a compound of molecular formula A_mB_n:
(a) number of equivalents of A, B and A_mB_n are same
(b) number of moles of A, B and A_mB_n are same
(c) $m \times$ moles of $A = n \times$ moles of $B = (m+n) \times$ moles of A_mB_n
(d) $n \times$ moles of $A = m \times$ moles of $B = (m+n) \times$ moles of A_mB_n

47. 4.4 g of CO_2 and 2.24 litre of H_2 at STP are mixed in a container. The total number of molecules present in the container will be:
(a) 6.022×10^{23} (b) 1.2044×10^{23}
(c) 6.023×10^{26} (d) 6.023×10^{24}

48. A partially dried clay mineral contains 8% water. The original sample contained 12% water and 45% silica. The % of silica in the partially dried sample is nearly:
(a) 50% (b) 49%
(c) 55% (d) 47%
[**Hint:**

	Clay	Silica	Water
Initial stage:	43%	45%	12%
Final stage:	$(92 - x)$	x	8%

Ratio of silica and clay will remain constant, before and after drying.
$$\therefore \quad \frac{45}{43} = \frac{x}{92 - x}$$
$$x = 47\%]$$

49. Which of the following is isomorphous with $MgSO_4 \cdot 7H_2O$?
(a) Green vitriol (b) Blue vitriol
(c) Red vitriol (d) Vitriol of mass

50. In the reaction;
$$CO + \frac{1}{2}O_2 \longrightarrow CO_2; \quad N_2 + O_2 \longrightarrow 2NO$$
10 mL of mixture containing carbon monoxide and nitrogen required 7 mL oxygen to form CO_2 and NO, on combustion. The volume of N_2 in the mixture will be:
(a) 7/2 mL (b) 17/2 mL (c) 4 mL (d) 7 mL
[**Hint:** $\underset{x}{CO(g)} + \underset{x/2}{\frac{1}{2}O_2(g)} \longrightarrow CO_2(g)$
$$\underset{y}{N_2(g)} + \underset{y}{O_2(g)} \longrightarrow 2NO(g)$$
$$x + y = 10 \qquad \ldots (i)$$
$$\frac{x}{2} + y = 7 \qquad \ldots (ii)$$
Solving eqs. (i) and (ii),
$$x = 6 \text{ and } y = 4]$$

51. 1.44 g of titanium (Ti) reacted with excess of O_2 and produced x gm of a nonstoichiometric compound $Ti_{1.44}O_1$. The value of x is:

(b) 1.77
(d) None of these

he reaction:

$$Ti + O_2 \longrightarrow Ti_{1.44}O_1$$

noles of titanium = Number of moles of $Ti_{1.44}O_1$

$$\frac{1.44}{48} = \frac{x}{48 \times 1.44 + 16}$$

$x = 1.77$ g]

e of 75% alcohol by mass ($d = 0.8$ g/cm^3) must be
pare 150 cc of 30% alcohol by mass ($d = 0.9$

L (b) 56.25 mL
(d) 33.56 mL

V mL of alcohol was used.

$$\frac{75}{100} \times V \times 0.8 = \frac{30}{100} \times 150 \times 0.9$$

$V = 67.5$ mL]

Following questions may have more than one correct options:

1. 11.2 L of a gas at STP weighs 14 g. The gas could be:
(a) N_2O (b) NO_2 (c) N_2 (d) CO
2. In which of the following pairs do 1 g of each have an equal number of molecules?
(a) N_2O and CO (b) N_2 and C_3O_2
(c) N_2 and CO (d) N_2O and CO_2
3. 8 g of oxygen has the same number of molecules as in:
(a) 11 g CO_2 (b) 22 g CO_2 (c) 7 g CO (d) 14 g CO
4. Which of the following has three significant figures?
(a) 6.02×10^{23} (b) 0.25
(c) 6.60×10^{-34} (d) 1.50
5. 1 mole of ${}^{14}_{7}N^{3-}$ ions contains:
(a) $7 \times 6.023 \times 10^{23}$ electrons (b) $7 \times 6.023 \times 10^{23}$ protons
(c) $7 \times 6.023 \times 10^{23}$ neutrons (d) $14 \times 6.023 \times 10^{23}$ protons
6. 1 g atom of nitrogen represents:
(a) 14 g nitrogen
(b) 11.2 litre of N_2 at NTP
(c) 22.4 litre of N_2 at NTP
(d) 6.023×10^{23} molecules of N_2

wers

ect option

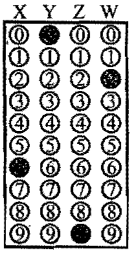
2. (b)	3. (a)	4. (d)	5. (b)	6. (c)	7. (a)	8. (a)
10. (a)	11. (d)	12. (b)	13. (c)	14. (d)	15. (c)	16. (b)
18. (b)	19. (d)	20. (a)	21. (a)	22. (d)	23. (c)	24. (c)
26. (a)	27. (c)	28. (d)	29. (a)	30. (b)	31. (d)	32. (c)
34. (a)	35. (b)	36. (d)	37. (b)	38. (c)	39. (b)	40. (c)
42. (c)	43. (c)	44. (a)	45. (c)	46. (a)	47. (b)	48. (d)
50. (c)	51. (b)	52. (c)				

e than one correct options

2. (c, d)	3. (a, c)	4. (a, c, d)	5. (b, c)	6. (a)

Integer Answer TYPE QUESTIONS

o each of the following questions is a teger, ranging from 0 to 9. If the correct estion numbers X, Y, Z and W (say) are respectively, then the correct darkening ll look like the given figure.



ous alkane (C_nH_{2n+2}) is exploded with oxygen. The e of O_2 used and CO_2 formed are in the ratio of 7 : 4. e the value of *n*.

$$C_nH_{2n+2} + \left[n + \frac{n+1}{2}\right]O_2 \longrightarrow nCO_2 + (n+1)\,H_2O$$

$$\frac{n + \frac{(n+1)}{2}}{n} = \frac{7}{4}$$

$$n = 2\,]$$

nany atoms do a mercury vapour molecule consist of, if pour density of mercury vapour relative to air is 6.92? ic mass of mercury is 200). The average molar mass of 9 g/mol.

$$\frac{\text{Vapour density of Hg vapour}}{\text{Vapour density of air}} = \frac{\text{Molar mass of Hg}}{\text{Molar mass of Air}}$$

$$\frac{6.92}{1} = \frac{m}{29}$$

$$m = 200 \text{ g/mol}$$

mass is same as that of atomic mass hence mercury vapour ns monoatomic mercury.]

nole of an element contains 4.2×10^{24} electrons. What is omic number of the element?

romolecule of iron has molar mass 2800 amu, it contains n by mass. The number of iron atom in one formula unit macromolecule is:

: Number of iron atoms in one formula unit of compound

$$= \frac{\%}{100} \times \frac{\text{Molecular mass}}{\text{Atomic mass}} = \frac{8}{100} \times \frac{2800}{56} = 4\,]$$

es of '*A*' and 10 moles of '*B*' are mixed and allowed to ccording to the equation :

$$A + 3B \longrightarrow 2C$$

nany moles of *C* are present when there are 4 moles of *A* container?

nany water molecules will be there in 3×10^{-23} g sample er?

7. 5 g H_2 is allowed to react with 14 g N_2 for the following reaction:

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

What mass of H_2 will be left unreacted at the end of reaction?

[**Hint :** N_2 is limiting reactant, thus 14 g N_2 will give 17 g NH_3 and *x*g H_2 remains unreacted.

Mass before reaction = Mass after reaction

$$(5+14) = (17+x)$$

$$x = 2\text{g}\,]$$

8. Calculate the number of moles of water in 976 g $BaCl_2 \cdot 2H_2O$.

9. If Avogadro's number be 3.01×10^{23} then the atomic mass of carbon will be:

10. How many moles of *R* will be produced when 8 mol of *P* and 5 mol of *Q* are allowed to react according to the equation :

$$2P + Q \longrightarrow R$$

11. The mass of 1×10^{22} molecules of blue vitriol ($CuSO_4 \cdot xH_2O$) is 4.144 g. The value of '*x*' will be:

12. What will be the mass (in kg) of 7.298×10^6 mol electrons?

13. Silver (Atomic weight = 108 g mol^{-1}) has a density of 10.5 g cm^{-3}. The number of silver atoms on a surface area of $10^{-12}m^2$ can be expressed in scientific notation as $y \times 10^x$. The value of *x* is: **(IIT 2010)**

[**Hint :** Mass of 1 cm^3 Ag = 1 × 10.5 g

$$\text{Number of atoms} = \frac{10.5}{108} \times 6.023 \times 10^{23}$$

$$\text{Number of atoms in 1 cm} = \left[\frac{10.5}{108} \times 6.023 \times 10^{23}\right]^{1/3}$$

$$\text{Number of atoms in 1 cm}^2 = \left[\frac{10.5}{108} \times 6.023 \times 10^{23}\right]^{2/3}$$

Number of atoms in $10^{-2}m^2$ or 10^{-8} cm^2

$$= \left[\frac{10.5}{108} \times 6.023 \times 10^{23}\right]^{2/3} \times 10^{-8} = 1.5 \times 10^7\,]$$

14. A student performs a titration with different burettes and finds titre values of 25.2 mL, 25.25 mL and 25.0 mL. The number of significant figures in average titre value is: **(IIT 2010)**

[**Hint :** Average titre value $= \dfrac{25.2 + 25.25 + 25.0}{3}$

$$= \frac{75.45}{3} = \frac{75.4}{3} = 25.1$$

(In addition, result is reported upto least place of decimal)]

nswers

2. (1) 3. (7) 4. (4) 5. (4) 6. (1) 7. (2) 8. (8)

10. (4) 11. (5) 12. (4) 13. (7) 14. (3)

LINKED COMPREHENSION TYPE QUESTIONS

1

y, 'mole' is an essential tool for the chemical t is a basic SI unit adopted by the 14th general weights and measurements in 1971. A mole contains ntary particles as the number of atoms present in 12 g of a gas at STP occupies 22.4 litre volume. Molar ids and liquids is not definite. Molar mass of a o called gram-atomic mass or gram molar mass. The g of mole is plenty, heap or the collection of large le of a substance contains 6.023×10^{23} elementary om or molecule. Atomic mass unit (amu) is the unit of g., atomic mass of single carbon is 12 amu.

lowing questions:

s of one amu is approximately:
(b) 0.5 g
10^{-24} g (d) 3.2×10^{-24} g

of a gas at STP are found to have a mass of 22 g. The r mass of the gas is:
(b) 44 (c) 88 (d) 33

s of one molecule of water is approximately:
(b) 0.5 g
10^{-24} g (d) 3.2×10^{-23} g

ny atoms are present in 49 g of H_2SO_4?
023×10^{23} (b) $5\times6.023\times10^{23}$
023×10^{23} (d) $7\times3.02\times10^{23}$

s at STP contains 3×10^{22} molecules. The number of s in x L ozone at STP will be:
22 (b) 4×10^{23} (c) 6.02×10^{23} (d) 3×10^{24}

lro's number is 1×10^{23} mol^{-1} then the mass of one xygen would be:
nu (b) 16×6.02 amu
ı (d) 16×10^{-23} amu

the Avogadro's number then number of valence n 4.8 g of O^{2-} is:
(b) 4.2 N_A (c) 1.6 N_A (d) 3.2 N_A

2

he atoms of same element; they have same atomic rent mass numbers. Isotopes have different number ieir nucleus. If an element exists in two isotopes asses 'a' and 'b' in the ratio $m:n$, then average be $\frac{m\times a+n\times b}{m+n}$.

opes of same element have same position in the he elements which have single isotope are called nents. Greater is the percentage composition of an

Answer the following questions:

1. The isotopes of chlorine with mass number 35 and 37 exist in the ratio of..... Its average atomic mass is 35.5.
(a) 1:1 (b) 2:1 (c) 3:1 (d) 3:2
2. Which of the following isotopes is/are used to decide the scale of atomic mass?
(a) ${}^{12}_{6}C$ (b) ${}^{14}_{6}C$ (c) ${}^{16}_{8}O$ (d) ${}^{14}_{7}N$
3. Atomic mass of boron is 10.81. It has two isotopes namely ${}^{11}_{5}B$ and ${}^{x}_{5}B$ with their relative abundance of 80% and 20% respectively. The value of x is:
(a) 10.05 (b) 10 (c) 10.01 (d) 10.02
4. The ratio of the mass of ${}^{12}C$ atom to that of an atom of element X (whose atomicity is four) is 1:9. The molecular mass of element X is:
(a) 480 g mol^{-1} (b) 432 g mol^{-1}
(c) 36 g mol^{-1} (d) 84 g mol^{-1}
5. ${}^{12}C$ and ${}^{14}C$ isotopes are found as 98% and 2% respectively in any sample. Then, the number of ${}^{14}C$ atoms in 12 g of the sample will be:
(a) 1.5 mole atoms (b) 1.032×10^{22} atoms
(c) 2.06×10^{21} atoms (d) 2 g atoms

• Passage 3

Empirical formula is the simplest formula of the compound which gives the atomic ratio of various elements present in one molecule of the compound. However, the molecular formula of the compound gives the number of atoms of various elements present in one molecule of the compound.

$$\text{Molecular formula} = (\text{Empirical formula})\times n$$

$$n=\frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

A compound may have same empirical and molecular formulae. Both these formulae are calculated by using percentage composition of constituent elements.

Answer the following questions:

1. Two metallic oxides contain 27.6% and 30% oxygen respectively. If the formula of first oxide is M_3O_4, that of second will be:
(a) MO (b) MO_2 (c) M_2O_5 (d) M_2O_3
2. Which of the following compounds have same empirical formula?
(a) Formaldehyde (b) Glucose
(c) Sucrose (d) Acetic acid
3. Which of the following represents the formula of a substance which contains 50% oxygen?
(a) N_2O (b) CO_2
(c) NO_2 (d) CH_3OH
4. An oxide of iodine (I = 127) contains 25.4 g of iodine and 8 g of oxygen. Its formula could be:
(a) I_2O_3 (b) I_2O
(c) I_2O_5 (d) I_2O_7

lrofluoric acid gas occupies 5.6 litres of volume at e empirical formula of the gas is HF, then its formula in the gaseous state will be:

(b) H_2F_2 (c) H_3F_3 (d) H_4F_4

species having different percentage composition of :

)OH and $C_6H_{12}O_6$ (b) CH_3COOH and C_2H_5OH
CH_3 and HCOOH (d) C_2H_5OH and CH_3OCH_3

nd of Na, C and O contains 0.0887 mol Na, 0.132 2.65×10^{22} atoms of carbon. The empirical formula ıpound is:

(b) $Na_3C_5O_2$
)$_3$ (d) $Na_{0.0887}C_{2.65 \times 10^{22}}O_{0.132}$

4

loids are extracted from the extracts of the plants na. Marijuana owes its activity to tetrahydro ıich contains 70% as many as carbon atoms as s and 15 times as many hydrogen atoms as oxygen m of tetrahydro cannabinol is 0.00318.

lowing questions:

ır mass of the compound is:
mu (b) 314 amu (c) 143 amu (d) 341 amu

ar formula of the compound is:
$_{30}O_2$ (b) $C_{21}H_{14}O_3$
$_{46}O$ (d) none of these

of oxygen atoms in 1 mol of the tetrahydro ıol is:
(b) N_A (c) $3\ N_A$ (d) $4\ N_A$
$_A = 6.023 \times 10^{23}$

ge composition of carbon in the compound is:
% (b) 70.85% (c) 80.25% (d) 59.64%

5

sity of a compound is defined as the ratio of mass of a of gas to the mass of the same volume of hydrogen tical conditions of temperature and pressure.

$$= \frac{\textit{Mass of certain volume of gas (22.4 L) at STP}}{\textit{Mass of same volume of } H_2 \textit{ gas (22.4 L) at STP}}$$

$$= \frac{Mw}{2}$$

ɔlecular mass of gas = Vapour density × 2

is a unitless quantity; it is unaffected by variation of nd pressure.

llowing questions:

density of a metal chloride is 66. Its oxide contains tal. The atomic mass of the metal is:
(b) 54 (c) 27.06 (d) 2.086

Number of equivalents of metal = Number of equivalents of oxygen

$$\frac{53}{E} = \frac{47}{8}$$

$$E = 9.02$$

Molecular formula of metal chloride = MCl_n

Molecular mass = $[n \times 9.02 + n \times 35.5] = 132$

$\therefore \quad n = 3$

$\therefore$ Atomic mass of metal = $3 \times 9.02 = 27.06$]

2. The vapour density of a mixture containing NO_2 and N_2O_4 is 38.3 at 27°C. The moles of NO_2 in 100 moles of mixture are:
(a) 33.48 (b) 53.52 (c) 38.3 (d) 76.6
3. At STP, 5.6 litre of a gas weighs 60 g. The vapour density of gas is:
(a) 60 (b) 120 (c) 30 (d) 240
4. Which of the following two substances have same vapour density?
(a) Glucose (b) Fructose (c) Sucrose (d) Starch
5. Let $NH_4HS(s)$ is heated in a closed vessel to decompose.

$$NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$$

The vapour density of the mixture will be:
(a) equal to that of NH_4HS
(b) lesser than that of NH_4HS
(c) greater than that of NH_4HS
(d) cannot be predicted

• Passage 6

Precision refers to the closeness of a set of values obtained for identical measurement of a quantity. Precision depends on the limitations of measuring devices and the skills with which it is used. However, accuracy refers to the closeness of a single measurement to its true value.

The digits in a properly recorded measurement are known as significant figures. These are meaningful digits in a measured or calculated quantity. The greater the number of significant figures in a reported result, smaller is the uncertainty and greater is the precision. The zeros at the beginning are not counted. The zeros to the right of a decimal point are counted. In the numbers that do not contain a decimal point, "trailing" zeros may or may not be significant. The purpose of zeros at the end of a number is to convey the correct range of uncertainty.

Answer the following questions:

1. If repeated measurements give values close to one another, the number is:
(a) surely precise (b) surely accurate
(c) surely precise and accurate (d) all of these are correct
2. The number of significant figures in a measured number contains how many uncertain number of digits?
(a) Zero (b) 1
(c) 2 (d) Cannot be predicted
3. In the number 2.4560, there are 5 significant digits. Which one is the least significant digit?
(a) 2 (b) 4 (c) 0 (d) 6
4. If we add 296.2 and 2.256, we get the answer as 298.456 g. The number of significant figures in the result are:
(a) 6 (b) 5 (c) 4 (d) 3
5. In which of the following numbers, all the zeros are not significant?
(a) 0.0010 (b) 0.00100 (c) 0.001000 (d) 0.001

Answers

Passage 1.	**1.** (c)	**2.** (c)	**3.** (d)	**4.** (d)	**5.** (a)	**6.** (c)	**7.** (a)
Passage 2.	**1.** (c)	**2.** (a, c)	**3.** (b)	**4.** (b)	**5.** (b)		
Passage 3.	**1.** (d)	**2.** (a, b)	**3.** (d)	**4.** (c)	**5.** (b)	**6.** (b, c)	**7.** (a)
Passage 4.	**1.** (b)	**2.** (a)	**3.** (a)	**4.** (c)			
Passage 5.	**1.** (c)	**2.** (c)	**3.** (b)	**4.** (a, b)	**5.** (b)		
Passage 6.	**1.** (a)	**2.** (b)	**3.** (c)	**4.** (c)	**5.** (d)		

SELF ASSESSMENT

ASSIGNMENT NO. 1

SECTION-I

Straight Objective Type Questions

This section contains 10 multiple choice questions. Each question has 4 choices (a), (b), (c) and (d), out of which only one is correct.

1. Cartisone is a molecular substance containing 21 atoms of carbon per molecule. The mass percentage of carbon in cartisone is 69.98%. What is the molecular mass of cartisone?
(a) 360.4 (b) 176.5 (c) 287.6 (d) 312.8

2. Total number of atoms present in 25 mg of camphor, $C_{10}H_{16}O$ is:
(a) 2.57×10^{21} (b) 9.89×10^{19}
(c) 2.67×10^{21} (d) 6.02×10^{20}

3. The oxide of a metal contains 60% of the metal. What will be the percentage of bromine in the bromide of the metal, if the valency of the metal is the same in both, the oxide and the bromide?
(a) 93% (b) 87% (c) 70% (d) 77%

4. The radius of water molecule having density 1 g mL^{-1} is:
(a) 1.925 Å (b) 73.46 Å (c) 19.25 Å (d) 7.346 Å

5. 3 g of an oxide of a metal is converted completely to 5 g chloride. Equivalent mass of metal is:
(a) 33.25 (b) 3.325 (c) 12 (d) 20

6. Quantitative analysis of a compound shows that it contains 0.110 mole of 'C', 0.055 mole of 'N' and 0.165 mole of 'O'. Its molecular mass is about 270. How many atoms of carbon are there in empirical and molecular formulae of the compound respectively?

	Empirical formula	Molecular formula
(a)	1	3
(b)	2	2
(c)	2	6
(d)	3	2

7. Total number of electrons present in 11.2 L of NH_3 at STP is:
(a) 6.02×10^{23} (b) 3.01×10^{23}
(c) 3.01×10^{24} (d) 5.1×10^{24}

8. Which one of the following is not a unit of length?
(a) Angstrom (b) Light-year
(c) Micron (d) Radian

9. Unit of J pa^{-1} is equivalent to:
(a) m^3 (b) cm^3
(c) dm^3 (d) none of these

10. The relative abundance of two isotopes of atomic masses 85 and 87 are 75% and 25% respectively. The average atomic mass of element is:
(a) 86 (b) 40 (c) 85.5 (d) 75.5

SECTION-II

Multiple Answers Type Objective Questions

11. Mass of one atom of oxygen is/are:
(a) 16 amu (b) 32 amu
(c) 16 gm (d) 2.656×10^{-23} gm

12. Which of the following compounds have same percentage composition of carbon?
(a) $C_6H_{12}O_6$ (b) CH_3COOH
(c) $HCOOCH_3$ (d) $C_{12}H_{22}O_{11}$

13. Which of the following is/are correct about 1 mole electrons?
(a) 6.023×10^{23} electrons (b) 5.48×10^{-7} kg
(c) 96500 coulomb charge (d) None of these

14. In which of the following numbers, all zeros are significant?
(a) 5.0005 (b) 0.0030 (c) 30.000 (d) 0.5200

15. Which of the following are correct SI units?
(a) Amount of substance in mol L^{-1}
(b) Pressure of gas in pascal
(c) Density of a solid in kg m^{-3}
(d) Force in newton

SECTION-III

Assertion-Reason Type Questions

This section contains 4 questions. Each question contains **Statement-1** (Assertion) and **Statement-2** (Reason). Each question has following 4 choices (a), (b), (c) and (d), out of which only one is correct.

(a) Statement-1 is true; statement-2 is true; statement-2 is a correct explanation for statement-1.
(b) Statement-1 is true; statement-2 is true; statement-2 is not a correct explanation for statement-1.
(c) Statement-1 is true; statement-2 is false.
(d) Statement-1 is false; statement-2 is true.

16. **Statement-1:** Avogadro's number is a dimensionless quantity.

Because

Statement-2: It is a number of atoms or molecules in one gram mole.

17. **Statement-1:** An element has variable equivalent mass.

Because

Statement-2: The valency of element is variable.

18. **Statement-1:** Vapour density of CH_4 is half of O_2.

Because

Statement-2: 1.6 g of CH_4 contains same number of electrons as 3.2 g of O_2.

19. **Statement-1:** Specific gravity is dimensionless quantity.

Because

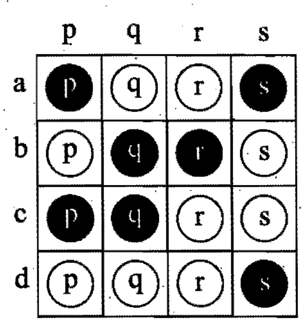
Statement-2: Specific gravity is relative density of a substance, measured with respect to density of water at 4°C.

SECTION-IV

Matrix-Matching Type Questions

This section contains 3 questions. Each question contains statement given in two columns which have to be matched. Statements (a, b, c and d) in Column-I have to be matched with statements (p, q, r and s) in Column-II. The answers to these questions have to be appropriately bubbled as illustrated in the following examples:

If the correct matches are (a-p,s); (b-q,r); (c-p,q) and (d-s); then the correctly bubbled 4 × 4 matrix should be as follows:



20. Match the Column-I with Column-II:

Column-I	Column-II
(a) N_2	(p) 40% carbon by mass
(b) CO	(q) Empirical formula CH_2O
(c) $C_6H_{12}O_6$	(r) Vapour density = 14
(d) CH_3COOH	(s) $14N_A$ ($N_A = 6.023 \times 10^{23}$) electrons in a mole

21. Match the Column-I with Column-II:

Column-I	Column-II
(a) 1 L	(p) 10^{-5} N
(b) 1 J	(q) 0.2389 cal
(c) 9.9×10^6 erg	(r) 10^{-3} m^3
(d) 1 Dyne	(s) 6.25×10^{18} eV

22. Match the Column-I with Column-II:

Column-I	Column-II
(a) 1 g mole of $O_2(g)$	(p) mass, 32 g
(b) 0.5 mole of $SO_2(g)$	(q) mass, 24 g
(c) 1 g of $H_2(g)$	(r) volume, 11.2 L at STP
(d) 0.5 mole of $O_3(g)$	(s) $1.5 \times 6.023 \times 10^{23}$ atoms

Answers

1. (a) **2.** (c) **3.** (b) **4.** (a) **5.** (a) **6.** (c) **7.** (c) **8.** (d)
9. (a) **10.** (c) **11.** (a, d) **12.** (a, b, c) **13.** (a, b, c) **14.** (a, c) **15.** (b, c, d) **16.** (a)
17. (b) **18.** (c) **19.** (a) **20.** (a-r, s) (b-r, s) (c-p, q) (d-p, q) **21.** (a-r) (b-q, s) (c-q) (d-p)
22. (a-p) (b-p, r, s) (c-r) (d-q, s)

CHAPTER 2

ATOMIC STRUCTURE

2.1 INTRODUCTION

The word **atom** is a Greek word meaning indivisible, *i.e.*, an ultimate particle which cannot be further subdivided. The idea that all matter ultimately consists of extremely small particles was conceived by ancient Indian and Greek philosophers. The old concept was put on firm footing by **John Dalton** in the form of **atomic theory** which he developed in the years 1803–1808. This theory was a landmark in the history of chemistry. According to this theory, atom is the smallest indivisible part of matter which takes part in chemical reactions. Atom is neither created nor destroyed. Atoms of the same element are similar in size, mass and characteristics; however, atoms of different elements have different size, mass and characteristics.

In 1833, Michael Faraday showed that there is a relationship between matter and electricity. This was the first major breakthrough to suggest that atom was not a simple indivisible particle of all matter but was made up of smaller particles. Discovery of **electrons, protons** and **neutrons** discarded the indivisible nature of the atom proposed by John Dalton.

The complexity of the atom was further revealed when the following discoveries were made in subsequent years:

(i) Discovery of cathode rays.
(ii) Discovery of positive rays.
(iii) Discovery of X-rays.
(iv) Discovery of radioactivity.
(v) Discovery of isotopes and isobars.
(vi) Discovery of quarks and the new atomic model.

During the past 100 years, scientists have made contributions which helped in the development of modern theory of atomic structure. The works of **J.J. Thomson** and **Ernest Rutherford** actually laid the foundation of the modern picture of the atom. It is now believed that the atom consists of several particles called **subatomic particles** like electron, proton, neutron, positron, neutrino, meson, etc. Out of these particles, the electron, the proton and the neutron are called **fundamental particles** and are the building blocks of the atoms.

2.2 CATHODE RAYS—DISCOVERY OF ELECTRON

The nature and existence of electron was established by experiments on conduction of electricity through gases. In 1859, **Julius Plucker** started the study of conduction of electricity through gases at low pressure in a discharge tube. [A common discharge tube consists of a hard glass cylindrical tube (about 50 cm long) with two metal electrodes sealed on both the ends. It is connected to a side tube through which it can be evacuated to any desired pressure with the help of a vacuum pump.] Air was almost completely removed from the discharge tube (pressure about 10^{-4} atmosphere). When a high voltage of the order of 10,000 volts or more was impressed across the electrodes, some sort of invisible rays moved from the negative electrode to the positive electrode (Fig. 2.1). Since, the negative electrode is referred to as cathode, these rays were called **cathode rays.**

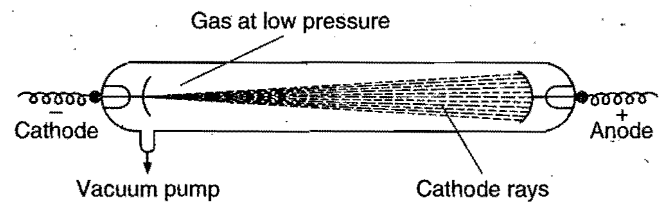


Fig. 2.1 Production of cathode rays

Further investigations were made by W. Crookes, J. Perrin, J.J. Thomson and others. Cathode rays possess the following properties:

(i) They travel in straight lines away from the cathode with very high velocities ranging from $10^9 - 10^{11}$ cm per second. A shadow of metallic object placed in the path is cast on the wall opposite to the cathode.

(ii) They produce a green glow when strike the glass wall beyond the anode. Light is emitted when they strike the zinc sulphide screen.

(iii) They produce heat energy when they collide with the matter. It shows that cathode rays possess kinetic energy which is converted into heat energy when stopped by matter.

(iv) They are deflected by the electric and magnetic fields. When the rays are passed between two electrically charged plates, these are deflected towards the positively charged plate. They discharge a positively charged gold leaf electroscope. It shows that **cathode rays carry negative charge.**

(v) They possess kinetic energy. It is shown by the experiment that when a small pin wheel is placed in their path, the blades of the wheel are set in motion. Thus, the **cathode rays consist of material particles which have mass and velocity.**

These particles carrying negative charge were called **negatrons** by Thomson.

The name 'negatron' was changed to 'electron' by **Stoney**.

(vi) Cathode rays produce X-rays. When these rays fall on a material having high atomic mass, new type of penetrating rays of very small wavelength are emitted which are called X-rays.

(vii) These rays affect the photographic plate.

(viii) These rays can penetrate through thin foils of solid materials and cause ionisation in gases through which they pass.

(ix) The nature of the cathode rays is independent of:

(a) the nature of the cathode and

(b) the gas in the discharge tube.

In 1897, **J. J. Thomson** determined the e/m value (charge/mass) of the electron by studying the deflections of cathode rays in electric and magnetic fields. The value of e/m has been found to be -1.7588×10^8 coulomb/g.

[The path of an electron in an electric field is parabolic, given as:

$$y = \frac{eE}{2mv^2}x^2$$

where, y = deflection in the path of electron in y-direction
e = charge on electron
E = intensity of applied electric field
m = mass of electron
v = velocity of electron
x = distance between two parallel electric plates within which electron is moving.

The path of an electron in a magnetic field is circular with radius 'r' given as:

$$r = \frac{mv}{eB}$$

where, m = mass of electron
v = velocity of electron
e = charge on electron
B = intensity of applied magnetic field

The radius of the path is proportional to momentum.]

By performing a series of experiments, **Thomson** proved that **whatever be the gas taken in the discharge tube and whatever be the material of the electrodes, the value of e/m is always the same.** Electrons are thus common universal constituents of all atoms.

J.J. Thomson gave following relation to calculate charge/mass ratio:

$$\frac{e}{m} = \frac{E}{rB^2}$$

where the terms have usual significance given before

$$= -1.7588 \times 10^{11}\ \text{C kg}^{-1}$$

Electrons are also produced by the action of ultraviolet light or X-rays on a metal and from heated filaments. β-particles emitted by radioactive materials are also electrons.

The first precise measurement of the charge on an electron was made by **Robert A. Millikan** in 1909 by oil drop experiment. The charge on the electron was found to be -1.6022×10^{-19} coulomb. **Since, an electron has the smallest charge known, it was, thus,** designated as unit negative charge.

Mass of the electron: The mass of the electron can be calculated from the value of e/m and the value of e.

$$m = \frac{e}{e/m} = \frac{-1.6022 \times 10^{-19}}{-1.7588 \times 10^8}$$

$$= 9.1096 \times 10^{-28}\ \text{g or } 9.1096 \times 10^{-31}\ \text{kg}$$

This is termed as the rest mass of the electron, *i.e.*, mass of the electron when moving with low speed. The mass of a moving electron may be calculated by applying the following formula:

$$\text{Mass of moving electron} = \frac{\text{rest mass of electron}}{\sqrt{1 - \left(\frac{v}{c}\right)^2}}$$

where v is the velocity of the electron and c is the velocity of light. When v becomes equal to c, mass of the moving electron becomes infinity and when the velocity of the electron becomes greater than c, mass of the electron becomes imaginary.

Mass of the electron relative to that of hydrogen atom:

Mass of hydrogen atom = 1.008 amu

$$= 1.008 \times 1.66 \times 10^{-24}\ \text{g} \quad (\text{since } 1\ \text{amu} = 1.66 \times 10^{-24}\ \text{g})$$

$$= 1.673 \times 10^{-24}\ \text{g}$$

$$\frac{\text{Mass of hydrogen atom}}{\text{Mass of the electron}} = \frac{1.673 \times 10^{-24}}{9.1096 \times 10^{-28}}$$

$$= 1837$$

$$\text{Thus, Mass of an electron} = \frac{1}{1837} \times \text{mass of hydrogen atom}$$

$$= \frac{1.008}{1837}$$

$$= 0.000549\ \text{amu}$$

An electron can, thus, be defined as a subatomic particle which carries charge -1.60×10^{-19} coulomb, *i.e.*, one unit negative charge and has mass 9.1×10^{-28} g, *i.e.*, $\frac{1}{1837}$th mass of the hydrogen atom (0.000549 amu).

[Millikan's oil drop method is used to determine the charge on an electron by measuring the terminal velocity of a charged spherical oil drop which is made stationary between two electrodes on which a very high potential is applied.

$$\text{Charge on an electron '}q\text{'} = \frac{6\pi\eta r}{E}(v_1 + v_2)$$

where, η = coefficient of viscosity of the gas medium
v_1, v_2 = terminal velocities
E = field strength

$$r = \text{radius of the oil drop} = \sqrt{\frac{9\eta v_1}{2(f - \sigma)g}}$$

(f = density of oil; σ = density of gas; g = gravitational force)]

2.3 POSITIVE RAYS—DISCOVERY OF PROTON

With the discovery of electrons, scientists started looking for positively charged particles which were naturally expected because matter is electrically neutral under ordinary conditions. The first experiment that led to the discovery of the positive particle was conducted by **Goldstein** in 1886. He used a perforated cathode in the modified cathode ray tube (Fig. 2.2). It was observed that when a high potential difference was applied between the electrodes, not only cathode rays were produced but also a new type of rays were produced simultaneously from anode moving towards cathode and passed through the holes or canals of the cathode. These rays were termed **canal rays** since these passed through the canals of the cathode. These were also named **anode rays** as these originated from anode. When the properties of these rays were studied by Thomson, he observed that these rays consisted of positively charged particles and named them as **positive rays.**

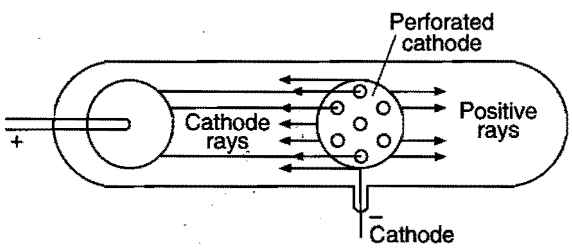


Fig. 2.2

The following characteristics of the positive rays were recognised:

(i) The rays travel in straight lines and cast a shadow of the object placed in their path.

(ii) Like cathode rays, these rays also rotate the wheel placed in their path and also have heating effect. Thus, the rays possess kinetic energy, *i. e.*, mass particles are present.

(iii) The rays produce flashes of light on zinc sulphide screen.

(iv) The rays are deflected by electric and magnetic fields in a direction opposite to that of cathode rays. These rays are attracted towards the negatively charged plate showing thereby that **these rays carry positive charge.**

(v) These rays can pass through thin metal foils.

(vi) These rays can produce ionisation in gases.

(vii) These rays are capable of producing physical and chemical changes.

(viii) Positive particles in these rays have e/m values much smaller than that of electron. This means either m is high or the value of charge is small in comparison to electron. Since, positive particle is formed by the loss of electron or electrons, the charge on the positive particle must be an integral multiple of the charge present on the electron. Hence, for a smaller value of e/m, it is definite that positive particles possess high mass.

(ix) e/m value is dependent on the nature of the gas taken in the discharge tube, *i. e.*, positive particles are different in different gases.

Accurate measurements of the charge and the mass of the particles obtained in the discharge tube containing hydrogen, the lightest of all gases, were made by J.J. Thomson in 1906. These particles were found to have the e/m value as $+9.579 \times 10^4$ coulomb/g. This was the maximum value of e/m observed for any positive particle. It was thus assumed that the positive particle given by hydrogen represents a fundamental particle of positive charge. This particle was named **proton** by **Rutherford** in 1911. Its charge was found to be equal in magnitude but opposite in sign to that of electron.

Thus, **proton carries a charge $+1.602 \times 10^{-19}$ coulomb, *i. e.*, one unit positive charge.**

The mass of the proton, thus, can be calculated.

$$\text{Mass of the proton} = \frac{e}{e/m} = \frac{1.602 \times 10^{-19}}{9.579 \times 10^4}$$

$$= 1.672 \times 10^{-24} \text{ g}$$

or

$$= 1.672 \times 10^{-27} \text{ kg}$$

$$\text{Mass of the proton in amu} = \frac{1.672 \times 10^{-24}}{1.66 \times 10^{-24}} = 1.0072 \text{ amu}$$

A proton is defined as a subatomic particle which has a mass nearly 1 amu and a charge of +1 unit ($+1.602 \times 10^{-19}$ coulomb).

Protons are produced in a number of nuclear reactions. On the basis of such reactions, proton has been recognised as a fundamental building unit of the atom.

2.4 RUTHERFORD EXPERIMENT—DISCOVERY OF NUCLEUS

After the discovery of electron and proton, the question arose how these charged particles are distributed in an atom. The answer was given by J.J. Thomson in the form of first model of the atom.

He proposed that the positive charge is spread over a sphere in which the electrons are embedded to make the atom as a whole neutral. This model was much like raisins in a pudding and is also known as *Thomson's plum pudding model.* This model was discarded as it was not consistent with the results of further investigations such as scattering of α-particles by thin metal foils.

In 1911, Ernest Rutherford and his co-workers carried out a series of experiments using α-particles* (Fig. 2.3 and 2.4). A beam of α-particles was directed against a thin foil of about 0.0004 cm thickness of gold, platinum, silver or copper

*The radiations emitted by radioactive substances consist of α-particles. These particles are positively charged. These particles are actually helium atoms from which electrons have been removed. Each α-particle consists of a mass equal to about 4 times that of hydrogen atom and carries a positive charge of two units. It is represented by the symbol α or ${}_2^4\text{He}$.

$$\underset{\text{Helium atom}}{\text{He}} \longrightarrow \underset{\alpha\text{-particle}}{\text{He}^{2+}} + \underset{\text{Electron}}{2e}$$

α-particles are usually obtained from a natural isotope of polonium-214.

respectively. The foil was surrounded by a circular fluorescent zinc sulphide screen. Whenever an α-particle struck the screen, it produced a flash of light.

The following observations were made:

(i) Most of the α-particles (nearly 99%) went straight without suffering any deflection.

(ii) A few of them got deflected through small angles.

(iii) A very few (about one in 20,000) did not pass through the foil at all but suffered large deflections (more than 90°) or even came back in more or less the direction from which they have come, *i.e.*, a deflection of 180°.

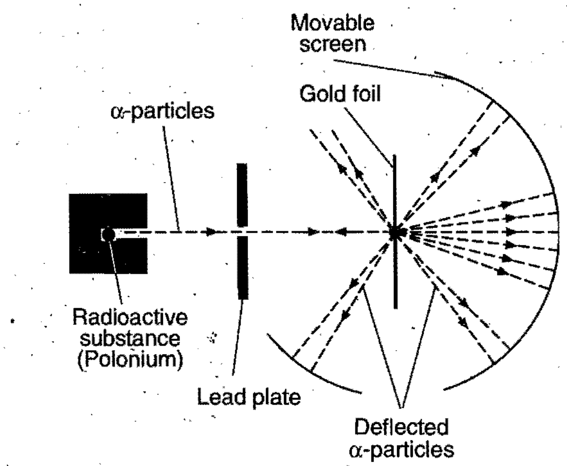


Fig. 2.3

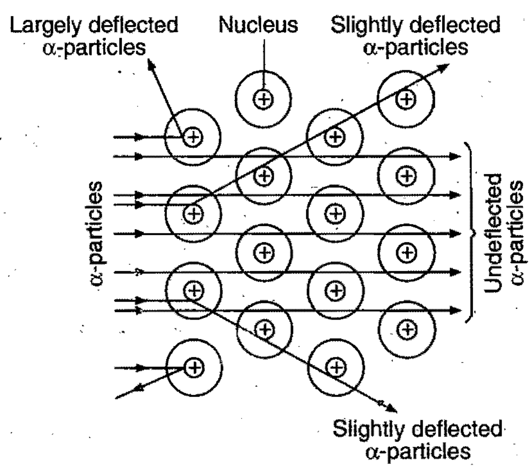


Fig. 2.4 (a)

Consider an α-particle of mass 'm' moving directly towards a nucleus with velocity 'v' at any given time. As this α-particle approaches the nucleus, its velocity and hence kinetic energy continues to decrease. At a certain distance r_0 from the nucleus, the α-particle will stop and then start retracing its depicted path. This distance is called the distance of closest approach. At this distance, the kinetic energy of the α-particle is transformed into electrostatic potential energy. If Z be the atomic number of the nucleus, then

$$\frac{1}{2}mv^2 = \frac{1}{4\pi\varepsilon_0}\frac{Z_e \times 2e}{r_0}$$

$$\because \text{ Electrostatic PE} = \frac{1}{4\pi\varepsilon_0}\frac{q_1 q_2}{r}$$

$$r_0 = \frac{1}{4\pi\varepsilon_0}\frac{4Ze^2}{mv^2}$$

$$r_0 = \frac{1}{4\pi\varepsilon_0}\frac{2Ze^2}{E_K}$$

where, E_K is the original kinetic energy of the α-particles.

Here, $\frac{1}{4\pi\varepsilon_0} = 9 \times 10^9 \text{ Nm}^2\text{C}^{-2}$ (MKS)

$= 1$(CGS)

The distance of closest approach is of the order of 10^{-14} m. So, the radius of the nucleus should be less than 10^{-14} m.

Following conclusions were drawn from the above observations:

(i) Since, most of the α-particles went straight through the metal foil undeflected, it means that **there must be very large empty space within the atom** or **the atom is extraordinarily hollow.**

(ii) A few of the α-particles were deflected from their original paths through moderate angles; it was concluded that whole of the positive charge is concentrated and the space occupied by this positive charge is very small in the atom. When α-particles come closer to this point, they suffer a force of repulsion and deviate from their paths.

The positively charged heavy mass which occupies only a small volume in an atom is called **nucleus.** It is supposed to be present at the centre of the atom.

(iii) A very few of the α-particles suffered strong deflections or even returned on their path indicating that the nucleus is rigid and α-particles recoil due to direct collision with the heavy positively charged mass.

The graph between angle of scattering and the number of α-particles scattering in the corresponding direction is as shown in Fig. 2.4 (b).

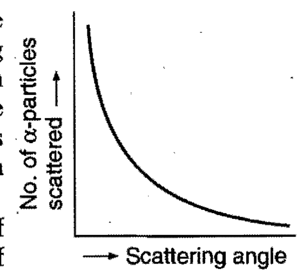


Fig. 2.4 (b)

Information of Rutherford's scattering equation can be memorised by the following relations:

(a) Kinetic energy of α-particles:

$$N = K_1 / [(1/2)mv^2]^2$$

(b) Scattering angle 'θ':

$$N = K_2 / [\sin^4 (\theta/2)]$$

(c) Nuclear charge:

$$N = K_3 (Ze)^2$$

Here, N = Number of α-particles striking the screen and K_1, K_2 and K_3 are the constants.

2.5 MOSELEY EXPERIMENT—ATOMIC NUMBER

Roentgen, in 1895, discovered that when high energy electrons in a discharge tube collide with the anode, penetrating radiations are produced which he named X-rays (Fig. 2.5).

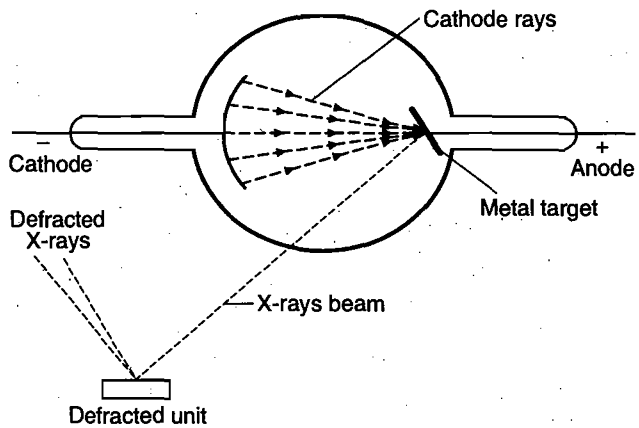


Fig. 2.5

X-rays are electromagnetic radiations of very small wavelengths (0.1–20 Å). X-rays are diffracted by diffraction gratings like ordinary light rays and X-ray spectra are, thus, produced. Each such spectrum is a characteristic property of the element used as anode.

Moseley (1912–1913), investigated the X-ray spectra of 38 different elements, starting from aluminium and ending in gold. He measured the frequency of principal lines of a particular series (the α-lines in the K series) of the spectra. It was observed that the frequency of a particular spectral line gradually increased with the increase of atomic mass of the element. But, it was soon realised that the frequency of the particular spectral line was more precisely related with the serial number of the element in the periodic table which he termed as atomic number (Z). He presented the following relationship:

$$\sqrt{\nu} = a(Z - b)$$

where, ν = frequency of X-rays, Z = atomic number, a and b are constants. When the values of square root of the frequency were plotted against atomic numbers of the elements producing X-rays, a straight line was obtained (Fig. 2.6).

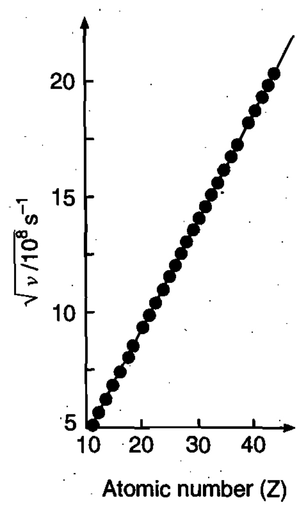


Fig. 2.6

van den Broek (1913) pointed out that the atomic number of an element is equal to the total positive charge contained in the nucleus of its atom. Rutherford was also having the same opinion that the atomic number of an element represents the number of protons in the nucleus of its atom. Thus,

Atomic number of the element
= Serial number of the element in periodic table
= Charge on the nucleus of the atom of the element
= Number of protons present in the nucleus of the atom of the element
= Number of extranuclear electrons present in the atom of the element

2.6 DISCOVERY OF NEUTRON

The discovery of neutron was actually made about 20 years after the structure of atom was elucidated by Rutherford. Atomic masses of different atoms could not be explained if it was accepted that atoms consisted only of protons and electrons. Thus, Rutherford (1920) suggested that in an atom, there must be present at least a third type of fundamental particles which should be electrically neutral and possess mass nearly equal to that of proton. He proposed the name for such fundamental particle as **neutron.** In 1932, **Chadwick** bombarded beryllium with a stream of α-particles. He observed that penetrating radiations were produced which were not affected by electric and magnetic fields. These radiations consisted of neutral particles, which were called neutrons. The nuclear reaction can be shown as:

$$\underset{\text{Beryllium}}{{}^{9}_{4}\text{Be}} + \underset{\alpha\text{-particle}}{{}^{4}_{2}\text{He}} \longrightarrow \underset{\text{Carbon}}{{}^{12}_{6}\text{C}} + \underset{\text{Neutron}}{{}^{1}_{0}n}$$

The mass of the neutron was determined. It was 1.675×10^{-24} g, *i.e.*, nearly equal to the mass of proton.

Thus, **a neutron is a subatomic particle which has a mass 1.675×10^{-24} g, approximately 1 amu, or nearly equal to the mass of proton or hydrogen atom and carrying no electrical charge.** The e/m value of a neutron is thus zero.

2.7 RUTHERFORD MODEL

On the basis of scattering experiments, Rutherford proposed a model of the atom which is known as nuclear atomic model. According to this model:

(i) An atom consists of a heavy positively charged nucleus where all the protons and neutrons are present. Protons and neutrons are collectively referred to as **nucleons.** Almost whole of the mass of the atom is contributed by these nucleons. The magnitude of the positive charge on the nucleus is different for different atoms.

(ii) The volume of the nucleus is very small and is only a minute fraction of the total volume of the atom. Nucleus has a diameter of the order of 10^{-12} to 10^{-13} cm and the atom has a diameter of the order of 10^{-8} cm.

$$\frac{\text{Diameter of the atom}}{\text{Diameter of the nucleus}} = \frac{10^{-8}}{10^{-13}} = 10^5$$

Thus, diameter (size) of an atom is 100,000 times the diameter of the nucleus.*

The radius of a nucleus is proportional to the cube root of the number of nucleons within it.

$$R = R_0 A^{1/3} \text{ cm}$$

where, $R_0 = 1.33 \times 10^{-13}$ cm; A = mass number; R = Radius of the nucleus.

Rutherford and **Marsden** calculated the density of the hydrogen nucleus containing only one proton.

$$d = \frac{\text{Mass}}{\text{Volume}} = \frac{[A \times 1.66 \times 10^{-24} \text{ g}]}{\frac{4}{3} \times \pi R^3 \text{ cm}^3}$$

$$= \frac{A \times 1.66 \times 10^{-24}}{\frac{4}{3} \times 3.14 \times (1.33 \times 10^{-13})^3 \times A}$$

$$= 1.685 \times 10^{14} \text{ g/cm}^3$$

(iii) There is an empty space around the nucleus called extranuclear part. In this part, electrons are present. The number of electrons in an atom is always equal to number of protons present in the nucleus. As the nucleus part of the atom is responsible for the mass of the atom, the extranuclear part is responsible for its volume. The volume of an atom is about 10^{15} times the volume of the nucleus.

$$\frac{\text{Volume of the atom}}{\text{Volume of the nucleus}} = \frac{(10^{-8})^3}{(10^{-13})^3} = \frac{10^{-24}}{10^{-39}} = 10^{15}$$

(iv) Electrons revolve round the nucleus in closed orbits with high speed. The centrifugal force acting on the revolving electrons is being counterbalanced by the force of attraction between the electrons and the nucleus.

This model was similar to the solar system, the nucleus representing the sun and revolving electrons as planets. The electrons are, therefore, generally referred to as **planetary electrons**.

Dissimilarities between Nuclear Atomic Model and Solar System

(i) The sun and the planets are very big bodies and uncharged while the nucleus and electrons are very small objects and charged.

(ii) The revolution of the planets in the solar system is governed by gravitational forces, while the revolution of electrons around the nucleus is governed by electrostatic forces.

(iii) In the solar system, there is only one planet which revolves in any particular orbit, but in the nuclear atomic model more than one electron may rotate in any particular orbit.

Drawbacks of Rutherford Model

(i) According to classical electromagnetic theory, when a charged particle moves under the influence of attractive force, it loses energy continuously in the form of electromagnetic radiations. Thus, when the electron (a charged particle) moves in an attractive field (created by protons present in the nucleus), it must emit radiations. As a result of this, the electron should lose energy at every turn and move closer and closer to the nucleus following a spiral path (Fig. 2.7). The ultimate result will be that it will fall into the nucleus, thereby making the atom unstable. Bohr made calculations and pointed out that an atom would collapse in 10^{-8} seconds. Since, the atom is quite stable, it means the electrons do not fall into the nucleus, thereby this model does not explain the stability of the atom.

(ii) If the electrons lose energy continuously, the observed spectrum should be continuous but the actual observed spectrum consists of well defined lines of definite frequencies. Hence, the loss of energy by the electrons is not continuous in an atom.

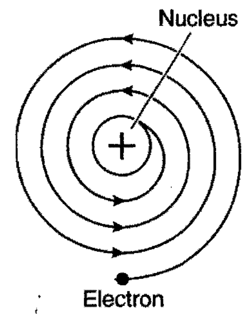


Fig. 2.7

2.8 ELECTROMAGNETIC RADIATIONS

These are energy radiations which do not need any medium for propagation, *e.g.*, visible, ultraviolet, infrared, X-rays, γ-rays, etc. An electromagnetic radiation is generated by oscillations of a charged body in a magnetic field or a magnet in an electrical field. The frequency of a wave is the frequency of oscillation of the oscillating charged particle. These radiations or waves have electrical and magnetic fields associated with them and travel at right angle to these fields. The following are thus the important characteristics of electromagnetic radiations:

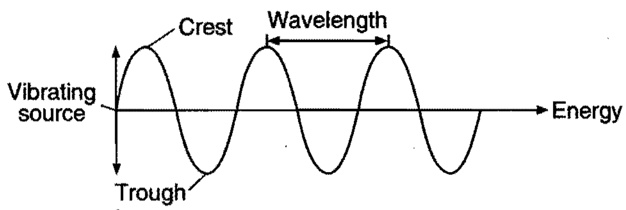
- All electromagnetic radiations travel with the velocity of light.
- These consist of electric and magnetic fields that oscillate in directions perpendicular to each other and perpendicular to the direction in which the wave is travelling.

S. No.	Name	Wavelength (Å)	Frequency (Hz)	Source
1.	Radio wave	$3 \times 10^{14} - 3 \times 10^7$	$1 \times 10^5 - 1 \times 10^9$	Alternating current of high frequency
2.	Micro wave	$3 \times 10^7 - 6 \times 10^6$	$1 \times 10^9 - 5 \times 10^{11}$	Klystron tube
3.	Infrared (IR)	$6 \times 10^6 - 7600$	$5 \times 10^{11} - 3.95 \times 10^{16}$	Incandescent objects
4.	Visible	7600–3800	$3.95 \times 10^{16} - 7.9 \times 10^{14}$	Electric bulbs, sun rays
5.	Ultraviolet (UV)	3800–150	$7.9 \times 10^{14} - 2 \times 10^{16}$	Sun rays, arc lamps with mercury vapours
6.	X-Rays	150–0.1	$2 \times 10^{16} - 3 \times 10^{19}$	Cathode rays striking metal plate
7.	γ-Rays	0.1–0.01	$3 \times 10^{19} - 3 \times 10^{20}$	Secondary effect of radioactive decay
8.	Cosmic rays	0.01–Zero	3×10^{20}–Infinity	Outer space

*The diameter of various atoms lies in the range of 0.74 to 4.70 Å (1 Å = 10^{-8} cm).

A wave is always characterized by the following six characteristics:

(i) Wavelength: The distance between two nearest crests or nearest troughs is called the wavelength. It is denoted by λ (lambda) and is measured in terms of centimetre (cm), angstrom (Å), micrometre (μm) or nanometre (nm).



$$1\ \text{Å} = 10^{-8}\ \text{cm} = 10^{-10}\ \text{m}$$
$$1\ \mu\text{m} = 10^{-4}\ \text{cm} = 10^{-6}\ \text{m}$$
$$1\ \text{nm} = 10^{-7}\ \text{cm} = 10^{-9}\ \text{m}$$
$$1\ \text{cm} = 10^{8}\ \text{Å} = 10^{4}\ \mu\text{m} = 10^{7}\ \text{nm}$$

(ii) Frequency: It is defined as the number of waves which pass through a point in one second. It is denoted by the symbol ν (nu) and is measured in terms of cycles (or waves) per second (cps) or hertz (Hz).

$$\lambda\nu = \text{distance travelled in one second}$$
$$= \text{velocity} = c$$

or $$\nu = \frac{c}{\lambda}$$

(iii) Velocity: It is defined as the distance covered in one second by the wave. It is denoted by the letter 'c'. All electromagnetic waves travel with the same velocity, *i.e.*, 3×10^{10} cm/sec.

$$\lambda\nu = 3 \times 10^{10}$$

Thus, a wave of higher frequency has a shorter wavelength while a wave of lower frequency has a longer wavelength.

(iv) Wave number: This is the reciprocal of wavelength, *i.e.*, the number of wavelengths per centimetre. It is denoted by the symbol $\bar{\nu}$ (nu bar).

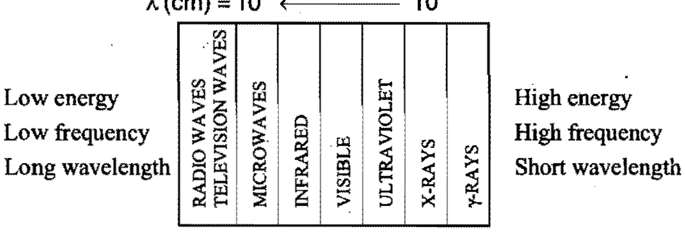
$$\bar{\nu} = \frac{1}{\lambda}$$

It is expressed in cm^{-1} or m^{-1}.

(v) Amplitude: It is defined as the height of the crest or depth of the trough of a wave. It is denoted by the letter 'a'. It determines the intensity of the radiation.

The arrangement of various types of electromagnetic radiations in the order of their increasing or decreasing wavelengths or frequencies is known as electromagnetic spectrum.

$$\nu = 3 \times 10^{7}\ (\text{cycle/sec}) \xrightarrow{\text{Frequency}} 3 \times 10^{21}$$
$$\lambda\ (\text{cm}) = 10^{3} \xleftarrow{\text{Wavelength}} 10^{-11}$$



(vi) Time period: Time taken by the wave for one complete cycle or vibration is called time period. It is denoted by T.

$$T = \frac{1}{\nu}$$

Unit: Second per cycle.

2.9 EMISSION SPECTRA—HYDROGEN SPECTRUM

Spectrum is the impression produced on a screen when radiations of particular wavelengths are analysed through a prism or diffraction grating. It is broadly of two types:

(i) Emission spectra (ii) Absorption spectra.

Difference between Emission and Absorption Spectrum

Emission spectrum	Absorption spectrum
1. It gives bright lines (coloured) on the dark background.	It gives dark lines on the bright background.
2. Radiations from emitting source are analysed by the spectroscope.	It is observed when the white light is passed through the substance and the transmitted radiations are analysed by the spectroscope.
3. It may be continuous (if source emits white light) and may be discontinuous (if the source emits coloured light).	These are always discontinuous.

Emission spectra: It is obtained from the substances which emit light on excitation, *i.e.*, either by heating the substances on a flame or by passing electric discharge through gases at low pressure or by passing electric current discharge through a thin filament of a high melting point metal. Emission spectra are of two types:

(a) Continuous spectra and (b) Discontinuous spectra.

(a) Continuous spectra: When white light is allowed to pass through a prism, it gets resolved into several colours (Fig. 2.8). The spectrum is a rainbow of colours, *i.e.*, violet merges into blue, blue into green, and so on. This is a continuous spectrum.

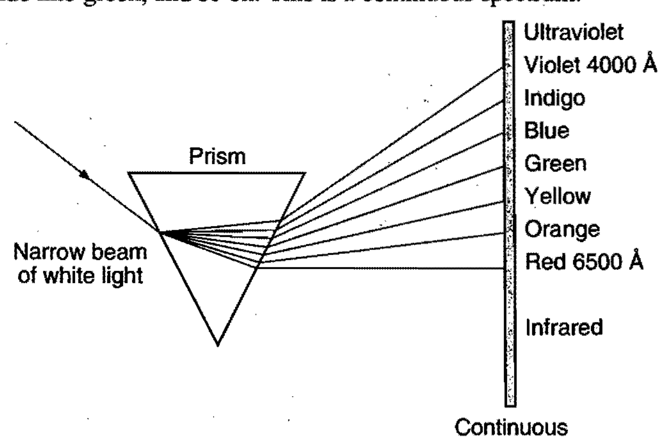


Fig. 2.8 Continuous spectrum of white light

(b) Discontinuous spectra: When gases or vapours of a chemical substance are heated in an electric arc or in a Bunsen flame, light is emitted. If a ray of this light is passed through a prism, a line spectrum is produced. This spectrum consists of a limited number of lines, each of which corresponds to a different wavelength of light. The line spectrum of each element is unique. Hydrogen spectrum is an example of line emission spectrum or atomic emission spectrum.

When an electric discharge is passed through hydrogen gas at low pressure, a bluish light is emitted. When a ray of this light is passed through a prism, a discontinuous line spectrum of several isolated sharp lines is obtained. The wavelengths of various lines show that these lines lie in visible, ultraviolet and infrared regions. All these lines observed in the hydrogen spectrum can be classified into six series.

Spectral series	Discovered by	Appearing in
Lyman series	Lyman	Ultraviolet region
Balmer series	Balmer	Visible region
Paschen series	Paschen	Infrared region
Brackett series	Brackett	Infrared region
Pfund series	Pfund	Infrared region
Humphrey series	Humphrey	Far infrared region

Ritz presented a mathematical formula to find the wavelengths of various hydrogen lines.

$$\bar{\nu} = \frac{1}{\lambda} = \frac{\nu}{c} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Where, R is a universal constant, known as Rydberg constant. Its value is 109,678 cm^{-1}, n_1 and n_2 are integers (such that $n_2 > n_1$). For a given spectral series, n_1 remains constant while n_2 varies from line to line in the same series.

The value of $n_1 = 1, 2, 3, 4$ and 5 for the Lyman, Balmer, Paschen, Brackett and Pfund series respectively. n_2 is greater than n_1 by at least 1.

Values of n_1 and n_2 for various series

Spectral series	Value of n_1	Value of n_2
Lyman series	1	2, 3, 4, 5, . . .
Balmer series	2	3, 4, 5, 6, . . .
Paschen series	3	4, 5, 6, 7, . . .
Brackett series	4	5, 6, 7, 8, . . .
Pfund series	5	6, 7, 8, 9, . . .

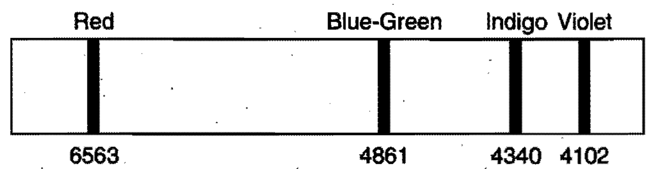


Fig. 2.9 (a) Balmer series in the hydrogen spectrum

Lyman series: $\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{1^2} - \frac{1}{n_2^2}\right]$

$n_2 = 2, 3, 4, 5, \ldots$

Obtained in emission as well as in absorption spectrum ratio of m^{th} to n^{th} wavelength of Lyman series.

$$\frac{\lambda_m}{\lambda_n} = \left(\frac{m+1}{n+1}\right)^2 \cdot \left\{\frac{(n+1)^2 - 1}{(m+1)^2 - 1}\right\}$$

Balmer series: $\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{2^2} - \frac{1}{n_2^2}\right]$

$n_2 = 3, 4, 5, 6, \ldots$

Obtained only in emission spectrum.

Paschen series: $\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{3^2} - \frac{1}{n_2^2}\right]$

$n_2 = 4, 5, 6, 7, \ldots$

Brackett series: $\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{4^2} - \frac{1}{n_2^2}\right]$

$n_2 = 5, 6, 7, 8, \ldots$

Pfund series: $\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{5^2} - \frac{1}{n_2^2}\right]$

$n_2 = 6, 7, 8, 9, \ldots$

Note : (i) Atoms give line spectra while molecules give band spectra.

(ii) Balmer, Paschen, Brackett, Pfund series are found in emission spectrum.

Electronic transition	Name of line	Wave no.	Wavelength and colour
$n_2 = \underset{(M)}{3} \longrightarrow n_1 = \underset{(L)}{2}$	H_α (First line)	$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{3^2}\right] = \frac{5R}{36}$	$\lambda = 6563$ Å (Red)
$n_2 = \underset{(N)}{4} \longrightarrow n_1 = \underset{(L)}{2}$	H_β (Second line)	$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{4^2}\right] = \frac{3R}{16}$	$\lambda = 4861$ Å (Blue-Green)
$n_2 = \underset{(O)}{5} \longrightarrow n_1 = \underset{(L)}{2}$	H_γ (Third line)	$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{5^2}\right] = \frac{21R}{100}$	$\lambda = 4340$ Å (Indigo)
$n_2 = \underset{(P)}{6} \longrightarrow n_1 = \underset{(L)}{2}$	H_δ (Fourth line)	$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{6^2}\right] = \frac{8R}{36}$	$\lambda = 4102$ Å (Violet)

(Above four lines were viewed in Balmer series by naked eye.)

Absorption Spectrum : Suppose the radiations from a continuous source like a hot body (sun light) containing the quanta of all wavelengths passes through a sample of hydrogen gas, then the wavelengths missing in the emergent light give dark lines on the bright background. This type of spectrum that contains lesser number of wavelengths in the emergent light than in incident light is called absorption spectrum.

Let the radiations of wavelengths $\lambda_1, \lambda_2, \lambda_3, \lambda_4, \lambda_5$ are passed through the sample of hydrogen gas such that λ_1 and λ_4 are absorbed then the absorption spectrum may be represented as:

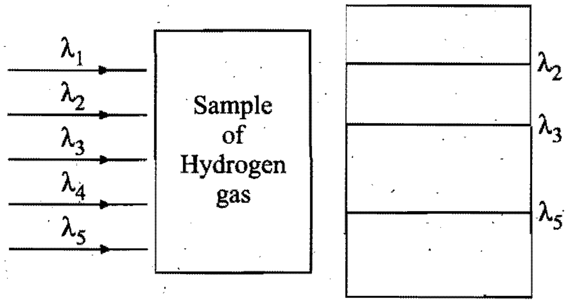


Fig. 2.9 (b) Absorption Spectrum

2.10 QUANTUM THEORY OF RADIATION

The wave theory successfully explains many properties of electromagnetic radiations such as reflection, refraction, diffraction, interference, polarisation, etc., but fails to explain some phenomena like black body radiation, photoelectric effect, etc.

In order to explain black body radiation and photoelectric effect, Max Planck in 1901 presented a new theory which is known as **quantum theory of radiation.** According to this theory, a hot body emits radiant energy not continuously but discontinuously in the form of small packets of energy called **quantum** (in plural quanta). The energy associated with each quantum of a given radiation is proportional to the frequency of the emitted radiation.

$$E \propto \nu$$

Or $E = h\nu$ where, h is a constant known as Planck's constant. Its numerical value is 6.624×10^{-27} erg-sec. The energy emitted or absorbed by a body can be either one quantum or any whole number multiple of $h\nu$, *i.e.*, $2h\nu, 3h\nu, 4h\nu, \ldots, nh\nu$ quanta of energy.

Thus, energy emitted or absorbed = $nh\nu$, where n can have values 1, 2, 3, 4, ... Thus, the energy emitted or absorbed is quantised.

In 1905, Einstein pointed out that light can be supposed to consist of a stream of particles, called **photons.** The energy of each photon of light depends on the frequency of the light, *i.e.*, $E = h\nu$. Energy is also related according to Einstein, as $E = mc^2$ where m is the mass of photon. Thus, it was pointed out that light has wave as well as particle characteristics (dual nature).

SOME SOLVED EXAMPLES

Example 1. *How many protons, electrons and neutrons are present in* 0.18 g $^{30}_{15}P$?

Solution: No. of protons in one atom of P

= No. of electrons in one atom of P = 15

No. of neutrons in one atom of P = $(A - Z) = (30 - 15) = 15$

$$0.18 \text{ g } ^{30}_{15}\text{P} = \frac{0.18}{30} = 0.006 \text{ mol}$$

No. of atoms in 0.006 mol = $0.006 \times 6.02 \times 10^{23}$

No. of protons in 0.006 mol $^{30}_{15}$P = $15 \times 0.006 \times 6.02 \times 10^{23}$

$= 5.418 \times 10^{22}$

So, No. of electrons = 5.418×10^{22}

and No. of neutrons = 5.418×10^{22}

Example 2. *Calculate the frequency and wave number of radiation with wavelength* 480 *nm.*

Solution: Given,

$$\lambda = 480 \text{ nm} = 480 \times 10^{-9} \text{ m} \qquad [\because 1 \text{ nm} = 10^{-9} \text{ m}]$$

$$c = 3 \times 10^8 \text{ m/sec}$$

$$\text{Frequency, } \nu = \frac{c}{\lambda} = \frac{3 \times 10^8 \text{ ms}^{-1}}{480 \times 10^{-9} \text{ m}} = 6.25 \times 10^{14} \text{ s}^{-1}$$

$$= 6.25 \times 10^{14} \text{ Hz}$$

Example 3. *Calculate the energy associated with photon of light having a wavelength* 6000 *Å.* [$h = 6.624 \times 10^{-27}$ *erg-sec.*]

Solution: We know that, $E = h\nu = h \cdot \frac{c}{\lambda}$

$h = 6.624 \times 10^{-27}$ erg-sec; $c = 3 \times 10^{10}$ cm/sec;

$\lambda = 6000 \text{ Å} = 6000 \times 10^{-8}$ cm

$$\text{So, } E = \frac{(6.624 \times 10^{-27}) \times (3 \times 10^{10})}{6 \times 10^{-5}} = 3.312 \times 10^{-12} \text{ erg.}$$

Example 4. *Which has a higher energy, a photon of violet light with wavelength* 4000 *Å or a photon of red light with wavelength* 7000 *Å?* [$h = 6.62 \times 10^{-34}$ *Js*]

Solution: We know that, $E = h\nu = h \cdot \frac{c}{\lambda}$

Given, $h = 6.62 \times 10^{-34}$ Js, $c = 3 \times 10^8 \text{ ms}^{-1}$

For a photon of violet light,

$$\lambda = 4000 \text{ Å} = 4000 \times 10^{-10} \text{ m}$$

$$E = 6.62 \times 10^{-34} \times \frac{3 \times 10^8}{4 \times 10^{-7}} = 4.96 \times 10^{-19} \text{ J}$$

For a photon of red light,

$$\lambda = 7000 \text{ Å} = 7000 \times 10^{-10} \text{ m}$$

$$E = 6.62 \times 10^{-34} \times \frac{3 \times 10^8}{7000 \times 10^{-10}} = 2.83 \times 10^{-19} \text{ J}$$

Hence, photon of violet light has higher energy than the photon of red light.

Example 5. *What is the ratio between the energies of two radiations one with a wavelength of* 6000 *Å and other with* 2000 *Å?*

Solution: $\lambda_1 = 6000 \text{ Å}$ and $\lambda_2 = 2000 \text{ Å}$

$$E_1 = h \cdot \frac{c}{\lambda_1} \text{ and } E_2 = h \cdot \frac{c}{\lambda_2}$$

$$\text{Ratio, } \frac{E_1}{E_2} = \frac{h \cdot c}{\lambda_1} \times \frac{\lambda_2}{h \cdot c} = \frac{\lambda_2}{\lambda_1} = \frac{2000}{6000} = \frac{1}{3}$$

or $E_2 = 3E_1$

Example 6. *Calculate the wavelength, wave number and frequency of photon having an energy equal to three electron volt.* ($h = 6.62 \times 10^{-27}$ *erg-sec.*)

Solution: We know that,

$$E = h \cdot \nu$$

$$\nu = \frac{E}{h} \quad (1\,\text{eV} = 1.602 \times 10^{-12}\text{ erg})$$

$$= \frac{3 \times (1.602 \times 10^{-12})}{6.62 \times 10^{-27}} = 7.26 \times 10^{14}\text{ s}^{-1} = 7.26 \times 10^{14}\text{ Hz}$$

$$\lambda = \frac{c}{\nu} = \frac{3 \times 10^{10}}{7.26 \times 10^{14}} = 4.132 \times 10^{-5}\text{ cm}$$

$$\overline{\nu} = \frac{1}{\lambda} = \frac{1}{4.132 \times 10^{-5}} = 2.42 \times 10^{4}\text{ cm}^{-1}$$

Example 7. *Calculate the energy in kilocalorie per mol of the photons of an electromagnetic radiation of wavelength 7600 Å.*

Solution: $\lambda = 7600\,\text{Å} = 7600 \times 10^{-8}$ cm

$$c = 3 \times 10^{10}\text{ cm s}^{-1}$$

$$\text{Frequency, } \nu = \frac{c}{\lambda} = \frac{3 \times 10^{10}}{7600 \times 10^{-8}} = 3.947 \times 10^{14}\text{ s}^{-1}$$

$$\text{Energy of one photon} = h\nu = 6.62 \times 10^{-27} \times 3.947 \times 10^{14}$$
$$= 2.61 \times 10^{-12}\text{ erg}$$

$$\text{Energy of one mole of photons} = 2.61 \times 10^{-12} \times 6.02 \times 10^{23}$$
$$= 15.71 \times 10^{11}\text{ erg}$$

Energy of one mole of photons in kilocalorie

$$= \frac{15.71 \times 10^{11}}{4.185 \times 10^{10}} [1\text{ kcal} = 4.185 \times 10^{10}\text{ erg}]$$
$$= 37.538\text{ kcal per mol}$$

Example 8. *Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy in kJ mol^{-1}, $h = 6.6256 \times 10^{-34}$ Js.* **(IIT 1992)**

Solution: $\lambda = 242\text{ nm} = 242 \times 10^{-9}$ m

$$c = 3 \times 10^{8}\text{ ms}^{-1}$$

$$E = h\nu = h \cdot \frac{c}{\lambda} = 6.6256 \times 10^{-34} \times \frac{3 \times 10^{8}}{242 \times 10^{-9}}$$
$$= 0.082 \times 10^{-17}\text{ J} = 0.082 \times 10^{-20}\text{ kJ}$$

$$\text{Energy per mole for ionisation} = 0.082 \times 10^{-20} \times 6.02 \times 10^{23}$$
$$= 493.6\text{ kJ mol}^{-1}$$

Example 9. *How many photons of light having a wavelength 4000 Å are necessary to provide 1.00 J of energy?*

Solution: Energy of one photon

$$= h\nu = h \cdot \frac{c}{\lambda}$$
$$= \frac{(6.62 \times 10^{-34})(3.0 \times 10^{8})}{4000 \times 10^{-10}}$$
$$= 4.965 \times 10^{-19}\text{ J}$$

$$\text{Number of photons} = \frac{1.00}{4.965 \times 10^{-19}} = 2.01 \times 10^{18}$$

Example 10. *Find the number of quanta of radiations of frequency 4.67×10^{13} s^{-1}, that must be absorbed in order to melt 5 g of ice. The energy required to melt 1g of ice is 333 J.*

Solution: Energy required to melt 5 g of ice

$$= 5 \times 333 = 1665\text{ J}$$

Energy associated with one quantum

$$= h\nu = (6.62 \times 10^{-34}) \times (4.67 \times 10^{13})$$
$$= 30.91 \times 10^{-21}\text{ J}$$

Number of quanta required to melt 5 g of ice

$$= \frac{1665}{30.91 \times 10^{-21}} = 53.8 \times 10^{21}$$
$$= 5.38 \times 10^{22}$$

Example 11. *Calculate the wavelength of the spectral line, when the electron in the hydrogen atom undergoes a transition from the energy level 4 to energy level 2.*

Solution: According to Rydberg equation,

$$\frac{1}{\lambda} = R\left(\frac{1}{x^2} - \frac{1}{y^2}\right)$$

$$R = 109678\text{ cm}^{-1}; \quad x = 2; \quad y = 4$$

$$\frac{1}{\lambda} = 109678\left[\frac{1}{4} - \frac{1}{16}\right]$$
$$= 109678 \times \frac{3}{16}$$

On solving, $\lambda = 486$ nm

Example 12. *A bulb emits light of wavelength $\lambda = 4500$ Å. The bulb is rated as 150 watt and 8% of the energy is emitted as light. How many photons are emitted by the bulb per second?* **(IIT 1995)**

Solution: Energy emitted per second by the bulb

$$= 150 \times \frac{8}{100}\text{ J}$$

$$\text{Energy of 1 photon} = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{4500 \times 10^{-10}}$$
$$= 4.42 \times 10^{-19}\text{ joule}$$

Let n photons be evolved per second.

$$\therefore \quad n \times 4.42 \times 10^{-19} = 150 \times \frac{8}{100}$$
$$n = 27.2 \times 10^{18}$$

Example 13. *A near ultraviolet photon of 300 nm is absorbed by a gas and then remitted as two photons. One photon is red with wavelength of 760 nm. What would be the wave number of the second photon?*

Solution:

Energy absorbed = Sum of energy of two quanta

$$\frac{hc}{300 \times 10^{-9}} = \frac{hc}{760 \times 10^{-9}} + \frac{hc}{\lambda \times 10^{-9}}$$

On solving, we get,

$$\bar{\nu}\,(\text{wave number}) = \frac{1}{\lambda} = 2.02 \times 10^{-3}\ \text{m}^{-1}$$

Example 14. *Calculate the wavelength of the radiation which would cause the photodissociation of chlorine molecule if the Cl—Cl bond energy is 243 kJ mol $^{-1}$.*

Solution: Energy required to break one Cl—Cl bond

$$= \frac{\text{Bond energy per mole}}{\text{Avogadro's number}}$$

$$= \frac{243}{6.023 \times 10^{23}}\ \text{kJ} = \frac{243 \times 10^3}{6.023 \times 10^{23}}\ \text{J}$$

Let the wavelength of the photon to cause rupture of one Cl—Cl bond be λ.

We know that,

$$\lambda = \frac{hc}{E} = \frac{6.6 \times 10^{-34} \times 3 \times 10^8 \times 6.023 \times 10^{23}}{243 \times 10^3}$$

$$= 4.90 \times 10^{-7}\ \text{m} = 490\ \text{nm}$$

Example 15. *How many moles of photon would contain sufficient energy to raise the temperature of 225 g of water 21 °C to 96 °C? Specific heat of water is 4.18 J g $^{-1}$ K $^{-1}$ and frequency of light radiation used is $2.45 \times 10^9\ s^{-1}$.*

Solution: Energy associated with one mole of photons

$$= N_0 \times h \times \nu$$

$$= 6.02 \times 10^{23} \times 6.626 \times 10^{-34} \times 2.45 \times 10^9$$

$$= 97.727 \times 10^{-2}\ \text{J mol}^{-1}$$

Energy required to raise the temperature of 225 g of water by 75°C $= m \times s \times t$

$$= 225 \times 4.18 \times 75 = 70537.5\ \text{J}$$

Hence, number of moles of photons required

$$= \frac{mst}{N_0 h\nu} = \frac{70537.5}{97.727 \times 10^{-2}} = 7.22 \times 10^4\ \text{mol}$$

Example 16. *During photosynthesis, chlorophyll-a absorbs light of wavelength 440 nm and emits light of wavelength 670 nm. What is the energy available for photosynthesis from the absorption-emission of a mole of photons?*

Solution: $\Delta E = \left[\frac{Nhc}{\lambda}\right]_{\text{absorbed}} - \left[\frac{Nhc}{\lambda}\right]_{\text{evolved}}$

$$= Nhc\left[\frac{1}{\lambda_{\text{absorbed}}} - \frac{1}{\lambda_{\text{evolved}}}\right]$$

$$= 6.023 \times 10^{23} \times 6.626 \times 10^{-34} \times 3 \times 10^8 \left[\frac{1}{440 \times 10^{-9}} - \frac{1}{670 \times 10^{-9}}\right]$$

$$= 0.1197[2.272 \times 10^6 - 1.492 \times 10^6]$$

$$= 0.0933 \times 10^6\ \text{J/mol} = 93.3\ \text{kJ/mol}$$

Example 17. *Photochromic sunglasses, which darken when exposed to light, contain a small amount of colourless AgCl (s) embedded in the glass. When irradiated with light, metallic silver atoms are produced and the glass darkens.*

$$AgCl(s) \longrightarrow Ag(s) + Cl$$

Escape of chlorine atoms is prevented by the rigid structure of the glass and the reaction therefore, reverses as soon as the light is removed. If 310 kJ / mol of energy is required to make the reaction proceed, what wavelength of light is necessary?

Solution: Energy per mole = Energy of one Einstein

i.e., energy of one mole quanta

$$= \frac{Nhc}{\lambda}$$

$$\therefore \quad 310 \times 1000 = \frac{6.023 \times 10^{23} \times 6.626 \times 10^{-34} \times 3 \times 10^8}{\lambda}$$

$$\lambda = 3.862 \times 10^{-7}\ \text{m} = 3862 \times 10^{-10}\ \text{m} = 3862\ \text{Å}$$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

1. The frequency of the radiation having wave number 10 m $^{-1}$ is:
   (a) $10\ s^{-1}$ (b) $3 \times 10^7\ s^{-1}$
   (c) $3 \times 10^{11}\ s^{-1}$ (d) $3 \times 10^9\ s^{-1}$
   [**Ans.** (d)]
   [**Hint:** $\bar{\nu} = \frac{1}{\lambda}$
   $\nu = c\bar{\nu} = \frac{c}{\lambda} = 3 \times 10^8 \times 10 = 3 \times 10^9\ s^{-1}$]

2. The energy of a photon of radiation having wavelength 300 nm is:
   (a) 6.63×10^{-29} J (b) 6.63×10^{-19} J
   (c) 6.63×10^{-28} J (d) 6.63×10^{-17} J
   [**Ans.** (b)]
   [**Hint:** $E = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{300 \times 10^{-9}} = 6.63 \times 10^{-19}$ J]

3. The maximum kinetic energy of the photoelectrons is found to be 6.63×10^{-19} J, when the metal is irradiated with a radiation of frequency 2×10^{15} Hz. The threshold frequency of the metal is about:
   (a) $1 \times 10^{15}\ s^{-1}$ (b) $2 \times 10^{15}\ s^{-1}$
   (c) $3 \times 10^{15}\ s^{-1}$ (d) $1.5 \times 10^{15}\ s^{-1}$
   [**Ans.** (a)]
   [**Hint:** $\text{KE} = h(\nu - \nu_0)$
   $\nu_0 = \nu - \frac{\text{KE}}{h}$
   $= 2 \times 10^{15} - \frac{6.63 \times 10^{-19}}{6.63 \times 10^{-34}} = 1 \times 10^{15}\ s^{-1}$]

4. The number of photons of light having wavelength 100 nm which can provide 1 J energy is nearly:
   (a) 10^7 photons (b) 5×10^{18} photons
   (c) 5×10^{17} photons (d) 5×10^7 photons
   [**Ans.** (c)]
   [**Hint:** $E = \frac{nhc}{\lambda}$
   $n = \frac{E\lambda}{hc} = \frac{1 \times 100 \times 10^{-9}}{6.626 \times 10^{-34} \times 3 \times 10^8} = 5 \times 10^{17}$]

5. The atomic transition gives rise to the radiation of frequency (10^4 MHz). The change in energy per mole of atoms taking place would be:
(a) 3.99×10^{-6} J (b) 3.99 J
(c) 6.62×10^{-24} J (d) 6.62×10^{-30} J
[**Ans.** (b)]
[**Hint:** $E = Nh\nu$
$= 6.023 \times 10^{23} \times 6.626 \times 10^{-34} \times 10^4 \times 10^6$
$= 3.99$ J]

2.11 BOHR'S ATOMIC MODEL

To overcome the objections of Rutherford's model and to explain the hydrogen spectrum, **Bohr** proposed a **quantum mechanical model** of the atom. This model was based on the quantum theory of radiation and the classical laws of physics. The important postulates on which Bohr's model is based are the following:

(i) The atom has a nucleus where all the protons and neutrons are present. The size of the nucleus is very small. It is present at the centre of the atom.

(ii) Negatively charged electrons are revolving around the nucleus in the same way as the planets are revolving around the sun. The path of the electron is circular. The force of attraction between the nucleus and the electron is equal to centrifugal force of the moving electron.

Force of attraction towards nucleus = centrifugal force

(iii) Out of infinite number of possible circular orbits around the nucleus, the electron can revolve only on those orbits whose angular momentum* is an integral multiple of $\frac{h}{2\pi}$, *i.e.*, $mvr = n\frac{h}{2\pi}$ where m = mass of the electron, v = velocity of electron, r = radius of the orbit and $n = 1, 2, 3, \ldots$ number of the orbit. The angular momentum can have values such as, $\frac{h}{2\pi}, \frac{2h}{2\pi}, \frac{3h}{2\pi}$, etc., but it cannot have a fractional value. Thus, the angular momentum is quantized. The specified or circular orbits (quantized) are called **stationary orbits.**

(iv) By the time, the electron remains in any one of the stationary orbits, it does not lose energy. Such a state is called **ground or normal state.**

In the ground state, potential energy of electron will be minimum, hence it will be the most stable state.

(v) Each stationary orbit is associated with a definite amount of energy. The greater is the distance of the orbit from the nucleus, more shall be the energy associated with it. These orbits are also called energy levels and are numbered as 1, 2, 3, 4, . . . or K, L, M, N, . . . from nucleus outwards.

i.e.,
$$E_1 < E_2 < E_3 < E_4 \ldots$$
$$(E_2 - E_1) > (E_3 - E_2) > (E_4 - E_3) \ldots$$

(vi) The emission or absorption of energy in the form of radiation can only occur when an electron jumps from one stationary orbit to another.

$$\Delta E = E_{\text{high}} - E_{\text{low}} = h\nu$$

Energy is absorbed when the electron jumps from inner to outer orbit and is emitted when it moves from outer to an inner orbit.

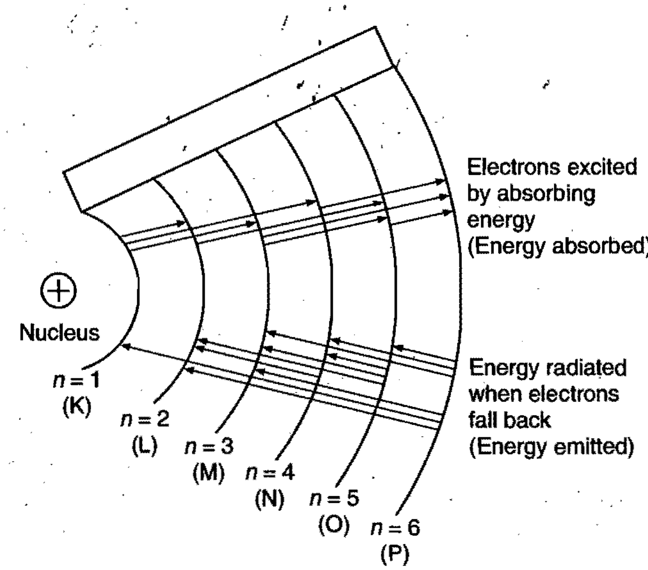


Fig. 2.10

When the electron moves from inner to outer orbit by absorbing definite amount of energy, the new state of the electron is said to be **excited state** (Fig. 2.10).

Using the above postulates, Bohr calculated the radii of various stationary orbits, the energy associated with each orbit and explained the spectrum of hydrogen atom.

Radii of various orbits: Consider an electron of mass 'm' and charge 'e' revolving around the nucleus of charge Ze (Z = atomic number). Let 'v' be the tangential velocity of the revolving electron and 'r' the radius of the orbit (Fig. 2.11). The electrostatic force of attraction between the nucleus and electron (applying Coulomb's law) $= \frac{kZe \times e}{r^2} = \frac{kZe^2}{r^2}$

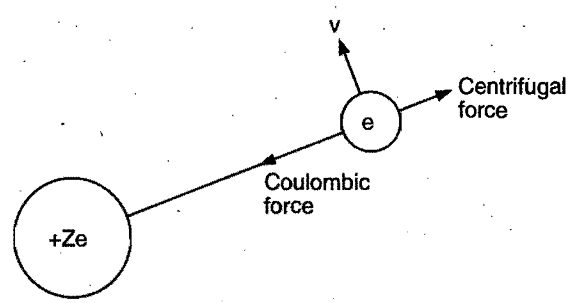


Fig. 2.11

*Angular momentum = $I\omega$, where, I = moment of inertia and ω = angular velocity;
$\omega = \frac{v}{r}$, where, v = linear velocity and r = radius; and $I = mr^2$. So angular momentum $= mr^2 \cdot \frac{v}{r} = mvr$.

where, k is a constant. It is equal to $\frac{1}{4\pi\varepsilon_0}$, ε_0 being absolute permittivity of medium. In SI units, the numerical value of $\frac{1}{4\pi\varepsilon_0}$ is equal to $9\times10^9\ Nm^2/C^2$.

[**Note:** In CGS units, value of k is equal to 1.]

As force of attraction = centrifugal force

So, $$\frac{kZe^2}{r^2}=\frac{mv^2}{r} \quad \text{or} \quad v^2=\frac{kZe^2}{rm}$$

$$v^2=\frac{1}{4\pi\varepsilon_0}\cdot\frac{Ze^2}{rm} \qquad \ldots \text{(i)}$$

According to one of the postulates,

Angular momentum $= mvr = n\frac{h}{2\pi}$

or $$v=\frac{nh}{2\pi mr} \qquad \ldots \text{(ii)}$$

Putting the value of 'v' in eq. (i),

$$\frac{n^2h^2}{4\pi^2m^2r^2}=\frac{kZe^2}{mr} \quad \text{or} \quad \frac{n^2h^2}{4\pi^2mr}=kZe^2$$

or $$r=\frac{n^2h^2}{4\pi^2mkZe^2} \qquad \ldots \text{(iii)}$$

Greater is the value of 'n', larger is the size of atom. On the other hand, greater is the value of 'Z', smaller is the size of the atom. Across a period from left to right, atomic number 'Z' increases with constant value of 'n' hence atomic radius decreases towards right. On moving down the group, both 'Z' and 'n' increase but due to shielding, Z^* (effective nuclear charge) remains same. Hence, on moving downwards, atomic radius increases due to increase in 'n'.

For hydrogen atom, $Z=1$; so $r=\frac{n^2h^2}{4\pi^2mke^2}$

Now putting the values of h, π, m, e and k,

$$r=\frac{n^2\times(6.625\times10^{-34})^2}{4\times(3.14)^2\times(9.1\times10^{-31})\times(1.6\times10^{-19})^2\times(9\times10^9)}$$

$$=0.529\times n^2\times10^{-10}\ \text{m}=0.529\times n^2\ \text{Å}$$

$$=0.529\times10^{-8}\times n^2\ \text{cm}$$

where $h=6.625\times10^{-34}$ J-sec, $\pi=3.14$

$m=9.1\times10^{-31}$ kg, $e=1.6\times10^{-19}$ coulomb

$k=9\times10^9\ Nm^2/C^2$

Thus, radius of 1st orbit

$=0.529\times10^{-8}\times1^2=0.529\times10^{-8}\ \text{cm}=0.529\times10^{-10}\ \text{m}$

Radius of 2nd orbit

$=0.529\times10^{-8}\times2^2=2.11\times10^{-8}\ \text{cm}=2.11\times10^{-10}\ \text{m}$

Radius of 3rd orbit

$=0.529\times10^{-8}\times3^2=4.76\times10^{-8}\ \text{cm}=4.76\times10^{-10}\ \text{m}$

and so on.

$$r_n=r_1\times n^2 \text{ for hydrogen atom}$$

and $$r_n=0.529\times\frac{n^2}{Z}\ \text{Å} \qquad \text{(for hydrogen like species)}$$

Energy of an electron: Let the total energy of the electron be E. It is the sum of kinetic energy and potential energy.

$$E = \text{kinetic energy} + \text{potential energy}$$
$$=\frac{1}{2}mv^2-\frac{kZe^2}{r}$$

Putting the value of mv^2 from eq. (i),

$$E=\frac{kZe^2}{2r}-\frac{kZe^2}{r}=-\frac{kZe^2}{2r}$$

Putting the value of r from eq. (iii),

$$E=-\frac{kZe^2}{2}\times\frac{4\pi^2mkZe^2}{n^2h^2}=-\frac{2\pi^2Z^2k^2me^4}{n^2h^2} \qquad \ldots\text{(iv)}$$

For hydrogen atom, $Z=1$

So, $$E=-\frac{2\pi^2k^2me^4}{n^2h^2}$$

Putting the values of π, k, m, e and h,

$$E=-\frac{2\times(3.14)^2\times(9\times10^9)^2\times(9.1\times10^{-31})\times(1.6\times10^{-19})^4}{n^2\times(6.625\times10^{-34})^2}$$

$$=-\frac{21.79\times10^{-19}}{n^2}\ \text{J per atom}$$

$$E=-\frac{R_H}{n^2} \qquad \text{(where, } R_H=2.18\times10^{-18}\text{ J)}$$

$$=-\frac{13.6}{n^2}\ \text{eV per atom} \qquad (1\ \text{J}=6.2419\times10^{18}\ \text{eV})$$

$$=-\frac{313.6}{n^2}\ \text{kcal/mol} \qquad (1\ \text{eV}=23.06\ \text{kcal/mol})$$

$$=-\frac{1312}{n^2}\ \text{kJ/mol}$$

$$\text{Kinetic energy in } n\text{th shell}=\frac{13.6\times Z^2}{n^2}\ \text{eV}$$

$$\text{Potential energy in } n\text{th shell}=\frac{-27.2\times Z^2}{n^2}\ \text{eV}$$

Substituting the values of $n=1,2,3,4,\ldots$, etc., the energy of electron in various energy shells in hydrogen atom can be calculated.

Energy shell	E (Joule per atom)	E (eV per atom)	E (kcal/mol)
1	-21.79×10^{-19}	-13.6	-313.6
2	-5.44×10^{-19}	-3.4	-78.4
3	-2.42×10^{-19}	-1.51	-34.84
4	-1.36×10^{-19}	-0.85	-19.6
—	—	—	—
—	—	—	—
∞	0	0	0

$$E_n=\frac{E_1}{n^2} \qquad \text{(for hydrogen atom)}$$

and $$E_n=E_1\times\frac{Z^2}{n^2} \qquad \text{(for hydrogen like species)}$$

where, E_1 = energy of hydrogen first orbit.

Since, n can have only integral values, it follows that total energy of the electron is quantised. The negative sign indicates that the electron is under attraction towards nucleus, *i.e.*, it is bound to the nucleus. The electron has minimum energy in the first orbit and its energy increases as n increases, *i.e.*, it becomes less negative. The electron can have a maximum energy value of zero when $n = \infty$. The zero energy means that the electron is no longer bound to the nucleus, *i.e.*, it is not under attraction towards nucleus.

For hydrogen like species such as He^+, Li^{2+}, etc., $E_n = Z^2 \times E_n$ for hydrogen atom.

Velocity of an electron: We know that,

Centrifugal force on electron

= force of attraction between nucleus and electron

$$\frac{mv^2}{r} = \frac{Ze^2}{r^2} \quad \text{(in CGS units)} \qquad \ldots \text{(i)}$$

The angular momentum of an electron is given as:

$$mvr = nh/2\pi \qquad \ldots \text{(ii)}$$

From eqs. (i) and (ii), we have

$$v\left(\frac{nh}{2\pi}\right) = Ze^2$$

$$v = \frac{Z}{n}\left(\frac{2\pi e^2}{h}\right)$$

$$v = \frac{Z}{n} \times 2.188 \times 10^8 \text{ cm/sec} \qquad \ldots \text{(iii)}$$

$$v = \frac{2.188 \times 10^8}{n} \text{ cm/sec} \quad \text{(For hydrogen, } Z = 1\text{)}$$

$$v_1 = 2.188 \times 10^8 \text{ cm/sec}$$

$$v_2 = \frac{1}{2} \times 2.188 \times 10^8 \text{ cm/sec}$$

$$v_3 = \frac{1}{3} \times 2.188 \times 10^8 \text{ cm/sec}$$

Here, v_1, v_2 and v_3 are the velocities of electron in first, second and third Bohr orbits in hydrogen.

From equation (iii),

$$\frac{v_1}{v_2} = \frac{2}{1} \quad \text{and} \quad \frac{v_3}{v_1} = \frac{1}{3} \text{ and so on.}$$

Orbital frequency: Number of revolutions per second by an electron in a shell is called **orbital frequency**; it may be calculated as,

Number of revolutions per second by an electron in a shell

$$= \frac{\text{Velocity}}{\text{Circumference}} = \frac{v}{2\pi r} = -\frac{E_1}{h}\left(\frac{2}{n^3}\right)$$

$$= \frac{Z^2}{n^3} \times 6.66 \times 10^{15}$$

where, E_1 = Energy of first shell.

Time period of revolution of electron in nth orbit (T_n):

$$T_n = \frac{2\pi r}{v_n} = \frac{n^3}{Z^2} \times 1.5 \times 10^{-16} \text{ sec}$$

Interpretation of hydrogen spectrum: The only electron in the hydrogen atom resides under ordinary conditions on the first orbit. When energy is supplied, the electron moves to higher energy shells depending on the amount of energy absorbed. When this electron returns to any of the lower energy shells, it emits energy. Lyman series is formed when the electron returns to the lowest energy state while Balmer series is formed when the electron returns to second energy shell. Similarly, Paschen, Brackett and Pfund series are formed when electron returns to the third, fourth and fifth energy shells from higher energy shells respectively (Fig. 2.12).

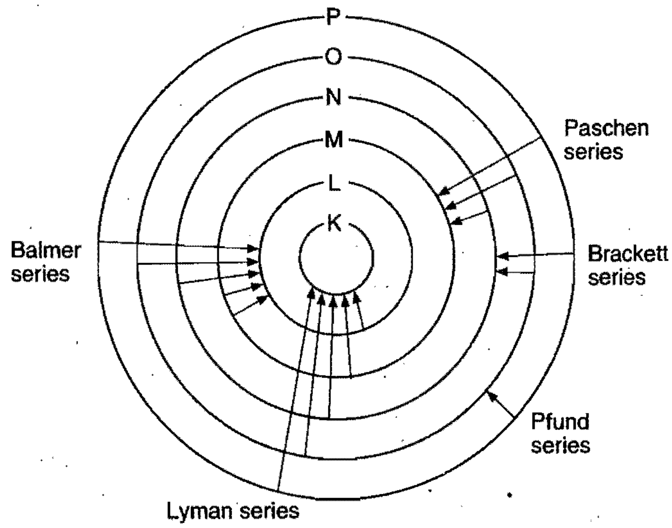


Fig. 2.12

Maximum number of lines produced when an electron jumps from n th level to ground level is equal to $\frac{n(n-1)}{2}$. For example, in the case of $n = 4$, number of lines produced is 6. ($4 \to 3, 4 \to 2, 4 \to 1, 3 \to 2, 3 \to 1, 2 \to 1$). When an electron returns from n_2 to n_1 state, the number of lines in the spectrum will be equal to.

$$\frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

If the electron comes back from energy level having energy E_2 to energy level having energy E_1, then the difference may be expressed in terms of energy of photon as:

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

or the frequency of the emitted radiation is given by

$$\nu = \frac{\Delta E}{h}$$

Since, ΔE can have only definite values depending on the definite energies of E_2 and E_1, ν will have only fixed values in an atom,

or $$\nu = \frac{c}{\lambda} = \frac{\Delta E}{h}$$

or $$\lambda = \frac{hc}{\Delta E}$$

Since, h and c are constants, ΔE corresponds to definite energy; thus, each transition from one energy level to another will produce a light of definite wavelength. This is actually observed as a line in the spectrum of hydrogen atom.

Thus, the different spectral lines in the spectra of atoms correspond to different transitions of electrons from higher energy levels to lower energy levels.

Derivation of Rydberg Equation

Let an excited electron from n_2 shell come to the n_1 shell with the release of radiant energy. The wave number $\bar{\nu}$ of the corresponding spectral line may be calculated in the following manner:

$$\Delta E = E_2 - E_1 = (-)\frac{2\pi^2 mZ^2e^4}{n_2^2 h^2} - (-)\frac{2\pi^2 mZ^2e^4}{n_1^2 h^2}$$

$$\frac{hc}{\lambda} = \frac{2\pi^2 mZ^2e^4}{h^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where, $\Delta E = h\nu = \frac{hc}{\lambda}$

$$\therefore \quad \bar{\nu} = \frac{1}{\lambda} = \frac{2\pi^2 mZ^2e^4}{ch^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

or $$\bar{\nu} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where, $R = \frac{2\pi^2 me^4}{ch^3}$ = Rydberg constant = 109743 cm^{-1}

This value of R is in agreement with experimentally determined value 109677.76 cm^{-1}. Rydberg equation for hydrogen may be given as,

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

Modification of Rydberg Equation

According to the Rydberg equation:

$$\underset{\text{wave number}}{\bar{\nu}} = \frac{2\pi^2 mZ^2e^4}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

It can be considered that the electron and the nucleus revolve around their common centre of mass. Therefore, instead of the mass of the electron, the reduced mass of the system was introduced and the equation becomes:

$$\bar{\nu} = \frac{2\pi^2 \mu Z^2e^4}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

Reduced mass 'μ' can be calculated as,

$$\frac{1}{\mu} = \frac{1}{m} + \frac{1}{M}$$

where, m = mass of electron

and M = mass of nucleus

$$\therefore \quad \mu = \frac{mM}{m+M}$$

(i) *First line of a series:* It is called 'line of longest wavelength' or 'line of shortest energy'.

For first line,

$$n_2 = (n_1 + 1)$$

$$\therefore \quad \bar{\nu}_{\text{first}} = \frac{1}{\lambda_{\text{first}}} = R\left[\frac{1}{n_1^2} - \frac{1}{(n_1+1)^2}\right]$$

Similarly for second, third and fourth lines,

$n_2 = n_1 + 2$; $n_2 = n_1 + 3$ and $n_2 = n_1 + 4$ respectively

$\therefore$ Rydberg equation may be written as,

$$\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{(n_1+x)^2}\right]$$

where, x = number of line in the spectrum.

e.g., $x = 1, 2, 3, 4, \ldots$ for first, second, third and fourth lines in the spectrum respectively.

(ii) *Series limit or last line of a series:* It is the line of shortest wavelength or line of highest energy.

For last line, $n_2 = \infty$

$$\bar{\nu}_{\text{last}} = \frac{1}{\lambda_{\text{last}}} = \frac{R}{n_1^2}$$

Lyman limit $= \frac{R}{1^2}$; Balmer limit $= \frac{R}{2^2}$

Paschen limit $= \frac{R}{3^2}$; Brackett limit $= \frac{R}{4^2}$

Pfund limit $= \frac{R}{5^2}$; Humphrey limit $= \frac{R}{6^2}$

(iii) *Intensities of spectral lines:* The intensities of spectral lines in a particular series decrease with increase in the value of n_2, *i.e.*, higher state.

e.g.,

Lyman series $\underset{(n_2 \to n_1)}{}$ $(2 \to 1) > (3 \to 1) > (4 \to 1) > (5 \to 1)$

Balmer series $\underset{(n_2 \to n_1)}{}$ $(3 \to 2) > (4 \to 2) > (5 \to 2) > (6 \to 2)$

$\xrightarrow{\hspace{6cm}}$

Decreasing intensity of the spectral lines

Ionization Energy and Excitation Energy

$$\text{Excitation potential for } n_1 \to n_2 = \frac{E_{n_2} - E_{n_1}}{\text{Electronic charge}}$$

$$\text{Ionization potential for } n_1 \to \infty = \frac{E_{n_1}}{\text{Electronic charge}}$$

The energy required to remove an electron from the ground state to form cation, *i.e.*, to take the electron to infinity, is called **ionization energy.**

$$\text{IE} = E_\infty - E_{\text{ground}}$$

$$\text{IE} = 0 - E_1(\text{H}) = 13.6\,\text{eV atom}^{-1}$$

$$= 2.17 \times 10^{-18}\ \text{J atom}^{-1}$$

$$\text{IE} = \frac{Z^2}{n^2} \times 13.6\,\text{eV}$$

$$\frac{I_1}{I_2} = \frac{Z_1^2}{n_1^2} \times \frac{n_2^2}{Z_2^2}$$

$$(\text{IE})_Z = \frac{(\text{IE})_H \times Z^2}{n^2}$$

If an electron is already present in the excited state, then the energy required to remove that electron is called **separation energy.**

$$E_{\text{separation}} = E_{\infty} - E_{\text{excited}}$$

The following points support Bohr theory:

(i) The frequencies of the spectral lines calculated from Bohr equation are in close agreement with the frequencies observed experimentally in hydrogen spectrum.

(ii) The value of Rydberg constant for hydrogen calculated from Bohr equation tallies with that determined experimentally.

(iii) The emission and absorption spectra of hydrogen like species such as He^+, Li^{2+} and Be^{3+} can be explained with the help of Bohr theory.

Limitations of Bohr Theory

(i) It does not explain the spectra of multi-electron atoms.

(ii) When a high resolving power spectroscope is used, it is observed that a spectral line in the hydrogen spectrum is not a simple line but a collection of several lines which are very close to one another. This is known as fine spectrum. Bohr theory does not explain the fine spectra of even the hydrogen atom.

(iii) It does not explain the splitting of spectral lines into a group of finer lines under the influence of magnetic field (Zeeman effect) and electric field (Stark effect).

(iv) Bohr theory is not in agreement with Heisenberg's uncertainty principle.

2.12 SOMMERFELD'S EXTENSION OF BOHR THEORY

To account for the fine spectrum of hydrogen atom, **Sommerfeld,** in 1915, proposed that the moving electron might describe elliptical orbits in addition to circular orbits and the nucleus is situated at one of the foci. During motion on a circle, only the angle of revolution changes while the distance from the nucleus remains the same but in elliptical motion both the angles of revolution and the distance of the electron from the nucleus change. The distance from the nucleus is termed as **radius vector** and the angle of revolution is known as **azimuthal angle.** The tangential velocity of the electron at a particular instant can be resolved into two components: one along the radius vector called **radial velocity** and the other perpendicular to the radius vector called **transverse** or **angular velocity.** These two velocities give rise to radial momentum and angular or azimuthal momentum. **Sommerfeld** proposed that both the momenta must be integral multiples of $\frac{h}{2\pi}$ [Fig. 2.13 (b)].

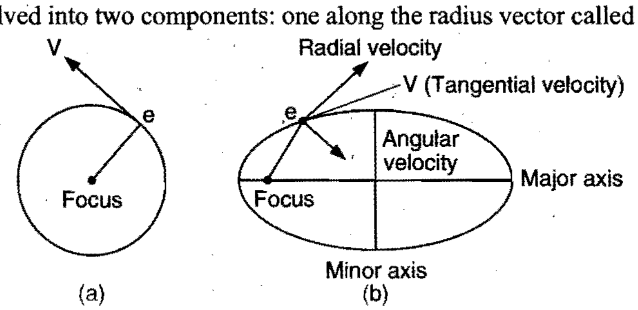


Fig. 2.13

$$\text{Radial momentum} = n_r \frac{h}{2\pi}$$

$$\text{Azimuthal momentum} = n_\phi \frac{h}{2\pi}$$

n_r and n_ϕ are related to the main orbit 'n' as:

$$n = n_r + n_\phi$$

or
$$\frac{n}{n_\phi} = \frac{n_r + n_\phi}{n_\phi} = \frac{\text{Length of major axis}}{\text{Length of minor axis}}$$

(i) 'n_ϕ' cannot be zero because under this condition, the ellipse shall take the shape of a straight line.

(ii) 'n_ϕ' cannot be more than 'n' because minor axis is always smaller than major axis.

(iii) 'n_ϕ' can be equal to 'n'. Under this condition, the major axis becomes equal to minor axis and the ellipse takes the shape of a circle. Thus, n_ϕ can have all integral values up to 'n' but not zero. When the values are less than 'n', orbits are elliptical and when it becomes equal to 'n', the orbit is circular in nature.

For $n = 1$, n_ϕ can have only one value, *i.e.*, 1. Therefore, the first orbit is circular in nature.

For $n = 2$, n_ϕ can have two values 1 and 2, *i.e.*, the second orbit has two sub-orbits, one is elliptical and the other is circular in nature.

For $n = 3$, n_ϕ can have three values 1, 2 and 3, *i.e.*, third orbit has three sub-orbits, two are elliptical and one is circular in nature.

For $n = 4$, n_ϕ can have four values 1, 2, 3 and 4, *i.e.*, fourth orbit has four sub-orbits, three are elliptical and fourth one is circular in nature (Fig. 2.14).

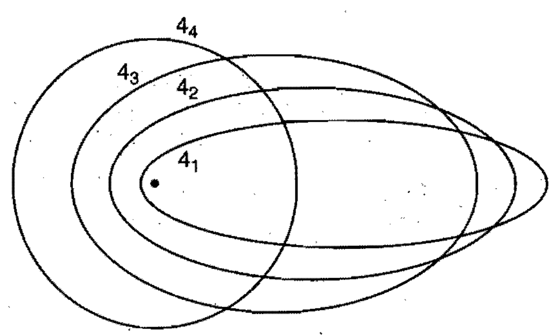


Fig. 2.14

Sommerfeld thus introduced the concept of subenergy shells. In a main energy shell, the energies of subshells differ slightly from one another. Hence, the jumping of an electron from one energy shell to another energy shell will involve slightly different amount of energy as it will depend on subshell also. This explains to some extent the fine spectrum of hydrogen atom. However, Sommerfeld extension fails to explain the spectra of multielectron atoms.

SOME SOLVED EXAMPLES

Example 18. *Calculate the wavelength and energy of radiation emitted for the electronic transition from infinite to stationary state of hydrogen atom. (Given, $R = 1.09678 \times 10^7\ m^{-1}$, $h = 6.6256 \times 10^{-34}$ J-s and $c = 2.9979 \times 10^8\ ms^{-1}$)*

Solution:
$$\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
$$n_1 = 1 \text{ and } n_2 = \infty$$
$$\frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{(\infty)^2}\right] = R$$
or
$$\lambda = \frac{1}{R} = \frac{1}{1.09678 \times 10^7} = 9.11 \times 10^{-8}\ \text{m}$$

We know that,
$$E = h\nu = h \cdot \frac{c}{\lambda} = 6.6256 \times 10^{-34} \times \frac{2.9979 \times 10^8}{9.11 \times 10^{-8}}$$
$$= 2.17 \times 10^{-18}\ \text{J}$$

Example 19. *Calculate the velocity (cm/sec) of an electron placed in the third orbit of the hydrogen atom. Also calculate the number of revolutions per second that this electron makes around the nucleus.*

Solution: Radius of 3rd orbit
$$= 3^2 \times 0.529 \times 10^{-8} = 4.761 \times 10^{-8}\ \text{cm}$$

We know that,
$$mvr = \frac{nh}{2\pi} \quad \text{or} \quad v = \frac{nh}{2\pi mr}$$
$$= \frac{3 \times 6.624 \times 10^{-27}}{2 \times 3.14 \times (9.108 \times 10^{-28}) \times (4.761 \times 10^{-8})}$$
$$= 0.729 \times 10^8\ \text{cm/sec}$$

Time taken for one revolution $= \frac{2\pi r}{v}$

Number of revolutions per second
$$= \frac{1}{\frac{2\pi r}{v}} = \frac{v}{2\pi r}$$
$$= \frac{0.729 \times 10^8}{2 \times 3.14 \times 4.761 \times 10^{-8}}$$
$$= 2.4 \times 10^{14}\ \text{revolutions/sec}$$

Example 20. *The electron energy in hydrogen atom is given by $E = -\frac{21.7 \times 10^{-12}}{n^2}$ erg. Calculate the energy required to remove an electron completely from $n = 2$ orbit. What is the longest wavelength (in cm) of light that be used to cause this transition?*

Solution:
$$E = -\frac{21.7 \times 10^{-12}}{n^2}\ \text{erg}$$

Electron energy in the 2nd orbit, *i.e.*, $n = 2$,
$$E_2 = -\frac{21.7 \times 10^{-12}}{2^2}\ \text{erg} = -5.425 \times 10^{-12}\ \text{erg}$$
and $E_\infty = 0$
$$\Delta E = \text{Change in energy} = E_\infty - E_2 = 5.425 \times 10^{-12}\ \text{erg}$$

Thus, energy required to remove an electron from 2nd orbit
$$= 5.425 \times 10^{-12}\ \text{erg}$$

According to quantum equation,
$$\Delta E = h \cdot \frac{c}{\lambda}$$
or
$$\lambda = \frac{hc}{\Delta E}$$
$$(h = 6.625 \times 10^{-27}\ \text{erg-sec};\ c = 3 \times 10^{10}\ \text{cm/sec})$$
and
$$\Delta E = 5.425 \times 10^{-12}\ \text{erg}$$
So,
$$\lambda = \frac{(6.625 \times 10^{-27}) \times (3 \times 10^{10})}{5.425 \times 10^{-12}}$$
$$= 3.7 \times 10^{-5}\ \text{cm}$$

Thus, the longest wavelength of light that can cause this transition is 3.7×10^{-5} cm.

Example 21. *Calculate the shortest and longest wavelengths in hydrogen spectrum of Lyman series.*

Or

Calculate the wavelengths of the first line and the series limit for the Lyman series for hydrogen. ($R_H = 109678\ cm^{-1}$)

Solution: For Lyman series, $n_1 = 1$.

For shortest wavelength in Lyman series (*i.e.*, series limit), the energy difference in two states showing transition should be maximum, *i.e.*, $n_2 = \infty$.

So,
$$\frac{1}{\lambda} = R_H\left[\frac{1}{1^2} - \frac{1}{\infty^2}\right] = R_H$$
$$\lambda = \frac{1}{109678} = 9.117 \times 10^{-6}\ \text{cm}$$
$$= 911.7\ \text{Å}$$

For longest wavelength in Lyman series (*i.e.*, first line), the energy difference in two states showing transition should be minimum, *i.e.*, $n_2 = 2$.

So,
$$\frac{1}{\lambda} = R_H\left[\frac{1}{1^2} - \frac{1}{2^2}\right] = \frac{3}{4}R_H$$
or
$$\lambda = \frac{4}{3} \times \frac{1}{R_H} = \frac{4}{3 \times 109678} = 1215.7 \times 10^{-8}\ \text{cm}$$
$$= 1215.7\ \text{Å}$$

Example 22. *Show that the Balmer series occurs between 3647 Å and 6563 Å. ($R = 1.0968 \times 10^7\ m^{-1}$)*

Solution: For Balmer series,
$$\frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{n^2}\right]$$

where, $n = 3, 4, 5, \ldots \infty$

To obtain the limits for Balmer series $n = 3$ and $n = \infty$ respectively.

$$\lambda_{max}\ (n=3) = \frac{1}{R\left[\frac{1}{2^2} - \frac{1}{3^2}\right]} = \frac{36}{5R}$$

$$= \frac{36}{5 \times 1.0968 \times 10^7}\ \text{m}$$

$$= 6563\ \text{Å}$$

$$\lambda_{min}\ (n=\infty) = \frac{1}{R\left[\frac{1}{2^2} - \frac{1}{\infty^2}\right]} = \frac{4}{R}$$

$$= \frac{4}{1.0968 \times 10^7}\ \text{m}$$

$$= 3647\ \text{Å}$$

Example 23. *Light of wavelength 12818 Å is emitted when the electron of a hydrogen atom drops from 5th to 3rd orbit. Find the wavelength of the photon emitted when the electron falls from 3rd to 2nd orbit.*

Solution: We know that,

$$\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

When, $n_1 = 3$ and $n_2 = 5$,

$$\frac{1}{12818} = R\left[\frac{1}{9} - \frac{1}{25}\right] = \frac{16R}{9 \times 25}$$

or $$12818 = \frac{9 \times 25}{16 \times R} \quad \ldots\text{(i)}$$

When, $n_1 = 2$ and $n_2 = 3$,

$$\frac{1}{\lambda} = R\left[\frac{1}{4} - \frac{1}{9}\right] = \frac{5R}{36}$$

$$\lambda = \frac{36}{5R} \quad \ldots\text{(ii)}$$

Dividing eqn. (ii) by eqn. (i),

$$\frac{\lambda}{12818} = \frac{36}{5R} \times \frac{16R}{9 \times 25} = \frac{64}{125}$$

$$\lambda = \frac{64}{125} \times 12818 = 6562.8\ \text{Å}$$

Example 24. *The ionisation energy of hydrogen atom is 13.6 eV. What will be the ionisation energy of He^+ and Li^{2+} ions?*

Solution: Ionisation energy = – (energy of the 1st orbit)

Energy of the 1st orbit of hydrogen $= -13.6$ eV

Energy of the 1st orbit of $He^+ = -13.6 \times Z^2$ (Z for $He^+ = 2$)

$= -13.6 \times 4$ eV $= -54.4$ eV

So, Ionisation energy of $He^+ = -(-54.4) = 54.4$ eV

Energy of 1st orbit of $Li^{2+} = -13.6 \times 9$ (Z for $Li^{2+} = 3$)

$= -122.4$ eV

Ionisation energy of $Li^{2+} = -(-122.4) = 122.4$ eV

Example 25. *If the energy difference between two electronic states is 46.12 kcal mol^{-1}, what will be the frequency of the light emitted when the electrons drop from higher to lower states? ($Nh = 9.52 \times 10^{-14}$ kcal sec mol^{-1}, where, N is the Avogadro's number and h is the Planck's constant)*

Solution: $\Delta E = 46.12$ kcal mol^{-1}

According to Bohr theory, $\Delta E = Nh\nu$

or $$\nu = \frac{\Delta E}{Nh} = \frac{46.12}{9.52 \times 10^{-14}}$$

$$= 4.84 \times 10^{14}\ \text{cycle sec}^{-1}$$

Example 26. *According to Bohr theory, the electronic energy of the hydrogen atom in the nth Bohr orbit is given by*

$$E_n = -\frac{21.76 \times 10^{-19}}{n^2}\ J$$

Calculate the longest wavelength of light that will be needed to remove an electron from the 3rd orbit of the He^+ ion.

(IIT 1990)

Solution: The electronic energy of He^+ ion in the nth Bohr orbit

$$= -\frac{21.76 \times 10^{-19}}{n^2} \times Z^2\ \text{J}$$

where, $Z = 2$

Thus, energy of He^+ in the 3rd Bohr orbit

$$= -\frac{21.76 \times 10^{-19}}{9} \times 4\ \text{J}$$

$$\Delta E = E_\infty - E_3$$

$$= 0 - \left[-\frac{21.76 \times 10^{-19} \times 4}{9}\right]$$

$$= \frac{21.76 \times 10^{-19} \times 4}{9}$$

We know that, $$\lambda = \frac{hc}{\Delta E} = \frac{6.625 \times 10^{-34} \times 3 \times 10^8 \times 9}{21.76 \times 10^{-19} \times 4}$$

$$= 2.055 \times 10^{-7}\ \text{m}$$

Example 27. *Calculate the ratio of the velocity of light and the velocity of electron in the first orbit of a hydrogen atom. (Given, $h = 6.624 \times 10^{-27}$ erg-sec; $m = 9.108 \times 10^{-28}$ g, $r = 0.529 \times 10^{-8}$ cm)*

Solution: $$v = \frac{h}{2\pi mr}$$

$$= \frac{6.624 \times 10^{-27}}{2 \times 3.14 \times 9.108 \times 10^{-28} \times 0.529 \times 10^{-8}}$$

$$= 2.189 \times 10^8\ \text{cm/sec}$$

$$\frac{c}{v} = \frac{3 \times 10^{10}}{2.189 \times 10^8} = 137$$

Example 28. *The wavelength of a certain line in Balmer series is observed to be 4341 Å. To what value of 'n' does this correspond? ($R_H = 109678\ cm^{-1}$)*

Solution:
$$\frac{1}{\lambda} = R_H\left[\frac{1}{2^2} - \frac{1}{n^2}\right]$$
$$\frac{1}{n^2} = \frac{1}{4} - \frac{1}{\lambda \times R_H}$$
$$= \frac{1}{4} - \frac{1}{4341 \times 10^{-8} \times 109678}$$
$$= 0.04$$
$$n^2 = \frac{1}{0.04} = 25$$
or
$$n = 5$$

Example 29. *Estimate the difference in energy between the first and second Bohr orbit for hydrogen atom. At what minimum atomic number would a transition from n = 2 to n = 1 energy level result in the emission of X-rays with $\lambda = 3.0 \times 10^{-8}$ m ? Which hydrogen-like species does this atomic number correspond to?* **(IIT 1993)**

Solution: $\Delta E = h\nu = \dfrac{h \cdot c}{\lambda}$

and
$$\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
$$\Delta E = R \cdot h \cdot c\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
$$\Delta E = h \cdot c \cdot \frac{3}{4} R$$
$$= \frac{6.625 \times 10^{-34} \times 3 \times 10^8 \times 1.09678 \times 10^7 \times 3}{4}$$
$$= 1.635 \times 10^{-18}\ J$$

For hydrogen-like species,
$$\Delta E = Z^2\ Rhc\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
$$\frac{1}{\lambda} = Z^2 R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
$$\frac{1}{3.0 \times 10^{-8}} = Z^2 \times 1.09678 \times 10^7 \times \left[\frac{1}{1^2} - \frac{1}{2^2}\right]$$
$$Z^2 = \frac{4}{3 \times 10^{-8} \times 1.09678 \times 10^7 \times 3} \approx 4$$
or $\quad Z = 2$

The species is He^+.

Example 30. *What transition in the hydrogen spectrum have the same wavelength as Balmer transition n = 4 to n = 2 of He^+ spectrum?* **(IIT 1993)**

Solution: For He^+ ion,
$$\frac{1}{\lambda} = Z^2 R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
$$= (2)^2\ R\left[\frac{1}{2^2} - \frac{1}{4^2}\right] = \frac{3R}{4}$$

For hydrogen atom,
$$\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
So,
$$\frac{3R}{4} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
or
$$\frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{3}{4}$$
i.e., $\quad n_1 = 1$ and $n_2 = 2$

Example 31. *Calculate the energy emitted when electrons of 1.0 g atom of hydrogen undergo transition giving the spectral line of lowest energy in the visible region of its atomic spectrum. ($R_H = 1.1 \times 10^7\ m^{-1}$; $c = 3 \times 10^8\ m\,s^{-1}$; $h = 6.62 \times 10^{-34}$ J-s)* **(IIT 1993)**

Solution: The transition occurs like Balmer series as spectral line is observed in visible region.

Thus, the line of lowest energy will be observed when transition occurs from 3rd orbit to 2nd orbit, *i.e.*, $n_1 = 2$ and $n_2 = 3$.
$$\frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{3^2}\right] = \frac{5}{36} R$$
$$E = h\nu = h \cdot \frac{c}{\lambda} = 6.62 \times 10^{-34} \times 3 \times 10^8 \times \frac{5}{36} \times 1.1 \times 10^7$$
$$= 3.03 \times 10^{-19}\ \text{J per atom}$$

Energy corresponding to 1.0 g atom of hydrogen
$$= 3.03 \times 10^{-19} \times \text{Avogadro's number}$$
$$= 3.03 \times 10^{-19} \times 6.023 \times 10^{23}\ J$$
$$= 18.25 \times 10^4\ J$$

Example 32. *How many times does the electron go around the first Bohr's orbit of hydrogen in one second?*

Solution: Number of revolutions per second $= \dfrac{v}{2\pi r}$ ... (i)
$$v = \frac{2.188 \times 10^8}{n}\ \text{cm/sec}$$
$$v = \frac{2.188 \times 10^8}{1} = 2.188 \times 10^8\ \text{cm/sec}$$
$$r = \frac{n^2}{Z} \times 0.529\ \text{Å}$$
$$= \frac{1^2}{1} \times 0.529 \times 10^{-8}\ \text{cm}$$
$$= 0.529 \times 10^{-8}\ \text{cm}$$

$\therefore$ Number of revolutions per sec $= \dfrac{2.188 \times 10^8}{2 \times 3.14 \times 0.529 \times 10^{-8}}$
$$= 6.59 \times 10^{15}$$

Example 33. *Calculate the wavelength of radiations emitted, produced in a line in Lyman series, when an electron falls from fourth stationary state in hydrogen atom.* ($R_H = 1.1 \times 10^7\ m^{-1}$) **(IIT 1995)**

Solution: $\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

$= 1.1 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{4^2}\right)$

$= 969.6 \times 10^{-10}$ metre

$\therefore \quad \lambda = 969.6\,\text{Å}$

Example 34. *What is the degeneracy of the level of the hydrogen atom that has the energy* $\left(-\frac{R_H}{9}\right)$?

Solution: $E_n = -\frac{R_H}{n^2} = -\frac{R_H}{9}$

$\therefore \quad n = 3$

Thus, $l = 0$ and $m = 0$ (one 3*s*-orbital)

$l = 1$ and $m = -1, 0, +1$ (three 3*p*-orbitals)

$l = 2$ and $m = -2, -1, 0, +1, +2$ (five 3*d*-orbitals)

Thus, degeneracy is nine ($1 + 3 + 5 = 9$ states).

Example 35. *Calculate the angular frequency of an electron occupying the second Bohr orbit of He^+ ion.*

Solution: Velocity of electron $(v) = \frac{2\pi Ze^2}{nh}$ ...(i)

Radius of He^+ ion in an orbit $(r_n) = \frac{n^2h^2}{4\pi^2 mZe^2}$ ...(ii)

Angular frequency or angular velocity (ω)

$$= \frac{v}{r_n} = \frac{2\pi Ze^2}{nh} \times \frac{4\pi^2 mZe^2}{n^2h^2} = \frac{8\pi^3 mZ^2e^4}{n^3h^3}$$

Given, $n = 2, m = 9.1 \times 10^{-28}$ g, $Z = 2, e = 4.8 \times 10^{-10}$ esu

$h = 6.626 \times 10^{-27}$ erg-sec

$$\therefore \quad \omega = \frac{8 \times \left(\frac{22}{7}\right)^3 \times 2^2 \times 9.1 \times 10^{-28} \times (4.8 \times 10^{-10})^4}{(2)^3 \times (6.626 \times 10^{-27})^3}$$

$= 2.067 \times 10^{16}\ \text{sec}^{-1}$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

6. If the speed of electron in first Bohr orbit of hydrogen be 'x', then speed of the electron in second orbit of He^+ is:
(a) $x/2$ (b) $2x$ (c) x (d) $4x$
[**Ans.** (c)]
[**Hint:** $v_n = \frac{v_1 \times Z}{n} = \frac{x \times 2}{2} = x$]

7. If first ionisation energy of hydrogen is E, then the ionisation energy of He^+ would be:
(a) E (b) $2E$ (c) $0.5E$ (d) $4E$
[**Ans.** (d)]
[**Hint:** $I_2(He^+) = Z^2 I_1(H)$
$= 2^2 \times E = 4E$]

8. The number of spectral lines that are possible when electrons in 7th shell in different hydrogen atoms return to the 2nd shell is:
(a) 12 (b) 15 (c) 14 (d) 10
[**Ans.** (b)]
[**Hint:** Number of spectral lines $= \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$
$= \frac{(7-2)(7-2+1)}{2} = 15$]

9. The ratio of radii of first orbits of H, He^+ and Li^{2+} is:
(a) $1:2:3$ (b) $6:3:2$ (c) $1:4:9$ (d) $9:4:1$
[**Ans.** (b)]
[**Hint:** $r = \frac{n^2}{Z} \times 0.529\,\text{Å}$
$r_H : r_{He^+} : r_{Li^{2+}}$
$1 : \frac{1}{2} : \frac{1}{3}$
$6 : 3 : 2$]

10. The energy of second orbit of hydrogen is equal to the energy of:
(a) fourth orbit of He^+ (b) fourth orbit of Li^{2+}
(c) second orbit of He^+ (d) second orbit of Li^{2+}
[**Ans.** (a)]
[**Hint:** $E = -\frac{Z^2}{n^2} \times 13.6$ eV
$E_2 = -\frac{13.6}{4}$ for 'H'
$E = -\frac{Z^2}{n^2} \times 13.6$ eV
$-\frac{13.6}{4} = -\frac{Z^2}{n^2} \times 13.6$
$\frac{Z^2}{n^2} = \frac{1}{4}$ $(Z = 1, n = 2)$]

11. What is the energy in eV required to excite the electron from $n = 1$ to $n = 2$ state in hydrogen atom? (n = principal quantum number) **[CET (J&K) 2006]**
(a) 13.6 (b) 3.4 (c) 17 (d) 10.2
[**Ans.** (d)]
[**Hint:** $\Delta E = E_2 - E_1$
$= \left(-\frac{13.6}{2^2}\right) - \left(-\frac{13.6}{1^2}\right)$
$= 13.6\left(1 - \frac{1}{4}\right) = \frac{3}{4} \times 13.6 = 10.2$ eV]

12. An electron in an atom undergoes transition in such a way that its kinetic energy changes from x to $\frac{x}{4}$, the change in potential energy will be:
(a) $+\frac{3}{2}x$ (b) $-\frac{3}{8}x$ (c) $+\frac{3}{4}x$ (d) $-\frac{3}{4}x$
[**Ans.** (a)]

[**Hint** : PE = − 2KE

$\therefore$ PE will change from $-2x$ to $-\frac{2x}{4}$

Change in potential energy $= \left(-\frac{2x}{4}\right) - (-2x)$

$= -\frac{x}{2} + 2x = \frac{3x}{2}$]

2.13 PARTICLE AND WAVE NATURE OF ELECTRON

In 1924, **de Broglie** proposed that an **electron,** like light, behaves both as a material particle and as a wave. This proposal gave birth to a new theory known as **wave mechanical theory of matter.** According to this theory, the electrons, protons and even atoms, when in motion, possess wave properties.

de Broglie derived an expression for calculating the wavelength of the wave associated with the electron.

According to Planck's equation,

$$E = h\nu = h\cdot\frac{c}{\lambda} \qquad \ldots \text{(i)}$$

The energy of a photon on the basis of Einstein's mass- energy relationship is

$$E = mc^2 \qquad \ldots \text{(ii)}$$

where, c is the velocity of the electron.

Equating both the equations, we get

$$h\frac{c}{\lambda} = mc^2$$

$$\lambda = \frac{h}{mc} = \frac{h}{p}$$

Momentum of the moving electron is inversely proportional to its wavelength.

Let kinetic energy of the particle of mass 'm' is E.

$$E = \frac{1}{2}mv^2$$

$$2Em = m^2v^2$$

$$\sqrt{2Em} = mv = p(\text{momentum})$$

$$\lambda = \frac{h}{p} = \frac{h}{\sqrt{2Em}}$$

Davisson and **Germer** made the following modification in de Broglie equation:

Let a charged particle, say an electron be accelerated with a potential of V; then the kinetic energy may be given as:

$$\frac{1}{2}mv^2 = eV$$

$$m^2v^2 = 2eVm$$

$$mv = \sqrt{2eVm} = p$$

$$\lambda = \frac{h}{\sqrt{2eVm}}$$

$$\lambda = \frac{h}{\sqrt{2qVm}} \text{ for charged particles of charge } q$$

de Broglie waves are not radiated into space, *i.e.*, they are always associated with electron. The wavelength decreases if the value of mass (m) increases, *i.e.*, in the case of heavier particles, the wavelength is too small to be measured. de Broglie equation is applicable in the case of smaller particles like electron and has no significance for larger particles.

(A) de Broglie wavelength associated with charged particles

(i) For electron:

$$\lambda = \frac{12.27}{\sqrt{V}} \text{Å}$$

(ii) For proton:

$$\lambda = \frac{0.286}{\sqrt{V}} \text{Å}$$

(iii) For α-particles:

$$\lambda = \frac{0.101}{\sqrt{V}} \text{Å}$$

where, V = accelerating potential of these particles.

(B) de Broglie wavelength associated with uncharged particles

(i) For neutrons:

$$\lambda = \frac{h}{\sqrt{2Em}} = \frac{6.62 \times 10^{-34}}{\sqrt{2 \times 1.67 \times 10^{-27} \times E}}$$

$$= \frac{0.286}{\sqrt{E\,(\text{eV})}} \text{Å}$$

(ii) For gas molecules:

$$\lambda = \frac{h}{m \times v_{\text{rms}}}$$

$$= \frac{h}{\sqrt{3mkT}}$$

where, k = Boltzmann constant

Bohr theory versus de Broglie equation: One of the postulates of Bohr theory is that angular momentum of an electron is an integral multiple of $\frac{h}{2\pi}$. This postulate can be derived with the help of de Broglie concept of wave nature of electron.

Consider an electron moving in a circular orbit around nucleus. The wave train would be associated with the circular orbit as shown in Fig. 2.15. If the two ends of an electron wave meet to give a regular series of crests and troughs, the electron wave is said to be in phase, *i.e.*, the circumference of Bohr orbit is equal to whole number multiple of the wavelength of the electron wave.

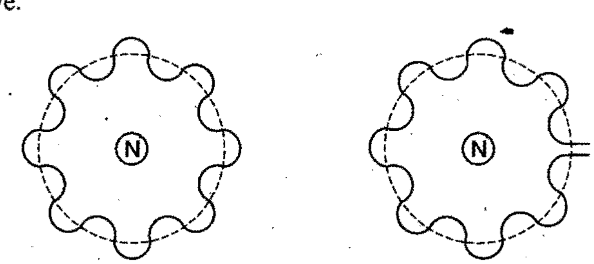


Fig. 2.15

So, $2\pi r = n\lambda$

or $$\lambda = \frac{2\pi r}{n} \quad \ldots \text{(i)}$$

From de Broglie equation,

$$\lambda = \frac{h}{mv} \quad \ldots \text{(ii)}$$

Thus, $$\frac{h}{mv} = \frac{2\pi r}{n}$$

or $$mvr = n \cdot \frac{h}{2\pi}$$ (v = velocity of electron and r = radii of the orbit)

i.e., $$\text{Angular momentum} = n \cdot \frac{h}{2\pi} \quad \ldots \text{(iii)}$$

This proves that the de Broglie and Bohr concepts are in perfect agreement with each other.

2.14 HEISENBERG UNCERTAINTY PRINCIPLE

Bohr theory considers an electron as a material particle. Its position and momentum can be determined with accuracy. But, when an electron is considered in the form of wave as suggested by de Broglie, it is not possible to ascertain simultaneously the exact position and velocity of the electron more precisely at a given instant since the wave is extending throughout a region of space. To locate the electron, radiation with extremely short wavelength is required. Radiation that has short wavelength is very energetic in nature. When it strikes the electron, the impact causes a change in the velocity of the electron. Thus, the attempt to locate the electron changes ultimately the momentum of the electron. Photons with longer wavelengths are less energetic and cause less effect on the momentum of the electron. Because of larger wavelength, such photons are not able to locate the position of an electron precisely.

In 1927, **Werner Heisenberg** presented a principle known as **Heisenberg uncertainty principle** which states: **"It is impossible to measure simultaneously the exact position and exact momentum of a body as small as an electron."**

The uncertainty of measurement of position, Δx and the uncertainty of momentum, Δp or Δmv are related by Heisenberg's relationship as:

$$\Delta x \cdot \Delta p \geq h / 4\pi$$

or $$\Delta x \cdot \Delta mv \geq h / 4\pi$$

where, h is Planck's constant.

For an electron of mass m (9.10×10^{-28} g), the product of uncertainty is quite large.

$$\Delta x \cdot \Delta v \geq \frac{6.626 \times 10^{-27}}{4\pi m}$$

$$\geq \frac{6.626 \times 10^{-27}}{4 \times 3.14 \times 9.10 \times 10^{-28}}$$

$$\approx 0.57 \text{ erg-sec per gram approximately}$$

$\Delta x \cdot \Delta v$ = uncertainty product

When $\Delta x = 0, \Delta v = \infty$ and *vice-versa*.

In the case of bigger particles (having considerable mass), the value of uncertainty product is negligible. If the position is known quite accurately, *i.e.*, Δx is very small, Δv becomes large and *vice-versa*. Thus, uncertainty principle is important only in the case of smaller moving particles like electrons.

For other canonical conjugates of motion, the equation for Heisenberg uncertainty principle may be given as:

$$\text{momentum} = \text{mass} \times \text{velocity}$$
$$= \text{mass} \times \frac{\text{velocity}}{\text{time}} \times \text{time}$$
$$= \text{force} \times \text{time}$$
$$\text{momentum} \times \text{distance} = \text{force} \times \text{distance} \times \text{time}$$
$$= \text{energy} \times \text{time}$$
$$\Delta p\, \Delta x = \Delta E\, \Delta t$$
$$\Delta E\, \Delta t \geq \frac{h}{4\pi} \quad \text{(for energy and time)}$$

Similarly, $$\Delta \phi\, \Delta \theta \geq \frac{h}{4\pi} \quad \text{(for angular motion)}$$

On the basis of this principle, therefore, Bohr picture of an electron in an atom, which gives a fixed position in a fixed orbit and definite velocity to an electron, is no longer tenable. The best we can think of in terms of probability of locating an electron with a probable velocity in a given region of space at a given time. The space or a three dimensional region round the nucleus where there is maximum probability of finding an electron of a specific energy is called an **atomic orbital**.

SOME SOLVED EXAMPLES

Example 36. *Calculate the wavelength associated with an electron moving with a velocity of 10^{10} cm per sec.*

Solution: Mass of the electron = 9.10×10^{-28} g

Velocity of electron = 10^{10} cm per sec

$h = 6.62 \times 10^{-27}$ erg-sec

According to de Broglie equation,

$$\lambda = \frac{h}{mv} = \frac{6.62 \times 10^{-27}}{9.10 \times 10^{-28} \times 10^{10}}$$
$$= 7.72 \times 10^{-10} \text{ cm}$$
$$= 0.0772 \text{ Å}$$

Example 37. *Calculate the uncertainty in the position of a particle when the uncertainty in momentum is:*
(a) 1×10^{-3} g cm sec^{-1} (b) zero.

Solution: (a) Given,

$$\Delta P = 1 \times 10^{-3} \text{ g cm sec}^{-1}$$
$$h = 6.62 \times 10^{-27} \text{ erg-sec}$$
$$\pi = 3.142$$

According to uncertainty principle,

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$$

So, $$\Delta x \geq \frac{h}{4\pi} \cdot \frac{1}{\Delta p} \geq \frac{6.62 \times 10^{-27}}{4 \times 3.142} \times \frac{1}{10^{-3}}$$
$$= 0.527 \times 10^{-24} \text{ cm}$$

(b) When the value of $\Delta p = 0$, the value of Δx will be infinity.

Example 38. *Calculate the momentum of a particle which has a de Broglie wavelength of* 2.5×10^{-10} *m.*
$(h = 6.6 \times 10^{-34}\ kg\ m^2\ s^{-1})$

Solution: Momentum $= \dfrac{h}{\lambda}$ (using de Broglie equation)

$$= \frac{6.6 \times 10^{-34}}{2.5 \times 10^{-10}}$$

$$= 2.64 \times 10^{-24} \text{ kg m sec}^{-1}$$

Example 39. *What is the mass of a photon of sodium light with a wavelength of* $5890\ \text{Å}$*?*
$(h = 6.63 \times 10^{-27}\ erg\text{-}sec, c = 3 \times 10^{10}\ cm/sec)$

Solution: $\lambda = \dfrac{h}{mc}$

or $m = \dfrac{h}{\lambda c}$

So, $m = \dfrac{6.63 \times 10^{-27}}{5890 \times 10^{-8} \times 3 \times 10^{10}}$

$= 3.752 \times 10^{-33}$ g

Example 40. *The uncertainty in position and velocity of a particle are* 10^{-10} *m and* $5.27 \times 10^{-24}\ ms^{-1}$ *respectively. Calculate the mass of the particle.* $(h = 6.625 \times 10^{-34}\ J\text{-}s)$

Solution: According to Heisenberg's uncertainty principle,

$$\Delta x \cdot m \Delta v = \frac{h}{4\pi}$$

or $$m = \frac{h}{4\pi\ \Delta x \cdot \Delta v}$$

$$= \frac{6.625 \times 10^{-34}}{4 \times 3.143 \times 10^{-10} \times 5.27 \times 10^{-24}}$$

$$= 0.099 \text{ kg}$$

Example 41. *Calculate the uncertainty in velocity of a cricket ball of mass* 150 g *if the uncertainty in its position is of the order of* 1 Å $(h = 6.6 \times 10^{-34}\ kg\ m^2\ s^{-1})$.

Solution: $\Delta x \cdot m \Delta v = \dfrac{h}{4\pi}$

$$\Delta v = \frac{h}{4\pi\ \Delta x \cdot m}$$

$$= \frac{6.6 \times 10^{-34}}{4 \times 3.143 \times 10^{-10} \times 0.150}$$

$$= 3.499 \times 10^{-24} \text{ ms}^{-1}$$

Example 42. *Find the number of waves made by a Bohr electron in one complete revolution in the 3rd orbit.* (IIT 1994)

Solution: Velocity of the electron in 3rd orbit $= \dfrac{3h}{2\pi mr}$

where, m = mass of electron and r = radius of 3rd orbit.

Applying de Broglie equation,

$$\lambda = \frac{h}{mv} = \frac{h}{m} \times \frac{2\pi mr}{3h} = \frac{2\pi r}{3}$$

No. of waves $= \dfrac{2\pi r}{\lambda} = \dfrac{2\pi r}{2\pi r} \times 3 = 3$

Example 43. *The kinetic energy of an electron is* 4.55×10^{-25} *J. Calculate the wavelength,* $(h = 6.6 \times 10^{-34}$ *J-sec, mass of electron* $= 9.1 \times 10^{-31}$ *kg).*

Solution: $KE = \dfrac{1}{2} mv^2 = 4.55 \times 10^{-25}$

or $\dfrac{1}{2} \times 9.1 \times 10^{-31} \times v^2 = 4.55 \times 10^{-25}$

or $\therefore\ v^2 = \dfrac{2 \times 4.55 \times 10^{-25}}{9.1 \times 10^{-31}}$

$v = 10^3 \text{ ms}^{-1}$

Applying de Broglie equation,

$$\lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34}}{9.1 \times 10^{-31} \times 10^3} = 0.72 \times 10^{-6} \text{ m}$$

Example 44. *The speeds of the Fiat and Ferrari racing cars are recorded to* $\pm 4.5 \times 10^{-4}\ m\,sec^{-1}$. *Assuming the track distance to be known within* $\pm 16\,m$, *is the uncertainty principle violated for a* 3500 *kg car?*

Solution: $\Delta x\, \Delta v = 4.5 \times 10^{-4} \times 16$

$= 7.2 \times 10^{-3} \text{ m}^2 \text{ sec}^{-1}$ ...(i)

$\dfrac{h}{4\pi m} = \dfrac{6.626 \times 10^{-34}}{4 \times 3.14 \times 3500}$ ...(ii)

$= 1.507 \times 10^{-38}$

Since, $\Delta x\, \Delta v \geq h/4\pi m$

Hence, Heisenberg uncertainty principle is not violated.

Example 45. *Alveoli are tiny sacs in the lungs whose average diameter is* 5×10^{-5} *m. Consider an oxygen molecule* $(5.3 \times 10^{-26}$ *kg) trapped within a sac. Calculate uncertainty in the velocity of oxygen molecule.*

Solution: Uncertainty in position Δx = Diameter of Alveoli

$= 5 \times 10^{-10}$ m

$$\Delta x\, \Delta v \geq \frac{h}{4\pi m}$$

$$\Delta v \geq \frac{6.626 \times 10^{-34}}{4 \times 3.14 \times 5.3 \times 10^{-26} \times 5 \times 10^{-10}}$$

$$\Delta v \approx 1.99 \text{ m/sec}$$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

13. If the kinetic energy of an electron is increased 4 times, the wavelength of the de Broglie wave associated with it would become:

(a) 4 times (b) 2 times

(c) $\dfrac{1}{2}$ times (d) $\dfrac{1}{4}$ times

[Ans. (c)]

[Hint: $\lambda = \dfrac{h}{\sqrt{2Em}}$, where, E = kinetic energy

When, the kinetic energy of electron becomes 4 times, the de Broglie wavelength will become half.]

14. The mass of photon having wavelength 1 nm is:

(a) 2.21×10^{-35} kg (b) 2.21×10^{-33} g

(c) 2.21×10^{-33} kg (d) 2.21×10^{-26} kg

[**Ans.** (c)]

[**Hint:** $\lambda=\frac{h}{mc}$

$$m=\frac{h}{\lambda c}=\frac{6.626\times10^{-34}}{1\times10^{-9}\times3\times10^{8}}$$

$$=2.21\times10^{-33}\text{ kg}]$$

15. The de Broglie wavelength of 1 mg grain of sand blown by a 20 ms^{-1} wind is:

(a) 3.3×10^{-29} m (b) 3.3×10^{-21} m

(c) 3.3×10^{-49} m (d) 3.3×10^{-42} m

[**Ans.** (a)]

[**Hint:** $\lambda=\frac{h}{mv}=\frac{6.626\times10^{-34}}{1\times10^{-6}\times20}=3.313\times10^{-29}$ m]

16. In an atom, an electron is moving with a speed of 600 m sec^{-1} with an accuracy of 0.005%. Certainty with which the position of the electron can be located is:

($h=6.6\times10^{-34}$ kg m^2 s^{-1}, mass of electron $=9.1\times10^{-31}$ kg)

(AIEEE 2009)

(a) 1.52×10^{-4} m (b) 5.1×10^{-3} m

(c) 1.92×10^{-3} m (d) 3.84×10^{-3} m

[**Ans.** (c)]

[**Hint:** Accuracy in velocity = 0.005%

$$\Delta v=\frac{600\times0.005}{100}=0.03$$

According to Heisenberg's uncertainty principle,

$$\Delta x\, m\Delta v\geq\frac{h}{4\pi}$$

$$\Delta x=\frac{6.6\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times0.03}$$

$$=1.92\times10^{-3}\text{ m}]$$

17. Velocity of de Broglie wave is given by:

(a) $\frac{c^2}{\nu}$ (b) $\frac{h\nu}{mc}$ (c) $\frac{mc^2}{h}$ (d) $\nu\lambda$

[**Ans.** (b)]

[**Hint:** $\lambda=\frac{h}{mv}=\frac{h}{p}$

$$p=\frac{h}{\lambda}$$

$$mv=\frac{h\nu}{c}$$

$$v=\frac{h\nu}{mc}]$$

2.15 WAVE MECHANICAL MODEL OF ATOM

The atomic model which is based on the particle and wave nature of the electron is known as **wave mechanical model of the atom.** This was developed by **Erwin Schrödinger** in 1926. This model describes the electron as a three-dimensional wave in the electronic field of positively charged nucleus. Schrödinger derived an equation which describes wave motion of an electron. The differential equation is:

$$\frac{d^2\psi}{dx^2}+\frac{d^2\psi}{dy^2}+\frac{d^2\psi}{dz^2}+\frac{8\pi^2m}{h^2}(E-V)\psi=0$$

where, x, y and z are cartesian coordinates of the electron; m = mass of the electron; E = total energy of the electron; V = potential energy of the electron; h = Planck's constant and ψ(psi) = wave function of the electron.

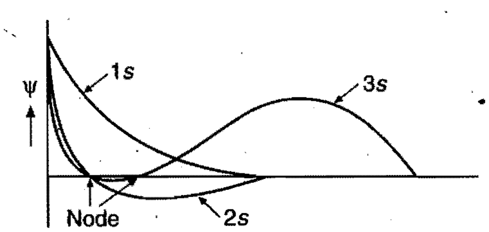
Significance of ψ: The wave function may be regarded as the amplitude function expressed in terms of coordinates x, y and z. The wave function may have positive or negative values depending upon the values of coordinates.

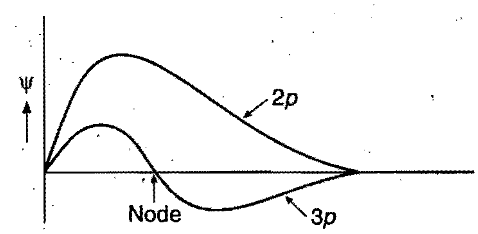
The main aim of Schrödinger equation is to give a solution for the probability approach. When the equation is solved, it is observed that for some regions of space the value of ψ is positive and for other regions the value of ψ is negative. But the probability must be always positive and cannot be negative. It is, thus, proper to use ψ^2 in favour of ψ.

Significance of ψ^2: ψ^2 is a probability factor. It describes the probability of finding an electron within a small space. The space in which there is maximum probability of finding an electron is termed as orbital.

The solution of the wave equation is beyond the scope of this book. The important point of the solution of this equation is that it provides a set of numbers, called **quantum numbers,** which describe energies of the electrons in atoms, information about the shapes and orientations of the most probable distribution of electrons around the nucleus.

Wave function ψ can be plotted against distance 'r' from nucleus as,





For hydrogen wave function, number of nodes can be calculated as,

(i) Number of radial nodes = $(n - l - 1)$
(ii) Number of angular nodes = l
(iii) Total number of nodes = $(n - 1)$
(iv) Number of nodal planes = l

Note: If the node at $r = \infty$ is also considered then no. of nodes will be 'n' (not $n - 1$).

Examples: *(i) For 1s-orbital $n = 1$, $l = 0$, it will have no radial or angular node.*

(ii) For 2s-orbital, $n = 2$, $l = 0$, it will have only one radial node.

(iii) For 3s-orbital, $n = 3$, $l = 0$, it will have two radial nodes.

(iv) For 2p-orbital, $n = 2$, $l = 1$, it will have no radial node but it has only one angular node.

(v) For 3p-orbital, $n = 3$, $l = 1$, it will have one radial and one angular node.

For *s*-orbitals:

$(n - 1)$ radial nodes + 0 angular node = $(n - 1)$ total nodes.

For *p*-orbitals:

$(n - 2)$ radial nodes + 1 angular node = $(n - 1)$ total nodes.

For *d*-orbitals:

$(n - 3)$ radial nodes + 2 angular nodes = $(n - 1)$ total nodes.

d_{z^2} like all d-orbitals has two angular nodes. The difference is that the angular nodes are cones in a d_{z^2} orbital, not planes.

Operator form Schrödinger Wave Equation

$$\hat{H}\Psi = E\Psi \qquad \text{(Operator form)}$$

where $\hat{H} = \left[-\dfrac{h^2}{8\pi^2 m}\nabla^2 + \hat{V}\right]$ = Hamiltonian operator

$$= \hat{T} + \hat{V}$$

Here, $\hat{T}$ = Kinetic energy operator

$\hat{V}$ = Potential energy operator

Complete wave function can be given as

$$\Psi(r,\theta,\phi) = \underbrace{R(r)}_{\text{Radial part}} \ \underbrace{\Theta(\theta)\Phi(\phi)}_{\text{Angular part}}$$

Dependence of the wave function on quantum number can be given as,

$$\Psi_{nlm}(r,\theta,\phi) = R_{nl}(r)\ \Theta_{lm}(\theta)\Phi_m(\phi)$$

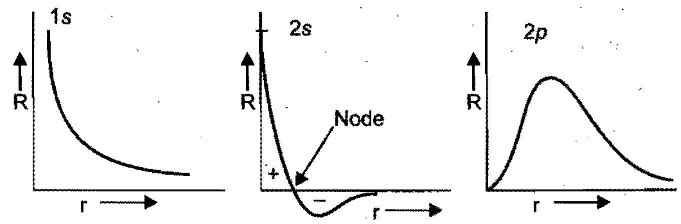
The function R depend only on r, therefore they describe the distribution of the electron as a function of r from the nucleus. These functions depend upon two quantum numbers, n and l. The two functions Θ and Φ taken together give the angular distribution of the electron.

The radial part of the wave function for some orbitals may be given as,

	n	l	R_{nl}
1s	1	0	$2\left(\dfrac{Z}{a_0}\right)^{3/2} e^{-Zr/a_0}$
2s	2	0	$\left(\dfrac{Z}{2a_0}\right)^{3/2}\left(2-\dfrac{Zr}{a_0}\right)e^{-Zr/2a_0}$
2p	2	1	$\dfrac{1}{\sqrt{3}}\left(\dfrac{Zr}{2a_0}\right)^{3/2}\left(\dfrac{Zr}{a_0}\right)e^{-Zr/2a_0}$

where, Z = atomic number, a_0 = radius of first Bohr orbit of hydrogen.

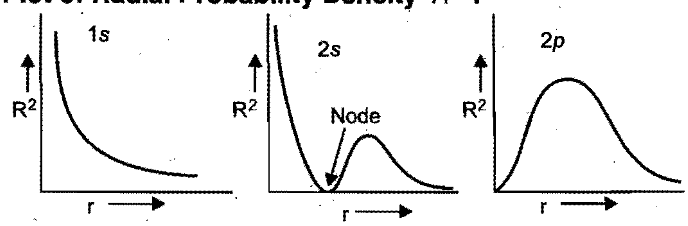
Plot of Radial Wave Function '*R*':



Number of radial nodes = $(n - l - 1)$.

At node, the value of 'R' changes from positive to negative.

Plot of Radial Probability Density 'R^2':



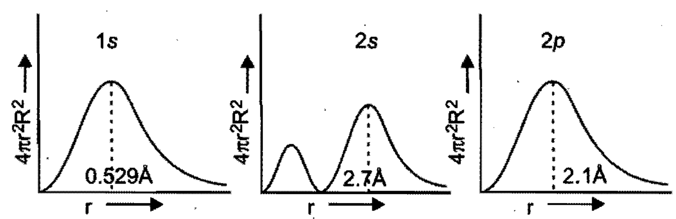
The plots of probability, *i.e.*, R^2 or Ψ^2 are more meaningful than the plots of functions themselves. It can be seen that for both 1s and 2s orbitals, the probability has a maximum value at $r = 0$, *i.e.*, in the nucleus. In case of 2s orbital, one more maximum in the probability plot is observed.

Plot of Radial Probability Function ($4\pi r^2 R^2$):

In order to visualize the electron cloud within a spherical shell is placed at radii 'r' and '$r + dr$' from the nucleus. Thus radial probability function describes the total probability of finding the electron in a spherical shell of thickness 'dr' located at the distance r from the nucleus.

R.P.F. = (Volume of spherical shell) × Probability density

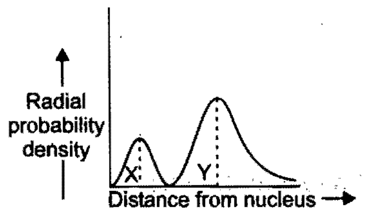
$$= (4\pi r^2 dr) \times R^2$$



In the plot of radial probability against 'r', number peaks, *i.e.*, region of maximum probability = $n - l$.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

18.



If the above radial probability curve indicates '2s' orbital, the distance between the peak points X, Y is :

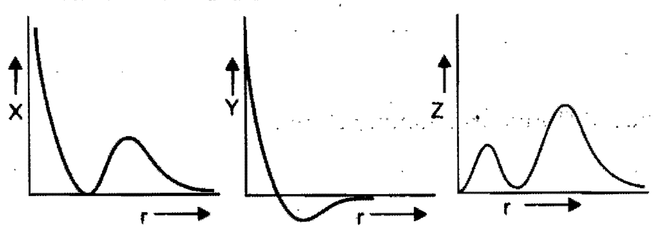
(a) 2.07Å (b) 1.59Å (c) 0.53Å (d) 2.12Å

[Ans. (a)]

[Hint : $X = 0.53\text{Å}$, $Y = 2.6\text{Å}$

$Y - X = 2.6 - 0.53 = 2.07\text{Å}$]

19. Plots for 2s orbital are :



X, Y and Z are respectively

(a) R, R^2 and $4\pi r^2R^2$ (b) R^2, R and $4\pi r^2R^2$

(c) $4\pi r^2R^2$, R^2 and R (d) R^2, $4\pi r^2R^2$ and R

[Ans. (b)]

[Hint : Y will be definitely 'R' because value of R cannot be negative, thus X will be R^2 and Z will be $4\pi r^2R$. Z represents radial probability function; its value will be zero at origin]

20. The wave function (Ψ) of 2s is given by :

$$\Psi_{2s} = \frac{1}{2\sqrt{2\pi}}\left(\frac{1}{a_0}\right)^{1/2}\left\{2 - \frac{r}{a_0}\right\}e^{-r/2a_0}$$

At $r = r_0$, radial node is formed. Thus for 2s, r_0 in terms of a_0 is :

(a) $r_0 = a_0$ (b) $r_0 = 2a_0$ (c) $r_0 = a_0/2$ (d) $r_0 = 4a_0$

[Ans. (b)]

[Hint : When $r = r_0$, $\psi_{2s} = 0$, then from the given equation:

$$2 - \frac{r}{a_0} = 0$$

$$r = 2a_0\,]$$

21. The wave function for 1s orbital of hydrogen atom is given by :

$$\Psi_{1s} = \frac{\pi}{\sqrt{2}}e^{-r/a_0}$$

where, a_0 = Radius of first Bohr orbit

r = Distance from the nucleus (Probability of finding the electron varies with respect to it)

What will be the ratio of probabilities of finding the electrons at the nucleus to first Bohr's orbit a_0?

(a) e (b) e^2 (c) $1/e^2$ (d) zero

[Ans. (d)]

[Hint : For 1s orbital, probability of finding the electron at the nucleus is zero.]

22. The radial wave equation for hydrogen atom is :

$$\Psi = \frac{1}{16\sqrt{4}}\left(\frac{1}{a_0}\right)^{3/2}[(x-1)(x^2 - 8x + 12)]e^{-x/2}$$

where, $x = 2r/a_0$; a_0 = radius of first Bohr orbit.

The minimum and maximum position of radial nodes from nucleus are :

(a) $a_0, 3a_0$ (b) $\frac{a_0}{2}, 3a_0$ (c) $\frac{a_0}{2}, a_0$ (d) $\frac{a_0}{2}, 4a_0$

[Ans. (b)]

[Hint : At radial node, $\Psi = 0$

$\therefore$ From given equation,

$$x - 1 = 0 \text{ and } x^2 - 8x + 12 = 0$$

$$x - 1 = 0 \Rightarrow x = 1$$

i.e., $\frac{2r}{a_0} = 1$; $r = \frac{a_0}{2}$ (Minimum)

$$x^2 - 8x + 12 = 0$$

$$(x-6)(x-2) = 0$$

when $x - 2 = 0$

$$x = 2$$

$\frac{2r}{a_0} = 2$, i.e., $r = a_0$ (Middle value)

when $x - 6 = 0$

$$x = 6$$

$$\frac{2r}{a_0} = 6$$

$r = 3a_0$ (Maximum)]

2.16 QUANTUM NUMBERS

As we know, to search a particular person in this world, four things are needed:

(i) The country to which the person belongs

(ii) The city in that country to which the person belongs

(iii) The street in that city where the person is residing

(iv) The house number

Similarly, four identification numbers are required to locate a particular electron in an atom. These identification numbers are called **quantum numbers.** The four quantum numbers are discussed below.

Principal Quantum Number

It was given by **Bohr;** it is denoted by 'n'. It represents the name, size and energy of the shell to which the electron belongs.

The value of 'n' lies between 1 to ∞.

$$n = 1, 2, 3, 4, \ldots \infty$$

Value of n	=	1	2	3	4	5	6	7
Designation of shell	=	K	L	M	N	O	P	Q

(i) Higher is the value of 'n', greater is the distance of the shell from the nucleus.

$$r_1 < r_2 < r_3 < r_4 < r_5 < \ldots$$

$$r = \frac{n^2}{Z} \times 0.529 \text{ Å}$$

(ii) Higher is the value of 'n', greater is the magnitude of energy.

$$E_1 < E_2 < E_3 < E_4 < E_5 \ldots$$

$$E = -\frac{Z^2}{n^2} \times 21.69 \times 10^{-19} \text{ J / atom}$$

$$= -\frac{Z^2}{n^2} \times 313.3 \text{ kcal per mol}$$

Energy separation between two shells decreases on moving away from nucleus.

$$(E_2 - E_1) > (E_3 - E_2) > (E_4 - E_3) > (E_5 - E_4)$$

(iii) Maximum number of electrons in a shell* $= 2n^2$

(iv) Angular momentum can also be calculated using principal quantum number

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

Azimuthal Quantum Number

It was given by **Sommerfeld;** it is also called **angular quantum number, subsidiary quantum number** or **secondary quantum number.** It is denoted by 'l'; its value lies between 0, 1, 2, . . . $(n-1)$.

It describes the spatial distribution of electron cloud and angular momentum. It gives the name of the subshell associated with the main shell

$l = 0$ s-subshell; $l = 1$ p-subshell;
$l = 2$ d-subshell; $l = 3$ f-subshell;
$l = 4$ g-subshell.

s, p, d, f and g are spectral terms and signify *sharp, principal, diffused, fundamental* and *generalised* respectively.

The energies of the various subshells in the same shell are in the order of $s < p < d < f < g$ (increasing order). Subshells having equal l values but with different n values have similar shapes but their sizes increase as the value of 'n' increases. 2s-subshell is greater in size than 1s-subshell. Similarly $2p, 3p, 4p$-subshells have similar shapes but their sizes increase in the order $2p < 3p < 4p$.

Orbital angular momentum of an electron is calculated using the expression

$$\mu_l = \sqrt{l(l+1)}\,\frac{h}{2\pi} = \sqrt{l(l+1)}\,\hbar$$

here, $$\hbar = \frac{h}{2\pi}$$

The magnitude of magnetic moment μ_L may be given as:

$$\mu_L = \sqrt{l(l+1)} \text{ BM}$$

where, BM = Bohr Magneton

$$1 \text{ BM} = \frac{eh}{4\pi mc} = 9.2732 \times 10^{-14} \text{ J}$$

Maximum electrons present in a subshell $= 2(2l+1)$

s-subshell → 2 electrons d-subshell → 10 electrons
p-subshell → 6 electrons f-subshell → 14 electrons
g-subshell → 18 electrons

Magnetic Quantum Number

This quantum number is designated by the symbol 'm'. To explain splitting of a single spectral line into a number of closely spaced lines in the presence of magnetic field **(Zeeman effect),** Linde proposed that electron producing a single line has several possible space orientations for the same angular momentum vector in a magnetic field, *i.e.*, under the influence of magnetic field each subshell is further sub-divided into orbitals. Magnetic quantum number describes the orientation or distribution of electron cloud. For each value of 'l', the magnetic quantum number 'm' may assume all integral values from $-l$ to $+l$ including zero, *i.e.*, total $(2l+1)$ values.

Thus, when $l = 0, m = 0$ (only one value)
when $l = 1, m = -1, 0, +1$ (three values)
i.e., three orientations.

One orientation corresponds to one orbital. Three orientations (orbitals) are designated as p_x, p_y and p_z.

When $l = 2, m = -2, -1, 0, +1, +2$ (five values), *i.e.*, five orientations.

The five orbitals are designated as:

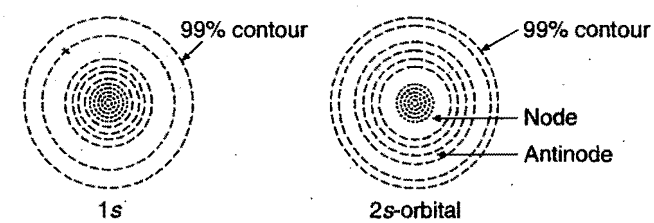
$$d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2} \text{ and } d_{z^2}.$$

When $l = 3, m = -3, -2, -1, 0, +1, +2, +3$ (seven values), *i.e.*, seven orientations.

Different values of 'm' for a given value of 'l' provide the total number of ways in which a given s, p, d and f subshells in presence of magnetic field can be arranged in space along x, y and z axes or total number of orbitals into which a given subshell can be divided.

When $l = 0, m = 0$, *i.e.*, one value implies that 's' subshell has only one space orientation and hence it can be arranged in space only in one way along x, y or z axes. Thus, 's' orbital has a symmetrical spherical shape and is usually represented as in Fig. 2.16.

In case of 1s-orbital, the electron cloud is maximum at the nucleus and decreases with the distance. The electron density at a particular distance is uniform in all directions. The region of maximum electron density is called **antinode.** In case of '2s'-orbital, the electron density is again maximum at the nucleus and decreases with increase in distance. The '2s'-orbital differs in detail from a 1s-orbital. The electron in a '2s'-orbital is likely to be found in two regions, one near the nucleus and other in a spherical shell about the nucleus. Electron density is zero in nodal region.



*No energy shell in atoms of known elements possesses more than 32 electrons.

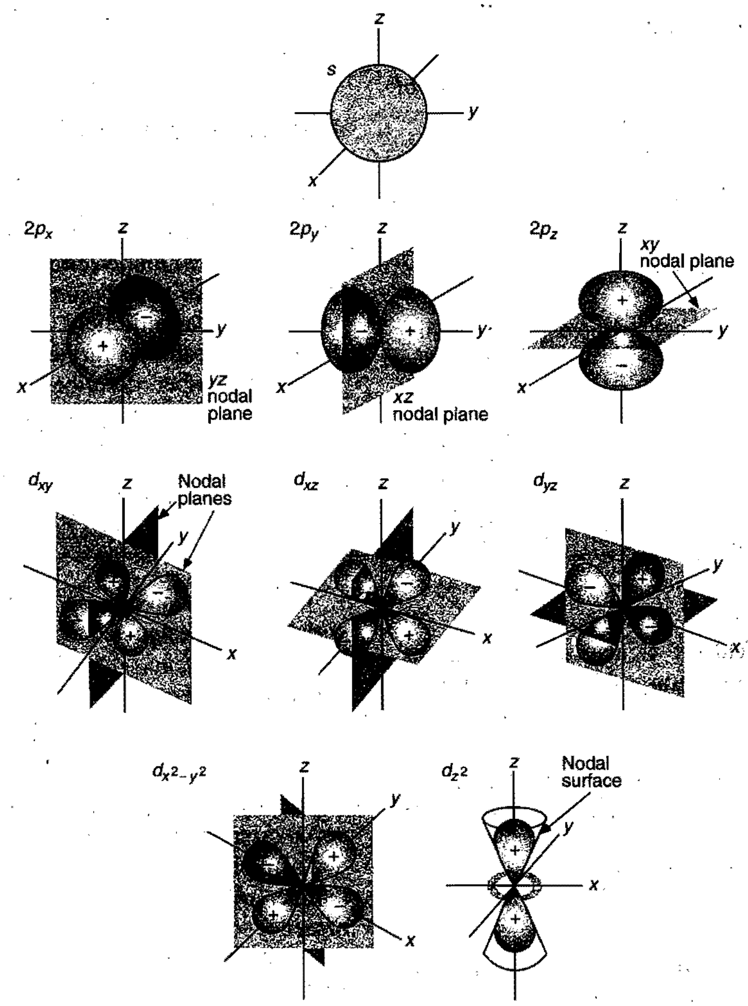


Fig. 2.16

When $l = 1$, 'm' has three values $-1, 0, +1$. It implies that 'p' subshell of any energy shell has three space orientations, *i.e.*, three orbitals. Each p-orbital has **dumb-bell** shape. Each one is disposed symmetrically along one of the three axes as shown in Fig. 2.16. p-orbitals have directional character.

Orbital	p_z	p_x	p_y
m	0	± 1	± 1
Nodal plane	xy	yz	zx

When $l = 2$, 'm' has five values $-2, -1, 0, +1, +2$. It implies that d-subshell of any energy shell has five orientations, *i.e.*, five orbitals. All the five orbitals are not identical in shape. Four of the d-orbitals $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ contain four lobes while fifth orbital d_{z^2} consists of only two lobes. The lobes of d_{xy} orbital lie between x and y-axes. Similar is the case for d_{yz} and d_{zx}. Four lobes of $d_{x^2-y^2}$ orbital are lying along x and y-axes while the two lobes of d_{z^2} orbital are lying along z-axis and contain a ring of negative charge surrounding the nucleus in xy-plane (Fig. 2.16).

Orbital	d_{xy}	d_{yz}	d_{zx}	$d_{x^2-y^2}$	d_{z^2}
m	± 2	± 1	± 1	± 2	0

Nodal planes :

Orbital	Nodal planes
d_{xy}	xz, yz
d_{yz}	xy, zx
d_{zx}	xy, yz
$d_{x^2-y^2}$	$x - y = 0, x + y = 0$
d_{z^2}	No nodal plane, it has a ring around the lobe

There are seven f-orbitals designated as $f_{x(x^2-y^2)}, f_{y(x^2-y^2)}, f_{z(x^2-y^2)}, f_{xyz}, f_{z^3}, f_{yz^2}$ and f_{xz^2}. Their shapes are complicated ones.

Positive values of m_l describes the orbital angular momentum component in the direction of applied magnetic field while the negative values of m_l are for the components in opposite direction to the applied magnetic field.

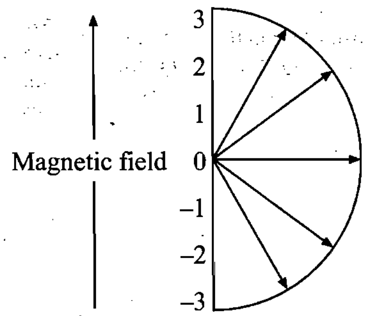


Fig. 21.6 (a) Space quantization in magnetic field

Characteristics of Orbitals

(i) All orbitals of the same shell in the absence of magnetic field possess same energy, *i.e.*, they are **degenerate.**

(ii) All orbitals of the same subshell differ in the direction of their space orientation.

(iii) Total number of orbitals in a main energy shell is equal to n^2 (but not more than 16 in any of the main shells of the known elements).

$n=1$ No. of orbitals $=(1)^2=1\,(1s)$

$n=2$ No. of orbitals $=(2)^2=4\,(2s, 2p_x, 2p_y, 2p_z)$

$n=3$ No. of orbitals $=(3)^2=9\,(3s, 3p_x, 3p_y, 3p_z, 3d_{xy}, 3d_{yz}, 3d_{zx}, 3d_{x^2-y^2}, 3d_{z^2})$

$n=4$ No. of orbitals $=(4)^2=16$

The division of main shells into subshells and that of subshell into orbitals has been shown below:

Note: Magnetic quantum number also represents quantized value of z-component of angular momentum of the electron in an orbital through the expression

$$Lz = m\left(\frac{h}{2\pi}\right)$$

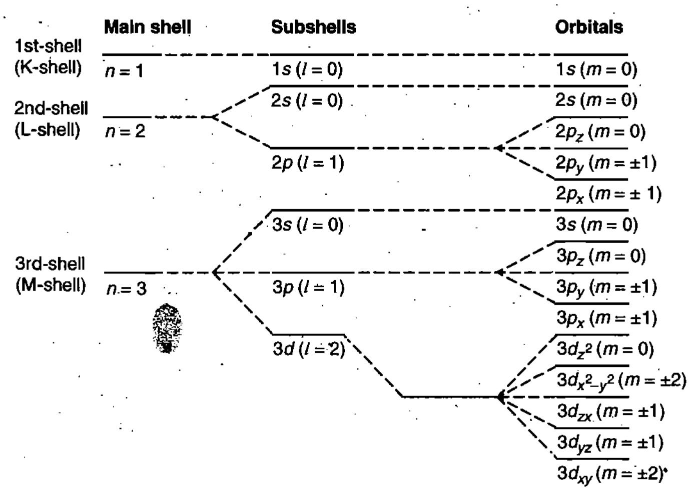
If θ is the angle between z-axis and angular momentum vector,

$$Lz = L\cos\theta$$

$$m\left(\frac{h}{2\pi}\right) = \sqrt{l(l+1)}\,\frac{h}{2\pi}\cos\theta$$

or

$$m = \sqrt{l(l+1)}\cos\theta$$



Degenerate Orbitals

Orbitals which are located at the same energy level on the energy level diagram are called degenerate orbitals. Thus, electrons have equal probability to occupy any of the degenerate orbitals.

p_x, p_y and $p_z \longrightarrow$ 3-fold degenerate

d-orbitals $\longrightarrow$ 5-fold degenerate

f-orbitals $\longrightarrow$ 7-fold degenerate

Degeneracy of p-orbitals remains unaffected in presence of external uniform magnetic field but degeneracy of d and f-orbitals is affected by external magnetic field.

Spin Quantum Number

It is denoted by 's' and it was given by **Goldschmidt.**

Spin quantum number represents the direction of electron spin around its own axis.

(i) For clockwise spin, $s = +\frac{1}{2}$ ($\uparrow$ arrow representation).

(ii) For anticlockwise spin, $s = -\frac{1}{2}$ ($\downarrow$ arrow representation).

Spin electron produces angular momentum equal to μ_s given by

$$\mu_s = \sqrt{s(s+1)}\,\frac{h}{2\pi}, \text{ where, } s = +\frac{1}{2}$$

Total spin of an atom $= n \times \frac{1}{2}$ (n = number of unpaired electrons)

Spin magnetic moment (μ_s) is given by

$$\mu_s = \sqrt{s(s+1)}\,\frac{eh}{2\pi mc}$$

Each orbital can accommodate two electrons with opposite spin or spin paired; paired electrons cancel the magnetic moment and develop mutual magnetic attraction as shown in the following Fig. 2.17.

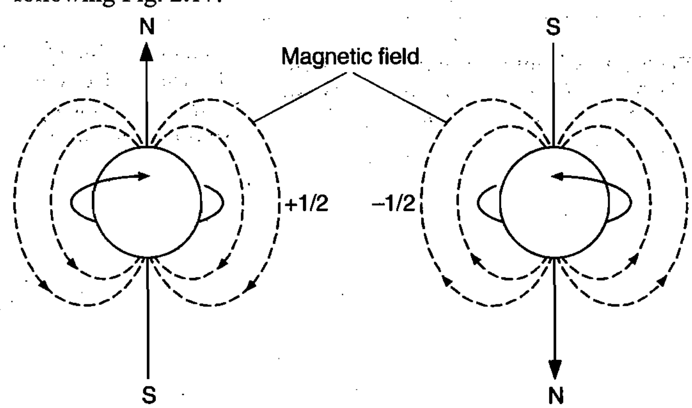


Fig. 2.17

Electrons having same spin are called spin parallel and those having opposite spin are called spin paired.

Spin paired $\longrightarrow$ $\boxed{\uparrow\downarrow}$

Spin parallel $\longrightarrow$ $\boxed{\uparrow\uparrow}$

Spin multiplicity:

Spin multiplicity $= [2\Sigma s + 1]$

where, s = spin quantum number

		$1s^2$	$2s^2$	$2p^2$		
e.g., carbon	=					
Normal state	=	↑↓	↑↓	↑	↑	
Excited state	=	↑↓	↑	↑	↑	↑

$$\text{Spin multiplicity} = 2\left[\left(+\frac{1}{2}\right)5+\left(-\frac{1}{2}\right)\right]+1=5$$

2.17 PAULI'S EXCLUSION PRINCIPLE

Each electron in an atom is designated by a set of four quantum numbers. In 1925, Pauli proposed that *no two electrons in an atom can have same values of all the four quantum numbers.*

An orbital accommodates two electrons with opposite spin; these two electrons have same values of principal, azimuthal and magnetic quantum number but the fourth, *i.e.*, spin quantum number will be different.

This principle can be illustrated by taking example of nitrogen.

$$N_7 = 1s^2 \; ; \; 2s^2 \; ; \; 2p^3$$

$$= 1s^2 \; ; \; 2s^2 \; ; \; 2p_x^1 \; 2p_y^1 \; 2p_z^1$$

$$= \boxed{\uparrow\downarrow} \; ; \boxed{\uparrow\downarrow} \; ; \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$$

Principal quantum number (n)	1	2	2	2	2
Azimuthal quantum number (l)	0	0	1	1	1
Magnetic quantum number (m)	0	0	+1	−1	0
Spin quantum number (s)	$+\frac{1}{2}-\frac{1}{2}$;	$+\frac{1}{2}-\frac{1}{2}$;	$+\frac{1}{2}$	$+\frac{1}{2}$	$+\frac{1}{2}$

Out of seven electrons no two have same values of all four quantum numbers. With the help of this principle, it is possible to calculate the maximum number of electrons which can be accommodated on main energy shells and subshells.

Principal Q. No. 'n'	Azimuthal Q. No. 'l'	Magnetic Q. No. 'm'	Spin Q. No. 's'	No. of electrons on a subshell	No. of electrons on a main shell
1	0(s)	0	$+\frac{1}{2},-\frac{1}{2}$	2	2
2	0(s)	0	$+\frac{1}{2},-\frac{1}{2}$	2	8
	1(p)	−1	$+\frac{1}{2},-\frac{1}{2}$	6	
		0	$+\frac{1}{2},-\frac{1}{2}$		
		+1	$+\frac{1}{2},-\frac{1}{2}$		
3	0(s)	0	$+\frac{1}{2},-\frac{1}{2}$	2	18
	1(p)	−1	$+\frac{1}{2},-\frac{1}{2}$	6	
		0	$+\frac{1}{2},-\frac{1}{2}$		
		+1	$+\frac{1}{2},-\frac{1}{2}$		
	2(d)	−2	$+\frac{1}{2},-\frac{1}{2}$	10	
		−1	$+\frac{1}{2},-\frac{1}{2}$		
		0	$+\frac{1}{2},-\frac{1}{2}$		
		+1	$+\frac{1}{2},-\frac{1}{2}$		
		+2	$+\frac{1}{2},-\frac{1}{2}$		
4	0(s)	0	$+\frac{1}{2},-\frac{1}{2}$	2	32
	1(p)			6	
	2(d)			10	
	3(f)	−3	$+\frac{1}{2},-\frac{1}{2}$	14	
		−2	$+\frac{1}{2},-\frac{1}{2}$		
		−1	$+\frac{1}{2},-\frac{1}{2}$		
		0	$+\frac{1}{2},-\frac{1}{2}$		
		+1	$+\frac{1}{2},-\frac{1}{2}$		
		+2	$+\frac{1}{2},-\frac{1}{2}$		
		+3	$+\frac{1}{2},-\frac{1}{2}$		

Conclusions:

(i) The maximum capacity of a main energy shell is equal to $2n^2$ electrons.

(ii) The maximum capacity of a subshell is equal to $2(2l+1)$ electrons.

Subenergy shell	Azimuthal Q. No. 'l'	Maximum capacity of electrons $2(2l+1)$
s	0	$2(2\times0+1)=2$
p	1	$2(2\times1+1)=6$
d	2	$2(2\times2+1)=10$
f	3	$2(2\times3+1)=14$

(iii) Number of subshells in a main energy shell is equal to the value of n.

Value of n	No. of subenergy shells	Designated as
1	1	1s
2	2	2s, 2p
3	3	3s, 3p, 3d
4	4	4s, 4p, 4d, 4f

(iv) Number of orbitals in a main energy shell is equal to n^2.

n	No. of orbitals	
1	$(1)^2=1$	s
2	$(2)^2=4$	s, p_x, p_y, p_z
3	$(3)^2=9$	$s, p_x, p_y, p_z, d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}, d_{z^2}$

(v) One orbital cannot have more than two electrons. If two electrons are present, their spins should be in opposite directions.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

23. The orbital angular momentum of an electron in a d-orbital is: **(DCE 2007)**

(a) $\sqrt{6}\,\frac{h}{2\pi}$ (b) $\sqrt{2}\,\frac{h}{2\pi}$ (c) $\frac{h}{2\pi}$ (d) $\frac{2h}{2\pi}$

[Ans. (a)]

[**Hint:** Orbital angular momentum $= \sqrt{l(l+1)}\,\dfrac{h}{2\pi}$

$$= \sqrt{2(2+1)}\,\frac{h}{2\pi} = \sqrt{6}\,\frac{h}{2\pi}$$

(Here, $l = 2$, for d-orbitals)]

24. Which of the following sets of quantum numbers is correct for an electron in $4f$-orbital?

(Jamia Millia Islamia Engg. Ent. 2007)

(a) $n = 4, l = 3, m = +4, s = +\frac{1}{2}$
(b) $n = 4, l = 4, m = -4, s = -\frac{1}{2}$
(c) $n = 4, l = 3, m = +1, s = +\frac{1}{2}$
(d) $n = 3, l = 2, m = -2, s = +\frac{1}{2}$

[**Ans.** (c)]

[**Hint:** For $4f$, $n = 4, l = 3, m = -3, -2, -1, 0, +1, +2, +3$
$s = -\frac{1}{2}$ or $+\frac{1}{2}$]

25. Match the List-I with List-II and select the correct set from the following sets given below:

List-I	List-II
(A) The number of sub-energy levels in an energy level	(1) n^2
(B) The number of orbitals in a sub-energy level	(2) $3d$
(C) The number of orbitals in an energy level	(3) $2l+1$
(D) $n = 3, l = 2, m = 0$	(4) n

[PET (Raj.) 2005]

Sets	(A)	(B)	(C)	(D)
(a)	4	3	1	2
(b)	3	1	2	4
(c)	1	2	3	4
(d)	3	4	1	2

[**Ans.** (a)]

[**Hint:** Number of orbitals in a shell $= n^2$

Number of subshells in a shell $= n$

Number of orbitals in a subshell $= (2l + 1)$

$n = 3, l = 2, m = 0$ represents $3d$]

26. Which of the following is not possible?

[BCECE (Medical) 2007]

(a) $n = 2, l = 1, m = 0$ (b) $n = 2, l = 0, m = -1$
(c) $n = 3, l = 0, m = 0$ (d) $n = 3, l = 1, m = -1$

[**Ans.** (b)]

[**Hint:** When $l = 0$, 'm' will also be equal to zero.]

27. What is the maximum number of electrons in an atom that can have the quantum numbers $n = 4$, $m_e = +1$?

[PMT (Kerala) 2007]

(a) 4 (b) 15 (c) 3 (d) 1 (e) 6

[**Ans.** (e)]

[**Hint:** $n = 4$; $l = 0$; $m_e = 0$
$l = 1$; $m_e = -1, 0, +1$
$l = 2$; $m_e = -2, -1, 0, +1, +2$
$l = 3$; $m_e = -3, -2, -1, 0, +1, +2, +3$

There are three orbitals having $m_e = +1$, thus maximum number of electrons in them will be 6.]

2.18 AUFBAU PRINCIPLE

Aufbau is a German word meaning 'building up'. This gives us a sequence in which various subshells are filled up depending on the relative order of the energy of the subshells. **The subshell with minimum energy is filled up first and when this obtains maximum quota of electrons, then the next subshell of higher energy starts filling.**

The sequence in which the various subshells are filled is the following:

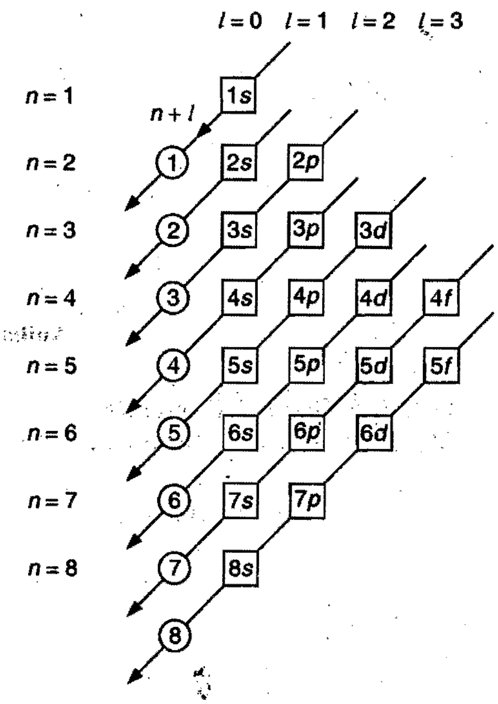


Fig. 2.18 Order of filling of various subshells

$1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p$.

The sequence in which various subshells are filled up can also be determined with the help of $(n + l)$ value for a given subshell. **The subshell with lowest $(n + l)$ value is filled up first. When two or more subshells have same $(n + l)$ value, the subshell with lowest value of 'n' is filled up first.**

Subshell	n	l	$(n+l)$	
1s	1	0	1	
2s	2	0	2	
2p	2	1	3	Lowest value of n
3s	3	0	3	
3p	3	1	4	Lowest value of n
4s	4	0	4	
3d	3	2	5	
4p	4	1	5	Lowest value of n
5s	5	0	5	

Subshell	n	l	$(n+l)$	
$4d$	4	2	6	Lowest value of n
$5p$	5	1	6	
$6s$	6	0	6	
$4f$	4	3	7	Lowest value of n
$5d$	5	2	7	
$6p$	6	1	7	
$7s$	7	0	7	
$5f$	5	3	8	Lowest value of n
$6d$	6	2	8	
$7p$	7	1	8	

The energy of electron in a hydrogen atom and other single electron species like He^+, Li^{2+} and Be^{3+} is determined solely by the principal quantum number. The energy of orbitals in hydrogen and hydrogen like species increases as follows:

$$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f < \ldots$$

The complete electronic configuration of all the known elements have been given in the table on next page. It is observed that few of the elements possess slightly different electronic configurations than expected on the basis of **Aufbau Principle.** These elements have been marked with asterisk (*) sign.

2.19 HUND'S RULE OF MAXIMUM MULTIPLICITY (Orbital Diagrams)

There is one more method of representing the electronic configuration which is usually called as **orbital diagram.** In this method, the electron is shown by an arrow: upward direction $\uparrow$ (clockwise spin) and downward direction $\downarrow$ (anti-clockwise spin).

To indicate the distribution of electrons among the orbitals of an atom, arrows are placed over bars that symbolise orbitals.

Hydrogen, for example, is represented as $\underset{1s}{\underline{\uparrow}}$. The next element with atomic number 2 is helium. It is represented as $\underset{1s}{\underline{\uparrow\downarrow}}$, *i.e.*, both the electrons are present on the same orbital 1s and are paired (spins are in opposite directions). The next two elements are Li and Be with three and four electrons, respectively. These are represented by orbital diagrams as:

Li $\quad \underset{1s}{\underline{\uparrow\downarrow}}\ \underset{2s}{\underline{\uparrow}}$

Be $\quad \underset{1s}{\underline{\uparrow\downarrow}}\ \underset{2s}{\underline{\uparrow\downarrow}}$

These can also be written as:

Li $\quad [He]\underset{2s}{\underline{\uparrow}}$

Be $\quad [He]\underset{2s}{\underline{\uparrow\downarrow}}$

In Be, 2s-orbital has been completed. The fifth electron in the case of boron enters the next available subshell which is $2p$. Thus, the electronic configuration of boron is $1s^2 2s^2 2p^1$. In the orbital diagram $[He]\underset{2s}{\underline{\uparrow\downarrow}} - \underset{2p}{\underline{\uparrow}} -$, the $2p$ subshell has three orbitals p_x, p_y and p_z. All the three have same energy. The electron can be accommodated on any one of the $2p$-orbitals. In the case of carbon, sixth electron is also accommodated on $2p$ subshell and its electronic configuration is represented as $1s^2 2s^2 2p^2$ but three orbital diagrams can be expected.

(i) $[He]\underset{2s}{\underline{\uparrow\downarrow}}\ \underbrace{\underline{\uparrow}\ \underline{\uparrow}\ \underline{\ \ }}_{2p}$ Electrons are present on two different orbitals with parallel spins.

(ii) $[He]\underset{2s}{\underline{\uparrow\downarrow}}\ \underbrace{\underline{\uparrow}\ \underline{\downarrow}\ \underline{\ \ }}_{2p}$ Electrons are present on two different orbitals with opposite spins.

(iii) $[He]\underline{\uparrow\downarrow}\ \underbrace{\underline{\uparrow\downarrow}\ \underline{\ \ }\ \underline{\ \ }}_{2p}$ Both the electrons are present on one orbital with opposite spins.

Experiments show that (i) orbital diagram is correct while (ii) and (iii) are not correct. This has given birth to a new rule known as **Hund's rule of maximum multiplicity.** It states that **electrons are distributed among the orbitals of a subshell in such a way as to give the maximum number of unpaired electrons with parallel spins.** Thus, the orbitals available in a subshell are first filled singly before they begin to pair. This means that pairing of electrons occurs with the introduction of second electron in s-orbitals, the fourth electron in p-orbitals, sixth electron in d-orbitals and eighth electron in f-orbitals. The orbital diagrams of nitrogen, oxygen, fluorine and neon are as given below:

			$2s$	$2p$		
Nitrogen	(7)	[He]	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	$\uparrow$
Oxygen	(8)	[He]	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$
Fluorine	(9)	[He]	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$
Neon	(10)	[He]	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$

The orbital diagrams of elements from atomic number 21 to 30 can be represented on similar lines as below:

		$3d$					$4s$
Sc	$[Ar]\,3d^1 4s^2$ [Ar]	$\uparrow$	—	—	—	—	$\uparrow\downarrow$
Ti	$[Ar]\,3d^2 4s^2$	$\uparrow$	$\uparrow$	—	—	—	$\uparrow\downarrow$
V	$[Ar]\,3d^3 4s^2$	$\uparrow$	$\uparrow$	$\uparrow$	—	—	$\uparrow\downarrow$
Cr	$[Ar]\,3d^5 4s^1$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$
Mn	$[Ar]\,3d^5 4s^2$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow\downarrow$
Fe	$[Ar]\,3d^6 4s^2$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow\downarrow$
Co	$[Ar]\,3d^7 4s^2$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow\downarrow$
Ni	$[Ar]\,3d^8 4s^2$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	$\uparrow\downarrow$
Cu	$[Ar]\,3d^{10} 4s^1$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$
Zn	$[Ar]\,3d^{10} 4s^2$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$

ELECTRONIC CONFIGURATION OF ELEMENTS

Element	At. No.	1s	2s	2p	3s	3p	3d	4s	4p	4d	4f	5s 5p 5d 5f
H	1	1										
He	2	2										(1s completed)
Li	3	2	1									
Be	4	2	2									(2s completed)
B	5	2	2	1								
C	6	2	2	2								
N	7	2	2	3								
O	8	2	2	4								
F	9	2	2	5								
Ne	10	2	2	6								(2p completed)
Na	11	2	2	6	1							
Mg	12	2	2	6	2							(3s completed)
Al	13	2	2	6	2	1						
Si	14	2	2	6	2	2						
P	15	2	2	6	2	3						
S	16	2	2	6	2	4						
Cl	17	2	2	6	2	5						
Ar	18	2	2	6	2	6						(3p completed)
K	19	2	2	6	2	6		1				
Ca	20	2	2	6	2	6		2				(4s completed)
Sc	21	2	2	6	2	6	1	2				
Ti	22	2	2	6	2	6	2	2				
V	23	2	2	6	2	6	3	2				
*Cr	24	2	2	6	2	6	5	1				
Mn	25	2	2	6	2	6	5	2				
Fe	26	2	2	6	2	6	6	2				
Co	27	2	2	6	2	6	7	2				
Ni	28	2	2	6	2	6	8	2				
*Cu	29	2	2	6	2	6	10	1				
Zn	30	2	2	6	2	6	10	2				(3d completed)
Ga	31	2	2	6	2	6	10	2	1			
Ge	32	2	2	6	2	6	10	2	2			
As	33	2	2	6	2	6	10	2	3			
Se	34	2	2	6	2	6	10	2	4			
Br	35	2	2	6	2	6	10	2	5			
Kr	36	2	2	6	2	6	10	2	6			(4p completed)
Rb	37	2	2	6	2	6	10	2	6			1
Sr	38	2	2	6	2	6	10	2	6			2 (5s completed)
Y	39	2	2	6	2	6	10	2	6	1		2
Zr	40	2	2	6	2	6	10	2	6	2		2
*Nb	41	2	2	6	2	6	10	2	6	4		1
*Mo	42	2	2	6	2	6	10	2	6	5		1
Tc	43	2	2	6	2	6	10	2	6	5		2
*Ru	44	2	2	6	2	6	10	2	6	7		1
*Rh	45	2	2	6	2	6	10	2	6	8		1
*Pd	46	2	2	6	2	6	10	2	6	10		
*Ag	47	2	2	6	2	6	10	2	6	10		1
Cd	48	2	2	6	2	6	10	2	6	10		2 (4d completed)
In	49	2	2	6	2	6	10	2	6	10		2 1
Sn	50	2	2	6	2	6	10	2	6	10		2 2
Sb	51	2	2	6	2	6	10	2	6	10		2 3
Te	52	2	2	6	2	6	10	2	6	10		2 4
I	53	2	2	6	2	6	10	2	6	10		2 5
Xe	54	2	2	6	2	6	10	2	6	10		2 6 (5p completed)

ELECTRONIC CONFIGURATION OF ELEMENTS

Element	At. No.	K	L	M	4s	4p	4d	4f	5s	5p	5d	5f	6s	6p	6d	6f	7s
Cs	55	2	8	18	2	6	10		2	6			1				
Ba	56	2	8	18	2	6	10		2	6			2	(6s completed)			
*La	57	2	8	18	2	6	10		2	6	1		2				
*Ce	58	2	8	18	2	6	10	1	2	6	1		2				
Pr	59	2	8	18	2	6	10	3	2	6			2				
Nd	60	2	8	18	2	6	10	4	2	6			2				
*Pm	61	2	8	18	2	6	10	5	2	6			2				
Sm	62	2	8	18	2	6	10	6	2	6			2				
Eu	63	2	8	18	2	6	10	7	2	6			2				
*Gd	64	2	8	18	2	6	10	7	2	6	1		2				
Tb	65	2	8	18	2	6	10	9	2	6			2				
Dy	66	2	8	18	2	6	10	10	2	6			2				
Ho	67	2	8	18	2	6	10	11	2	6			2				
Er	68	2	8	18	2	6	10	12	2	6			2				
Tm	69	2	8	18	2	6	10	13	2	6			2				
Yb	70	2	8	18	2	6	10	14	2	6			2				
Lu	71	2	8	18	2	6	10	14	2	6	1		2	(4f completed)			
Hf	72	2	8	18	32				2	6	2		2				
Ta	73	2	8	18	32				2	6	3		2				
W	74	2	8	18	32				2	6	4		2				
Re	75	2	8	18	32				2	6	5		2				
Os	76	2	8	18	32				2	6	6		2				
Ir	77	2	8	18	32				2	6	7		2				
*Pt	78	2	8	18	32				2	6	9		1				
*Au	79	2	8	18	32				2	6	10		1				
Hg	80	2	8	18	32				2	6	10		2	(5d completed)			
Tl	81	2	8	18	32				2	6	10		2	1			
Pb	82	2	8	18	32				2	6	10		2	2			
Bi	83	2	8	18	32				2	6	10		2	3			
Po	84	2	8	18	32				2	6	10		2	4			
At	85	2	8	18	32				2	6	10		2	5			
Rn	86	2	8	18	32				2	6	10		2	6(6p completed)			
Fr	87	2	8	18	32				2	6	10		2	6			1
Ra	88	2	8	18	32				2	6	10		2	6			2 (7s completed)
*Ac	89	2	8	18	32				2	6	10		2	6	1		2
*Th	90	2	8	18	32				2	6	10	0	2	6	2		2
*Pa	91	2	8	18	32				2	6	10	2	2	6	1		2
*U	92	2	8	18	32				2	6	10	3	2	6	1		2
*Np	93	2	8	18	32				2	6	10	4	2	6	1		2
Pu	94	2	8	18	32				2	6	10	6	2	6			2
Am	95	2	8	18	32				2	6	10	7	2	6			2
*Cm	96	2	8	18	32				2	6	10	7	2	6	1		2
*Bk	97	2	8	18	32				2	6	10	8	2	6	1		2
Cf	98	2	8	18	32				2	6	10	10	2	6			2
Es	99	2	8	18	32				2	6	10	11	2	6			2
Fm	100	2	8	18	32				2	6	10	12	2	6			2
Md	101	2	8	18	32				2	6	10	13	2	6			2
No	102	2	8	18	32				2	6	10	14	2	6			2
*Lr	103	2	8	18	32				2	6	10	14	2	6	1		2 (5f completed)
Ku or Rf	104	2	8	18	32				2	6	10	14	2	6	2		2
Ha or Db	105	2	8	18	32				2	6	10	14	2	6	3		2
Sg	106	2	8	18	32				2	6	10	14	2	6	4		2
Bh	107	2	8	18	32				2	6	10	14	2	6	5		2
Hs	108	2	8	18	32				2	6	10	14	2	6	6		2
Mt	109	2	8	18	32				2	6	10	14	2	6	7		2
*Uun or Ds	110	2	8	18	32				2	6	10	14	2	6	9		1
*Uuu or Rg	111	2	8	18	32				2	6	10	14	2	6	10		1
Uub	112	2	8	18	32				2	6	10	14	2	6	10		2 (6d completed)

Predicted electronic configurations

All those atoms which consist of at least one of the orbitals singly occupied behave as **paramagnetic** materials because these are weakly attracted to a magnetic field, while all those atoms in which all the orbitals are doubly occupied behave as **diamagnetic** materials because they have no attraction for magnetic field. However, these are slightly repelled by magnetic field due to induction.

Magnetic moment may be calculated as,

$$\mu = \sqrt{n(n+2)}\ \text{BM}$$

$$1\ \text{BM (Bohr Magneton)} = \frac{eh}{4\pi mc}$$

where, n = no. of unpaired electron

Exceptions to Aufbau Principle

In some cases, it is seen that actual electronic arrangement is slightly different from arrangement given by aufbau principle. A simple reason behind this is that half-filled and full-filled subshells have got extra stability.

Cr_{24}	$\longrightarrow$	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^4, 4s^2$	(wrong)
	$\longrightarrow$	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^5, 4s^1$	(right)
Cu_{29}	$\longrightarrow$	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^9, 4s^2$	(wrong)
	$\longrightarrow$	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^1$	(right)

Similarly the following elements have slightly different configurations than expected:

$$Nb_{41} \longrightarrow [Kr]4d^4 5s^1$$
$$Mo_{42} \longrightarrow [Kr]4d^5 5s^1$$
$$Ru_{44} \longrightarrow [Kr]4d^7 5s^1$$
$$Rh_{45} \longrightarrow [Kr]4d^8 5s^1$$
$$Pd_{46} \longrightarrow [Kr]4d^{10} 5s^0$$
$$Ag_{47} \longrightarrow [Kr]4d^{10} 5s^1$$
$$Pt_{78} \longrightarrow [Xe]4f^{14} 5d^9 6s^1$$
$$Au_{79} \longrightarrow [Xe]4f^{14} 5d^{10} 6s^1$$
$$La_{57} \longrightarrow [Kr]\, 4d^{10} 5s^2 5p^6 5d^1 6s^2$$
$$Ce_{58} \longrightarrow [Kr]\, 4d^{10} 4f^2 5s^2 5p^6 5d^0 6s^2$$
$$Gd_{64} \longrightarrow [Kr]\, 4d^{10} 4f^7 5s^2 5p^6 5d^1 6s^2$$

2.20 PHOTOELECTRIC EFFECT

Emission of electrons from a metal surface when exposed to light radiations of appropriate wavelength is called **photoelectric effect**. The emitted electrons are called **photoelectrons**.

Work function or threshold energy may be defined as the minimum amount of energy required to eject electrons from a metal surface.

According to Einstein,

Maximum kinetic energy of the ejected electron
= absorbed energy − work function

$$\frac{1}{2} mv^2_{\max} = h\nu - h\nu_0$$

$$= hc\left[\frac{1}{\lambda} - \frac{1}{\lambda_0}\right]$$

where, ν_0 and λ_0 are threshold frequency and threshold wavelength respectively.

Stopping potential: The minimum potential at which the plate photoelectric current becomes zero is called **stopping potential.**

If V_0 is the stopping potential, then

$$eV_0 = h(\nu - \nu_0)$$

Laws of Photoelectric Effect

(i) Rate of emission of photoelectrons from a metal surface is directly proportional to the intensity of incident light.

(ii) The maximum kinetic energy of photoelectrons is directly proportional to the frequency of incident radiation; moreover, it is independent of the intensity of light used.

(iii) There is no time lag between incidence of light and emission of photoelectrons.

(iv) For emission of photoelectrons, the frequency of incident light must be equal to or greater than the threshold frequency.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

28. The maximum kinetic energy of photoelectrons ejected from a metal, when it is irradiated with radiation of frequency $2 \times 10^{14}\ s^{-1}$ is 6.63×10^{-20} J. The threshold frequency of the metal is: [PMT (Kerala) 2008]

(a) $2 \times 10^{14}\ s^{-1}$ (b) $3 \times 10^{14}\ s^{-1}$

(c) $2 \times 10^{-14}\ s^{-1}$ (d) $1 \times 10^{-14}\ s^{-1}$

(e) $1 \times 10^{14}\ s^{-1}$

[Ans. (e)]

[Hint : Absorbed energy = Threshold energy + Kinetic energy of photoelectrons

$$h\nu = h\nu_0 + \text{KE}$$
$$h\nu_0 = h\nu - \text{KE}$$
$$6.626 \times 10^{-34} \times \nu_0 = 6.626 \times 10^{-34} \times 2 \times 10^{14} - 6.63 \times 10^{-20}$$
$$\nu_0 = \frac{1.3252 \times 10^{-19} - 6.63 \times 10^{-20}}{6.626 \times 10^{-34}}$$
$$\nu_0 = 9.99 \times 10^{13} = 10^{14}\ s^{-1}\ ]$$

29. If λ_0 and λ be the threshold wavelength and the wavelength of incident light, the velocity of photoelectrons ejected will be:

(a) $\sqrt{\frac{2h}{m}(\lambda_0 - \lambda)}$ (b) $\sqrt{\frac{2hc}{m}(\lambda_0 - \lambda)}$

(c) $\sqrt{\frac{2hc}{m}\left(\frac{\lambda_0 - \lambda}{\lambda\lambda_0}\right)}$ (d) $\sqrt{\frac{2h}{m}\left(\frac{1}{\lambda_0} - \frac{1}{\lambda}\right)}$

[Ans. (c)]

[Hint : Absorbed energy = Threshold energy + Kinetic energy of photoelectrons

$$\frac{hc}{\lambda} = \frac{hc}{\lambda_0} + \frac{1}{2}mv^2$$

$$v = \sqrt{\frac{2hc}{m}\frac{(\lambda_0 - \lambda)}{\lambda\lambda_0}}]$$

30. A radiation of wavelength λ illuminates a metal and ejects photoelectrons of maximum kinetic energy of 1eV. Another radiation of wavelength $\frac{\lambda}{3}$, ejects photoelectrons of maximum kinetic energy of 4 eV. What will be the work function of metal?
(a) 1eV (b) 2eV (3) 0.5eV (d) 3eV
[Ans. (c)]
[Hint : Absorbed energy = Threshold energy + Kinetic energy of photoelectrons

$$h\frac{c}{\lambda} = E_0 + 1\text{ eV} \quad ...(i)$$

$$3h\frac{c}{\lambda} = E_0 + 4\text{eV} \quad ...(ii)$$

$$3(E_0 + 1\text{ eV}) = E_0 + 4\text{eV}$$

$$E_0 = 0.5\text{eV}]$$

31. The ratio of slopes of maximum kinetic energy *versus* frequency and stopping potential (V_0) *versus* frequency, in photoelectric effect gives:
(a) charge of electron (b) planck's constant
(c) work function (d) threshold frequency
[Ans. (a)]
[Hint :

$$h\nu = h\nu_0 + eV_0$$

$$eV_0 = h\nu - h\nu_0$$

$$V_0 = \frac{h}{e}\nu - \frac{h}{e}\nu_0 \quad ...(i)$$

$$(\text{Slope})_1 = h/e$$

$$(\text{KE})_{max} = h\nu - h\nu_0 \quad ...(ii)$$

$$(\text{Slope})_2 = h$$

$$(\text{Slope})_2/(\text{Slope})_1 = \frac{h}{h/e} = e]$$

32. Ground state energy of H-atom is $(-E_1)$, the velocity of photoelectrons emitted when photon of energy E_2 strikes stationary Li^{2+} ion in ground state will be:
(a) $v = \sqrt{\frac{2(E_2 - E_1)}{m}}$ (b) $v = \sqrt{\frac{2(E_2 + 9E_1)}{m}}$
(c) $v = \sqrt{\frac{2(E_2 - 9E_1)}{m}}$ (d) $v = \sqrt{\frac{2(E_2 - 3E_1)}{m}}$
[Ans. (c)]
[Hint: Threshold energy of $Li^{2+} = 9E_1$
Absorbed energy = Threshold energy + Kinetic energy of photoelectrons

$$E_2 = 9E_1 + \frac{1}{2}mv^2$$

$$mv^2 = 2(E_2 - 9E_1)$$

$$v = \sqrt{\frac{2(E_2 - 9E_1)}{m}}]$$

2.21 SOME OTHER FUNDAMENTAL PARTICLES

Besides protons, neutrons and electrons, many more elementary particles have been discovered. These particles are also called **Fundamental particles.** Some of these particles are stable while the others are unstable. Out of stable particles, the electron, the proton, the antiproton and the positron are four mass particles while neutrino, photon and graviton are three energy particles. Among these, unstable particles are neutron, meson and ν-particles. The main characteristics of the particles are given in table 2.1 below.

Table 2.1

Particle	Symbol	Nature	Charge esu $\times 10^{-10}$	Mass (amu)	Discovered by
Positron	$e^+, 1e^0, \beta^+$	+	+ 4.8029	0.0005486	Anderson (1932)
Neutrino	ν	0	0	< 0.00002	Pauli
Antiproton	p^-	–	– 4.8029	1.00787	Chamberlain Sugri and Weighland (1955)
Photon	$h\nu$	0	0	0	Planck
Graviton	G	0	0	0	
Positive mu meson	μ^+	+	+ 4.8029	0.1152	Yukawa (1935)
Negative mu meson	μ^-	–	– 4.8029	0.1152	Anderson (1937)
Positive pi meson	π^+	+	+ 4.8029	0.1514	Powell (1947)
Negative pi meson	π^-	–	– 4.8029	0.1514	
Neutral pi meson	π^0	0	0	0.1454	

2.22 ISOTOPES

Isotopes are the atoms of the same element having different atomic masses (see determination of isotopic mass). The term 'isotope' was introduced by **Soddy**. This is a Greek word meaning same position (*Isos* = same, *topes* = position), since all the isotopes of an element occupy the same position in the periodic table. Isotopes of an element possess identical chemical properties but differ slightly in physical properties which depend on atomic mass. Isotopes were first identified in radioactive elements by **Soddy**. In 1919, Thomson established the existence of isotopes in a non-radioactive element, neon. Until now, more than 1000 isotopes have been identified (natural as well as artificial). Out of these about 320 occur in nature, approximately 280 of these are stable and the remaining 40 are radioactive.

Conclusions

(i) Number of neutrons present in the nuclei of various isotopes of an element is always different. The number of neutrons is determined by applying the formula $N = A - Z$ where A is mass number and Z is atomic number.
Hydrogen has three isotopes, ${}_1^1H$, ${}_1^2H$ and ${}_1^3H$.

	A (Mass number)	**Z**	**No. of neutrons**
${}_1^1H$	1	1	0
${}_1^2H$	2	1	1
${}_1^3H$	3	1	2

Oxygen has three isotopes, ${}^{16}O$, ${}^{17}O$ and ${}^{18}O$.

	A	**Z**	**No. of neutrons**
${}_8^{16}O$	16	8	8
${}_8^{17}O$	17	8	9
${}_8^{18}O$	18	8	10

(ii) In a neutral atom, the number of protons and the number of electrons are always the same, *i.e.*, the electronic configuration of all the isotopes of an element is the same. Thus, all the isotopes of an element show the same chemical properties. However, the rates of reactions may be different for different isotopes of an element.

(iii) All the isotopes of an element occupy the same position in the periodic table.

(iv) The isotopes of an element differ slightly in physical properties. The compounds formed by these isotopes will also have different physical properties.

Determination of Isotopic Mass

Chlorine has two isotopes ${}_{17}Cl^{35}$ and ${}_{17}Cl^{37}$; these are found in nature in 3 : 1 ratio or 75% : 25% respectively. Isotopic mass may be calculated as:

Isotopic mass of chlorine

$$= \frac{\%\text{ of } Cl^{35}}{100} \times \text{mass of } Cl^{35} + \frac{\%\text{ of } Cl^{37}}{100} \times \text{mass of } Cl^{37}$$

$$= \frac{75}{100} \times 35 + \frac{25}{100} \times 37 = 35.5$$

OR

Isotopic mass of chlorine

$$= \frac{\text{Ratio of } Cl^{35} \times \text{mass of } Cl^{35} + \text{Ratio of } Cl^{37} \times \text{mass of } Cl^{37}}{\text{Sum of ratio}}$$

$$= \frac{3 \times 35 + 1 \times 37}{4} = 35.5$$

2.23 THEORIES OF NUCLEAR STABILITY

Since, a nucleus contains positively charged protons, there must exist a strong repulsive force between them. It has been calculated that there exists an electrostatic repulsion of approximately six tons between two protons situated at a nuclear distance but at the same time the forces which bind the nucleus are very high. It has been found that nuclear forces attracting the same two particles (*i.e.*, protons) are at least forty times greater than the repulsive forces. Thus, two major forces exist in the nucleus. These are electrostatic and nuclear. The nuclear forces are stronger and the range of these forces is extremely small. The forces which operate between nucleons are referred to as exchange forces. In order to account for the stability of the nucleus, a theory known as **meson theory** was put forward by **Yukawa,** in 1935. Yukawa pointed out that neutrons and protons are held together by very rapid exchange of nuclear particles called **pi mesons.** These mesons may be electrically neutral, positive or negative (designated as π^0, π^+ and π^-) and possess a mass 275 times the mass of an electron. Nuclear forces arise from a constant exchange of mesons between nucleons with very high velocity (practically the velocity of light).

Let a neutron be converted into a proton by the emission of a negative meson. The emitted meson is accepted by another proton and converted into a neutron.

$$n_A \rightarrow p_A^+ + \pi^-$$
$$\pi^- + p_B^+ \rightarrow n_B$$

Similarly, a proton after emitting a positive meson is converted into a neutron and *vice-versa.*

$$p_A^+ \rightarrow n_A + \pi^+$$
$$\pi^+ + n_B \rightarrow p_B^+$$

or simply $\quad p \rightleftharpoons \pi^+ + n$

$$n \rightleftharpoons \pi^- + p$$

There may be two more types of exchange, *i.e.*, between neutron-neutron and proton-proton, involving neutral pi mesons.

$$\begin{matrix} p \searrow \\ p \swarrow \end{matrix} \pi^0 \qquad \begin{matrix} n \searrow \\ n \swarrow \end{matrix} \pi^0 \qquad \text{or simply} \quad p \underset{\pi^0}{\overset{\pi^0}{\rightleftharpoons}} p \quad \text{and} \quad n \underset{\pi^0}{\overset{\pi^0}{\rightleftharpoons}} n$$

Mass Defect—Binding Energy

It is observed that the atomic mass of all nuclei (except hydrogen) is different from the sum of the masses of protons and neutrons. For example, the helium nucleus consists of 2 protons and 2 neutrons. The combined mass of 2 protons and 2 neutrons should be

$$= 2 \times 1.00758 + 2 \times 1.00893$$
$$= 4.03302 \text{ amu}$$

The actual observed mass of helium nuclei is 4.0028 amu. A difference of 0.0302 amu is observed between these two values. This difference is termed as **mass defect.**

Mass defect = Total mass of nucleons − Observed atomic mass

This decrease in mass (*i.e.*, mass defect) is converted into energy according to Einstein equation $E = mc^2$. The energy released when a nucleus is formed from protons and neutrons is called the **binding energy.** This is the force which holds all the nucleons together in the nucleus. Binding energy can be defined in other ways also, *i.e.*, the energy required to break the nucleus into constituent protons and neutrons. Binding energy is measured in MeV (Million Electron-Volts), *i.e.*, 1 amu = 931 MeV.

Binding energy = Mass defect × 931 MeV

Binding energy can also be calculated in erg. This is

$= \text{Mass defect (amu)} \times 1.66 \times 10^{-24} \times (3 \times 10^{10})^2$ erg

$(1\,\text{MeV} = 1.60 \times 10^6 \text{ erg})$

The binding energy increases with the increase in atomic number of the element. This indicates that heavier nuclei should be more stable than lighter nuclei. But, it is not so because heavier nuclei above atomic number 82 are unstable. It is thus clear that total binding energy of a nucleus does not explain the stability of the nucleus.

The total binding energy of a nucleus when divided by the number of nucleons gives the average or mean binding energy per nucleon. The binding energy per nucleon is actually the measure of the stability of the nucleus. The greater the binding energy per nucleon, more stable is the nucleus.

$$\text{Binding energy per nucleon} = \frac{\text{Total binding energy}}{\text{Total number of nucleons}}$$

When binding energy per nucleon of a number of nuclei is plotted against the corresponding mass number, a graph is obtained (Fig. 2.19) whose characteristics are as follows:

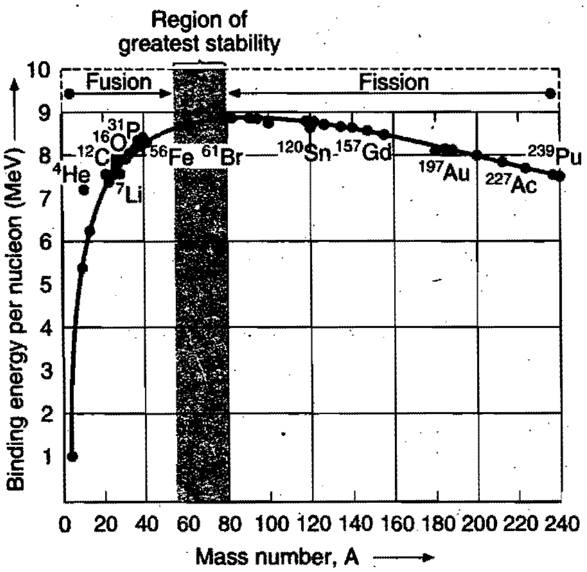


Fig. 2.19

(i) Binding energy per nucleon increases from 1.1 to 8.0 MeV from mass number 2 to 20.

(ii) Binding energy per nucleon increases from 8 to 8.6 MeV from mass number 20 to 40.

(iii) Binding energy per nucleon remains 8.6 – 8.7 MeV from mass number 40 to 90. Iron (56) has the maximum value of 8.7 MeV per nucleon.

(iv) The value of binding energy per nucleon decreases from 8.6 to 7.5 MeV from mass number 90 to 240.

(v) Points for helium, carbon, oxygen lie quite high in the graph showing that these nuclei are highly stable.

The binding energy per nucleon can be increased in two ways:

(i) Either by breaking heavy nucleus to those of intermediate mass numbers (process of fission) or

(ii) By fusing lighter nuclei to form heavier nuclei (process of fusion).

2.24 THE WHOLE NUMBER RULE AND PACKING FRACTION

Aston believed that mass number values (sum of protons and neutrons) of isotopes should be whole numbers on the scale of oxygen ($^{16}O = 16$) but actually it was observed that these were not integers. The difference in the atomic mass of an isotope and mass number was expressed by **Aston** (1927) as **packing fraction** by the following expression:

$$\text{Packing fraction} = \frac{\text{Isotopic atomic mass} - \text{Mass number}}{\text{Mass number}} \times 10^4$$

Thus, the packing fraction of $^1H = \frac{1.0078 - 1}{1} \times 10^4 = 78$ and the packing fraction of $^{35}Cl = \frac{34.980 - 35.0}{35.0} \times 10^4 = -5.7$. The packing fraction of oxygen is zero.

It is clear that the value of packing fraction varies from one atom to other. This is sometime positive or zero but more often negative.

A negative packing fraction means that atomic mass is less than nearest whole number and this suggests that some mass has been converted into energy when the particular isotope has been constituted. This energy is responsible for nuclear stability. All those having negative values of packing fraction are stable nuclei.

A positive packing fraction generally indicates instability of the nucleus. However, this statement is not correct for lighter nuclei.

In general, lower the value of packing fraction, the greater is the stability of the nucleus. The lowest values of packing fractions are observed for transition elements or iron family indicating thereby maximum stability of their nuclei.

2.25 THE MAGIC NUMBERS

It has been observed that atoms with an even number of nucleons in their nuclei are more plentiful than those with odd number. This indicates that a nucleus made up of even number of nucleons is more stable than a nuclei which consists of odd number of nucleons. It has also been observed that a stable nuclei results when either the number of neutrons or that of protons is equal to one of the numbers 2, 8, 20, 50, 82, 126. These numbers are called **magic numbers.** It is thought that the magic numbers form closed nuclear shells in the same way as the atomic numbers of inert gases form stable electronic configuration. In general, elements that have nuclei with magic number of protons as well as magic number of neutrons such as ^4_2He, $^{16}_8O$, $^{40}_{20}Ca$, $^{208}_{82}Pb$ are highly stable and found in abundance in nature.

A survey of stable nuclei found in nature shows the following trend:

Protons	Even	Even	Odd	Odd
Neutrons	Even	Odd	Even	Odd
No. of stable nuclei	157	52	50	5

Only five stable odd-odd nuclides are known; these nuclides are ^2_1H, ^6_3Li, $^{10}_5B$, $^{14}_7N$ and $^{180}_{73}Ta$.

SOME SOLVED EXAMPLES

Example 46. *The minimum energy required to overcome the attractive forces between an electron and the surface of Ag metal is 5.52×10^{-19} J. What will be the maximum kinetic energy of electrons ejected out from Ag which is being exposed to UV light of $\lambda = 360$ Å?*

Solution: Energy of the photon absorbed

$$= \frac{h \cdot c}{\lambda} = \frac{6.625 \times 10^{-27} \times 3 \times 10^{10}}{360 \times 10^{-8}}$$

$$= 5.52 \times 10^{-11} \text{ erg}$$

$$= 5.52 \times 10^{-18} \text{ J}$$

$$E(\text{photon}) = \text{work function} + \text{KE}$$

$$\text{KE} = 5.52 \times 10^{-18} - 7.52 \times 10^{-19}$$

$$= 47.68 \times 10^{-19} \text{ J}$$

Example 47. *Let a light of wavelength λ and intensity 'I' strikes a metal surface to emit x electrons per second. Average energy of each electron is 'y' unit. What will happen to 'x' and 'y' when (a) λ is halved (b) intensity I is doubled?*

Solution: (a) Rate of emission of electron is independent of wavelength. Hence, 'x' will be unaffected.

Kinetic energy of photoelectron = Absorbed energy − Threshold energy

$$y = \frac{hc}{\lambda} - w_0$$

when, λ is halved, average energy will increase but it will not become double.

(b) Rate of emission of electron per second 'x' will become double when intensity I is doubled. Average energy of ejected electron, *i.e.*, 'y' will be unaffected by increase in the intensity of light.

Example 48. *How many orbits, orbitals and electrons are there in an atom having atomic mass 24 and atomic number 12?*

Solution:

Atomic number = No. of protons = No. of electrons = 12

Electronic configuration = 2, 8, 2

No. of orbits = (K, L and M)

No. of orbitals on which electrons are present

= (one 1s + one 2s + three 2p + one 3s)

Example 49. *A neutral atom has 2K electrons, 8L electrons and 6 M electrons. Predict from this:*

(a) its atomic number, (b) total number of s-electrons, (c) total number of p-electrons, (d) total number of d-electrons.

Solution: (a) Total number of electrons

$$= (2 + 8 + 6) = 16$$

So, Atomic number = 16

Electronic configuration = $1s^2, 2s^2 2p^6, 3s^2 3p^4$

(b) Total number of s-electrons $= (1s^2 + 2s^2 + 3s^2) = 6$

(c) Total number of p-electrons $= (2p^6 + 3p^4) = 10$

(d) Total number of d-electrons = 0

Example 50. *Write down the values of quantum numbers of all the electrons present in the outermost orbit of argon (At. No. 18).*

Solution: The electronic configuration of argon is

$$1s^2, 2s^2 2p^6, 3s^2 3p_x^2 3p_y^2 3p_z^2$$

Values of quantum numbers are:

	n	l	m	s
$3s^2$	3	0	0	$+\frac{1}{2}, -\frac{1}{2}$
$3p_x^2$	3	1	± 1	$+\frac{1}{2}, -\frac{1}{2}$
$3p_y^2$	3	1	± 1	$+\frac{1}{2}, -\frac{1}{2}$
$3p_z^2$	3	1	0	$+\frac{1}{2}, -\frac{1}{2}$

Example 51. *(a) An electron is in 5f-orbital. What possible values of quantum numbers n, l, m and s can it have?*

(b) What designation is given to an orbital having

(i) $n = 2, l = 1$ and (ii) $n = 3, l = 0$?

Solution: (a) For an electron in $5f$-orbital, quantum numbers are:

$$n = 5; \quad l = 3; \quad m = -3, -2, -1, 0, +1, +2, +3$$

and $s = \text{either } +\frac{1}{2} \text{ or } -\frac{1}{2}$

(b) (i) $2p$, (ii) $3s$

Example 52. *Atomic number of sodium is 11. Write down the four quantum numbers of the electron having highest energy.*

Solution: The electronic configuration of sodium is:

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

$3s$-electron has the highest energy. Its quantum numbers are:

$$n = 3, l = 0, m = 0, s = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

Example 53. *An element has 8 electrons in 4d-subshell. Show the distribution of 8 electrons in the d-orbitals of the element within small rectangles.*

Solution: $4d$-subshell has five d-orbitals. These are first occupied singly and then pairing occurs. The distribution can be shown in the following manner:

4d

↑↓	↑↓	↑↓	↑	↑

Example 54. *How many elements would be in the third period of the periodic table if the spin quantum number m_s could have the value $-\frac{1}{2}$, 0 and $+\frac{1}{2}$?*

Solution:

$n = 3, l = 0, m = 0 \qquad m_s = -\frac{1}{2}, 0, +\frac{1}{2}$

$l = 1; m = -1, 0, +1 \quad \begin{cases} m_s = -\frac{1}{2}, 0, +\frac{1}{2} \\ m_s = -\frac{1}{2}, 0, +\frac{1}{2} \\ m_s = -\frac{1}{2}, 0, +\frac{1}{2} \end{cases}$

$l = 2; \quad m = -2, -1, 0, +1, +2 \left\{ \begin{array}{c} m_s = -\frac{1}{2}, 0, +\frac{1}{2} \\ \text{for each value of} \\ \text{magnetic} \\ \text{quantum no.} \end{array} \right\}$

Number of elements = $3s\ (3e)$
$3p\ (9e)$
$3d\ (15e)$

$\therefore$ 27 elements will be there in third period of periodic table.

Example 55. *The binding energy of ^4_2He is 28.57 MeV. What shall be the binding energy per nucleon of this element?*

Solution: The nucleus of $^4_2$He consists of 4 nucleons.

$$\text{So, Binding energy per nucleon} = \frac{\text{Total binding energy}}{\text{No. of nucleons}}$$

$$= \frac{28.57}{4} = 7.14 \text{ MeV}$$

Example 56. *Calculate the binding energy of the oxygen isotope $^{16}_8O$. The mass of the isotope is 16.0 amu. (Given $e = 0.0005486$ amu, $p = 1.00757$ amu and $n = 1.00893$ amu.)*

Solution: The isotope $^{16}_8$O contains 8 protons, 8 neutrons and 8 electrons.

Actual mass of the nucleus of $^{16}_8$O

$$= 16 - \text{mass of 8 electrons}$$

$$= 16 - 8 \times 0.0005486 = 15.9956 \text{ amu}$$

Mass of the nucleus of $^{16}_8$O

$$= \text{mass of 8 protons} + \text{mass of 8 neutrons}$$

$$= 8 \times 1.00757 + 8 \times 1.00893 = 16.132 \text{ amu}$$

$$\text{Mass defect} = (16.132 - 15.9956) = 0.1364 \text{ amu}$$

$$\text{Binding energy} = 0.1364 \times 931 = 127 \text{ MeV}$$

Example 57. *There are four atoms which have mass numbers 9, 10, 11 and 12 respectively. Their binding energies are 54, 70, 66 and 78 MeV respectively. Which one of the atoms is most stable?*

Solution: Stability depends on the value of binding energy per nucleon.

	A	*B*	*C*	*D*
Binding energy (MeV)	54	70	66	78
No. of nucleons	9	10	11	12
Binding energy per nucleon (MeV)	6	7	6	6.5

Thus, *B* is most stable.

MISCELLANEOUS NUMERICAL EXAMPLES

These examples will give the sharp edge to the aspirants for IIT and various other entrance examinations.

Example 1. *The Schrödinger wave equation for hydrogen atom is*

$$\psi_{2s} = \frac{1}{4\sqrt{2\pi}}\left(\frac{1}{a_0}\right)^{3/2}\left[2 - \frac{r_0}{a_0}\right]e^{-r/a_0}$$

where a_0 is Bohr radius. If the radial node in 2s be at r_0, then find r in terms of a_0. **(IIT 2004)**

Solution: Given,

$$\psi_{2s} = \frac{1}{4\sqrt{2\pi}}\left(\frac{1}{a_0}\right)^{3/2}\left[2 - \frac{r_0}{a_0}\right]e^{-r/a_0}$$

$$\psi^2_{2s} = 0 \text{ at node}$$

$$\therefore \quad 2 - \frac{r_0}{a_0} = 0$$

$$r_0 = 2a_0$$

Example 2. *Consider the hydrogen atom to be a proton embedded in a cavity of radius a_0 (Bohr radius) whose charge is neutralized by the addition of an electron to the cavity in vacuum infinitely slowly. Estimate the average total energy of an electron in its ground state in a hydrogen atom as the work done in the above neutralization process. Also, if the magnitude of average KE is half the magnitude of average potential energy, find the average potential energy.* **(IIT 1996)**

Solution:
Coulombic force of attraction = Centrifugal force

$$\frac{1}{4\pi\varepsilon_0}\frac{Ze \times e}{a_0^2} = \frac{mv^2}{a_0}$$

where, v = velocity of electron

a_0 = distance between electron and nucleus

$$\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{a_0} = mv^2$$

$$\text{KE} = \frac{1}{2}mv^2 = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{2a_0}$$

$$\text{PE} = -2 \times \text{KE}$$

$$= -2 \times \frac{1}{4\pi\varepsilon_0} \times \frac{Ze^2}{2a_0} = -\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{a_0}$$

Example 3. *Hydrogen atoms are excited from ground state. Its spectrum contains wavelength 486 nm. Find, what transition does the line corresponds to. Also find from this information what other wavelengths will be present in the spectrum?*

Solution: Wavelength 486 nm, *i.e.*, 4860 Å indicates that the spectrum is in visible region, *i.e.*, Balmer series.

$$\frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\frac{1}{4860\times10^{-8}} = 109677.76\times1^2\left[\frac{1}{2^2}-\frac{1}{n_2^2}\right]$$

On solving, we get

$$n_2^2 = 16$$

$$n_2 = 4$$

Thus, transition is from, $4 \rightarrow 2$

Other transitions in the spectrum are

$$4 \rightarrow 3 \rightarrow 2$$

$$\frac{1}{\lambda} = 109677.76\times1^2\times\left[\frac{1}{3^2}-\frac{1}{4^2}\right]$$

$$\lambda = 1875\times10^{-7}\text{ cm}$$

Example 4. *If uncertainties in the measurement of position and momentum of an electron are equal, calculate uncertainty in the measurement of velocity.*

Solution: According to Heisenberg's uncertainty principle,

$$\Delta x.\Delta p \geq \frac{h}{4\pi}$$

Given, $\Delta x = \Delta p = \sqrt{\frac{h}{4\pi}} = 0.726\times10^{-17}$

$$\Delta p = m\Delta V$$

or $$\Delta V = \frac{\Delta p}{m} = \frac{0.726\times10^{-17}}{9.1\times10^{-31}} = 7.98\times10^{12}\text{ ms}^{-1}$$

Example 5. *How much energy will be released when a sodium ion and a chloride ion, originally at infinite distance are brought together to a distance of 2.76 Å (the shortest distance of approach in a sodium chloride crystal)? Assume that ions act as point charges, each with a magnitude of 1.6×10^{-19} C. Permittivity constant of the medium is 9×10^9 Nm^2C^{-2}.*

Solution: Energy released

$$= -K\frac{q_1 q_2}{r} = -\frac{9\times10^9\times(1.6\times10^{-19})^2}{2.76\times10^{-10}} = -8.35\times10^{-19}\text{ J}$$

Example 6. *The angular momentum of an electron in a Bohr orbit of H-atom is 4.2178×10^{-34} kg m²/sec. Calculate the spectral line emitted when an electron falls from this level to the next lower level.*

Solution: We know, $mvr = n\frac{h}{2\pi}$

$$4.2178\times10^{-34} = n\times\frac{6.626\times10^{-34}}{2\times3.14}$$

$$\therefore \quad n = 4$$

$$\frac{1}{\lambda} = R_H\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$$

$$= 109678\left[\frac{1}{3^2}-\frac{1}{4^2}\right]$$

$$\lambda = 1.8\times10^{-4}\text{ cm}$$

Example 7. *A negatively charged particle called Negatron was discovered. In the Millikan's oil-drop experiment, the charges of the oil-drops in five experiments are reported as 3.2×10^{-19} coulomb; 4.8×10^{-19} coulomb; 6.4×10^{-19} coulomb; 8×10^{-19} coulomb and 9.6×10^{-19} coulomb. Calculate the charge on the negatron.*

Solution: In Millikan's oil-drop experiment; the charges on the oil-drops are integral multiples of the charge of the particle. Dividing the charges of droplets by the lowest charge:

(i) $\frac{3.2\times10^{-19}}{3.2\times10^{-19}} = 1$ (ii) $\frac{4.8\times10^{-19}}{3.2\times10^{-19}} = 1.5$

(iii) $\frac{6.4\times10^{-19}}{3.2\times10^{-19}} = 2$ (iv) $\frac{8\times10^{-19}}{3.2\times10^{-19}} = 2.5$

(v) $\frac{9.6\times10^{-19}}{3.2\times10^{-19}} = 3$

All the values are not integral; they can be converted to integers on multiplying by 2.

$\therefore$ Charge of the negatron will be

$$\frac{3.2\times10^{-19}}{2} = 1.6\times10^{-19}\text{ C}$$

Example 8. *When a certain metal was irradiated with light of frequency 3.2×10^{16} Hz, the photoelectrons emitted had twice the kinetic energy as did photoelectrons emitted when the same metal was irradiated with light of frequency 2.0×10^{16} Hz. Calculate ν_0 for the metal.*

Solution: Applying photoelectric equation,

$$\text{KE} = h\nu - h\nu_0$$

or $$(\nu - \nu_0) = \frac{\text{KE}}{h}$$

Given, $$\text{KE}_2 = 2\text{KE}_1$$

$$\nu_2 - \nu_0 = \frac{\text{KE}_2}{h} \quad \ldots\text{(i)}$$

and $$\nu_1 - \nu_0 = \frac{\text{KE}_1}{h} \quad \ldots\text{(ii)}$$

Dividing equation (i) by equation (ii),

$$\frac{\nu_2-\nu_0}{\nu_1-\nu_0} = \frac{\text{KE}_2}{\text{KE}_1} = \frac{2\text{KE}_1}{\text{KE}_1} = 2$$

or $$\nu_2 - \nu_0 = 2\nu_1 - 2\nu_0$$

or $$\nu_0 = 2\nu_1 - \nu_2 = 2(2.0\times10^{16}) - (3.2\times10^{16})$$

$$= 8.0\times10^{15}\text{ Hz}$$

Example 9. *An electron moves in an electric field with a kinetic energy of 2.5 eV. What is the associated de Broglie wavelength?*

Solution: Kinetic energy

$$= \frac{1}{2}mv^2 \left(v = \frac{h}{m\lambda}\right)$$

$$= \frac{1}{2}m\left(\frac{h}{m\lambda}\right)^2$$

$$= \frac{1}{2}\frac{h^2}{m\lambda^2}$$

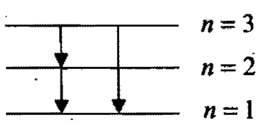
or $$\lambda^2 = \frac{1}{2}\frac{h^2}{m \times \text{KE}}$$

$$\lambda = \frac{h}{\sqrt{2m \times \text{KE}}} \quad \begin{pmatrix} m = 9.108 \times 10^{-28} \text{ g} \\ h = 6.626 \times 10^{-27} \text{ erg-sec} \\ 1\text{eV} = 1.602 \times 10^{-12} \text{ erg} \end{pmatrix}$$

$$= \frac{6.626 \times 10^{-27}}{\sqrt{2 \times 9.108 \times 10^{-28} \times 2.5 \times 1.602 \times 10^{-12}}}$$

$$= 7.7 \times 10^{-8} \text{ cm}$$

Example 10. *Consider the following two electronic transition possibilities in a hydrogen atom as pictured below:*



(*a*) *The electron drops from third Bohr orbit to second Bohr orbit followed with the next transition from second to first Bohr orbit.*

(*b*) *The electron drops from third Bohr orbit to first Bohr orbit directly. Show that the sum of energies for the transitions $n = 3$ to $n = 2$ and $n = 2$ to $n = 1$ is equal to the energy of transition for $n = 3$ to $n = 1$.*

Solution: Applying, $\Delta E = R_H\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

For $n = 3$ to $n = 2$;

$$\Delta E_{3\to 2} = R_H\left[\frac{1}{2^2} - \frac{1}{3^2}\right] = R_H \times \frac{5}{36} \quad \ldots \text{(i)}$$

For $n = 2$ to $n = 1$;

$$\Delta E_{2\to 1} = R_H\left[\frac{1}{1^2} - \frac{1}{2^2}\right] = R_H \times \frac{3}{4} \quad \ldots \text{(ii)}$$

For $n = 3$ to $n = 1$;

$$\Delta E_{3\to 1} = R_H\left[\frac{1}{1^2} - \frac{1}{3^2}\right] = R_H \times \frac{8}{9} \quad \ldots \text{(iii)}$$

Adding equations (i) and (ii),

$$R_H\left(\frac{5}{36} + \frac{3}{4}\right) = R_H\left(\frac{5+27}{36}\right) = R_H \times \frac{8}{9}$$

Thus, $\Delta E_{3\to 1} = \Delta E_{3\to 2} + \Delta E_{2\to 1}$

Example 11. *If an electron is moving with velocity $500\, ms^{-1}$, which is accurate up to 0.005% then calculate uncertainty in its position. [$h = 6.63 \times 10^{-34}$ Js, mass of electron $= 9.1 \times 10^{-31}$ kg]* **[AIPMT (Mains) 2008]**

Solution : Uncertainty in velocity

$$\Delta v = \frac{600 \times 0.005}{100} = 3 \times 10^{-2} \text{ms}^{-1}$$

According to Heisenberg's uncertainty principle

$$\Delta x\, \Delta v \geq \frac{h}{4\pi m}$$

$$\Delta x \geq \frac{h}{4\pi m \Delta v}$$

$$\geq \frac{6.63 \times 10^{-34}}{4 \times 3.14 \times 9.1 \times 10^{-31} \times 3 \times 10^{-2}}$$

$$= 1.9 \times 10^{-3} \text{m}$$

Example 12. *Applying Bohr's model when H atom comes from $n = 4$ to $n = 2$, calculate its wavelength. In this process, write whether energy is released or absorbed? Also write the range of radiation. $R_H = 2.18 \times 10^{-18}$ J, $h = 6.63 \times 10^{-34}$ Js.* **(AIPMT 2008)**

Solution : Energy is released in this process; and the radiation will belong to visible region (Balmer series)

$$E = \frac{hc}{\lambda} = R_H Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\frac{1}{\lambda} = \frac{R_H Z^2}{hc}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$= \frac{2.18 \times 10^{-18} \times 1^2}{6.63 \times 10^{-34} \times 3 \times 10^8}\left[\frac{1}{4} - \frac{1}{16}\right]$$

$$= \frac{2.18 \times 10^{-18} \times 1^2}{6.63 \times 10^{-34} \times 3 \times 10^8}\left[\frac{3}{16}\right]$$

$$\lambda = \frac{6.63 \times 10^{-34} \times 3 \times 10^8 \times 16}{3 \times 2.18 \times 10^{-18}} = 4866 \times 10^{-10} \text{m}$$

$$= 4866 \text{ Å}$$

SUMMARY AND IMPORTANT POINTS TO REMEMBER

1. Atom is the smallest indivisible particle of matter (proposed by John Dalton in 1808).

2. All atoms except hydrogen atom are composed of three fundamental particles, namely, electron, proton and neutron. Hydrogen atom has one electron and one proton but no neutron.

(a) Electron: The nature and existence of electron was established by experiments on conduction of electricity through gases, *i.e.*, discovery of cathode rays. In 1897, **J.J. Thomson** determined e/m value (-1.7588×10^8 coulomb/g) and proved that whatever gas be taken in the discharge tube and whatever be the material of the electrodes, the value of e/m is always the same. Electrons are, thus, common universal constituents of all atoms.

Electron is a subatomic particle which carries charge -1.60×10^{-19} coulomb, *i.e.*, one unit negative charge and has mass 9.1×10^{-28} g (or 9.1×10^{-31} kg), *i.e.*, $\frac{1}{1837}$th mass of hydrogen atom (0.000549 amu). The name electron was given by **Stoney.**

(b) Proton: The nature and existence of proton was established by the discovery of positive rays (**Goldstein** 1886). Proton is a subatomic particle which carries $+1.6 \times 10^{-19}$ coulomb or one unit positive charge and has mass 1.672×10^{-24} g (or 1.672×10^{-27} kg), *i.e.*, 1.0072 amu. The e/m was determined by **Thomson** in 1906 and the value is $+9.579 \times 10^4$ coulomb/g. It was named as proton by **Rutherford.**

(c) Neutron: It is a subatomic particle which carries no charge. Its mass is 1.675×10^{-24} g (1.675×10^{-27} kg) or 1.0086 amu. It is slightly heavier than proton. It was discovered by **Chadwick** in 1932 by bombarding beryllium with α-particles.

$${}^{9}_{4}\mathrm{Be} + {}^{4}_{2}\mathrm{He} \rightarrow {}^{12}_{6}\mathrm{C} + {}^{1}_{0}n$$

The e/m value of neutron is zero.

3. According to the Rutherford's model of atom, (i) it consists of **nucleus** of very small size and high density (ii) electrons revolve round the nucleus in a circular path.

$$\text{Radius of nucleus} = 10^{-15}\ \text{m}$$

$$\text{Density of nucleus} = 10^{8}\ \text{tonnes/cc}$$

4. Atomic number (Z) = Number of protons in the nucleus

5. Mass number (A) = Number of protons + Number of neutrons

6. **Isotopes:** These are atoms of same element having same atomic number but different mass numbers, *e.g.*,

$$({}^{1}_{1}\mathrm{H}, {}^{2}_{1}\mathrm{H}, {}^{3}_{1}\mathrm{H}) : ({}^{35}_{17}\mathrm{Cl}, {}^{37}_{17}\mathrm{Cl})$$

7. **Isobars:** These are atoms of different elements having same mass number but different atomic numbers, *e.g.*,

$${}^{40}_{18}\mathrm{Ar}, {}^{40}_{19}\mathrm{K}, {}^{40}_{20}\mathrm{Ca}$$

8. **Isotones:** These are atoms of different elements having same number of neutrons in the nucleus, *e.g.*,

$${}^{14}_{6}\mathrm{C} : {}^{15}_{7}\mathrm{N}, {}^{16}_{8}\mathrm{O}$$

9. Electromagnetic radiations are energy waves containing both electric and magnetic vector perpendicular to each other.

(i) These are transverse waves.

(ii) They do not need any medium for their propagation. They travel with the velocity of light.

(iii) $\nu = \frac{c}{\lambda}$, ν = frequency, c = velocity of light,

λ = wavelength

$\bar{\nu} = \frac{1}{\lambda}$ = wave number, $T = \frac{1}{\nu}$ = time period.

(iv) According to Planck's quantum theory, the energy is emitted or absorbed in the form of energy packets called **quanta.** Quantum of visible light is called **photon.**

Energy of one quantum $= h\nu$

$$= h\frac{c}{\lambda}$$

$$h = \text{Planck's constant}$$

$$= 6.626 \times 10^{-34}\ \text{J sec}$$

10. **Hydrogen spectrum:** Hydrogen spectrum is a line spectrum. The lines lie in visible, ultraviolet and infrared regions. All the lines can be classified into five series. Ritz presented a mathematical formula to find the wavelengths of various lines,

$$\frac{1}{\lambda} = \bar{\nu} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

where, R is Rydberg constant ($R = 109678\ \text{cm}^{-1}$).

		n_1	n_2
Lyman series	(UV region)	1	2, 3, 4, 5,...
Balmer series	(Visible region)	2	3, 4, 5, 6,...
Paschen series	(IR region)	3	4, 5, 6, 7,...
Brackett series	(IR region)	4	5, 6, 7, 8,...
Pfund series	(IR region)	5	6, 7, 8, 9,...

Balmer series consists of four prominent lines H_α, H_β, H_γ and H_δ having wavelength 6563 Å, 4861 Å, 4340 Å and 4102 Å respectively.

Balmer equation is,

$$\frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{n^2}\right]$$

where, $n = 3, 4, 5, 6, \ldots$

The Rydberg formula is used to calculate the wavelength of any line of the spectrum

$$\frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{(n_1+x)^2}\right]$$

where, x = number of lines in the spectrum; $x = \infty$ for series limit or last line. Let, transition of electrons takes place from n_2 to n_1 shell; then the number of lines can be calculated as:

$$\text{Number of lines} = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

11. Bohr's atomic model: It is based on Planck's quantum theory. Its main postulates are summarised as:

(i) Electrons revolve round the nucleus in circular path of fixed energy called **stationary states.**

(ii) Angular momentum of electrons are quantised, *i.e.*,

$$mvr = n\left(\frac{h}{2\pi}\right)$$

(iii) The energy as well as angular momentum both are quantised for electrons. It means they can have only certain values of energy and angular momenta.

12. Important formulations obtained from Bohr's atomic model which are valid for single electron species like H, He^+, Li^{2+}, Be^{3+}, etc.:

(i) $E_1 < E_2 < E_3 < E_4$

(ii) $(E_2 - E_1) > (E_3 - E_2) > (E_4 - E_3)\ldots$

where, $E_1, E_2, E_3, \ldots$ are energies of corresponding shells.

(iii) $$r_n = \frac{n^2h^2}{4\pi^2 Ke^2 mZ}$$

$$K = \frac{1}{4\pi\varepsilon_0} = 9 \times 10^9 \text{ Nm}^2/\text{C}^2$$

$r = \dfrac{n^2}{Z} \times 0.529$ Å (where, r is the radius of Bohr orbit of electrons.)

(iv) Energy of electrons in a particular shell can be calculated as:

$$E = -\frac{Z^2}{n^2}\frac{2\pi^2 mK^2e^4}{h^2}$$

$$E = -\frac{Z^2}{n^2} \times 21.79 \times 10^{-19} \text{ J/atom}$$

$$= -\frac{Z^2}{n^2} \times 13.6 \text{eV}$$

$$= -\frac{Z^2}{n^2} \times 1312 \text{ kJ/mol}$$

$$E_n = \frac{Z^2 R_E}{n^2}$$

$R_E = -13.6$ eV (Rydberg energy)

(v) $E_n = E_1/n^2$; $E_n = E_1 \times \dfrac{Z^2}{n^2}$ for hydrogen-like species.

(vi) Velocity of electrons in a particular shell or orbit can be calculated as:

$$v = \sqrt{\frac{Ke^2}{mr}}$$

where, $K = \dfrac{1}{4\pi\varepsilon_0} = 9 \times 10^9 \text{ Nm}^2/\text{C}^2$

$$v = \frac{Z}{n} \times 2.188 \times 10^8 \text{ cm/sec}$$

(vii) Potential energy of electrons in a particular shell:

$$\text{PE} = \frac{-KZe^2}{r} = -\frac{27.2}{n^2} \times Z^2 \text{ eV}$$

(viii) Kinetic energy of electrons in a particular shell:

$$\text{KE} = \frac{1}{2}\frac{KZe^2}{r} = +\frac{13.6}{n^2} Z^2 \text{ eV}$$

$$\text{Total Energy, TE} = -\frac{1}{2}\frac{KZe^2}{r}$$

$$\text{TE} = \frac{1}{2}\text{PE}$$

$$\text{TE} = -\text{KE}$$

(ix) Number of revolutions per second by an electron in a shell:

$$= \frac{\text{Velocity}}{\text{Circumference}} = \frac{v}{2\pi r} = -\frac{E_1}{h} \times \frac{2}{n^3}$$

(x) Frequency of electrons in nth orbit:

$$= \frac{v}{2\pi r}$$

$$= \frac{6.62 \times 10^{15} Z^2}{n^3}$$

(xi) Period of revolution of electrons in nth orbit (T_n),

$$T_n = \frac{2\pi r}{V_n} = \frac{1.5 \times 10^{-16} n^3}{Z^2} \text{ sec}$$

$$T \propto \frac{n^3}{Z^2}$$

(xii) Ionization energy $= E_\infty - E_n$

$$= 0 - \left(-\frac{Z^2}{n^2} \times 13.6 \text{eV}\right)$$

$$= \frac{Z^2}{n^2} \times 13.6 \text{eV}$$

$$= \frac{Z^2}{n^2} \times 21.79 \times 10^{-19} \text{ J/atom}$$

(xiii) $$\frac{I_1}{I_2} = \frac{Z_1^2}{Z_2^2} \times \frac{n_2^2}{n_1^2}$$

I_1 and I_2 are ionization energies of two elements 1 and 2.

(xiv) ΔE (Energy of transition) $= R_E\left(\dfrac{1}{n_1^2} - \dfrac{1}{n_2^2}\right)$

$R_E = -13.6$ eV

R_H = Rydberg constant $= \frac{R_E}{hc} = 109677\ cm^{-1}$

Defects of Bohr theory: (i) It fails to explain the spectra of multi-electron atoms. (ii) It fails to explain fine spectrum of hydrogen. (iii) It does not provide an explanation why angular momentum should always be an integral multiple of $h/2\pi$. (iv) It does not explain splitting of spectral lines under the influence of magnetic field (Zeeman effect) and electric field (Stark effect).

13. Sommerfeld's extension: Sommerfeld (1915) introduced the idea of elliptical orbits. Except first orbit which is only circular, the other orbits are elliptical. The second orbit has one elliptical and one circular suborbit. The third orbit has two elliptical and one circular suborbit.

14. Dual nature: Light has dual character, *i.e.*, it behaves sometimes like particles and sometimes like waves. de Broglie (1924) predicted that small particles such as electrons should show wave-like properties along with particle character. The wavelength (λ) associated with a particle of mass m and moving with velocity v is given by the relationship $\lambda = \frac{h}{mv}$; where, h is Planck's constant.

The wave nature was confirmed by Davisson and Germer's experiment.

Davisson and Germer gave some modified equations for calculation of de Broglie wavelength:

$\lambda = \frac{h}{\sqrt{2Em}}$; where, E = kinetic energy of the particle.

$\lambda = \frac{h}{\sqrt{2qVm}}$; where, q = charge of the particle accelerated by the potential of V volt.

15. Heisenberg uncertainty principle: It is impossible to measure simultaneously both the position and momentum of any microscopic particle with accuracy. Mathematically, $\Delta x\, \Delta p \approx \frac{h}{4\pi}$; where, Δx = uncertainty in position and Δp = uncertainty in momentum. It introduces the concept of probability of locating the electron in space around the nucleus.

16. de Broglie concept as well as uncertainty principle have no significance in everyday life because they are significant for only microscopic systems.

17. When radiations of a certain minimum frequency (ν_0), called threshold frequency, strike the surface of a metal, electrons called photoelectrons are ejected from the surface. The minimum energy required to eject the electrons from the metal surface is called threshold energy or work function.

Absorbed energy = Threshold energy + Kinetic energy of photoelectrons

$$E = E_0 + \text{KE}$$

$$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

$$\frac{hc}{\lambda} = \frac{hc}{\lambda_0} + \frac{1}{2}mv^2$$

ν_0 and λ_0 are called threshold frequency and threshold wavelength respectively.

18. Quantum numbers: The set of four integers required to define an electron completely in an atom are called quantum numbers. The first three have been derived from Schrödinger's wave equation.

(i) Principal quantum number: It describes the name, size and energy of the shell to which the electron belongs.

$n = 1, 2, 3, 4, \ldots$ represent K, L, M, N, ... shells respectively.

Formulae for radius, energy and angular momentum of electrons are given earlier.

(ii) Azimuthal quantum number: It is denoted by 'l'. It describes the shape of electron cloud and number of subshells in a shell.

$$l = 0, 1, 2, 3, \ldots, (n-1)$$

$l = 0$ (s-subshell); $l = 1$ (p-subshell); $l = 2$ (d-subshell); $l = 3$ (f-subshell).

Orbital angular momentum of electron

$$= \sqrt{l(l+1)}\,\frac{h}{2\pi} = \sqrt{l(l+1)}\,\hbar$$

when $l = 0$, electrons revolve in a circular orbit and when $l \neq 0$, the electrons revolve round the nucleus in an elliptical path.

(iii) Magnetic quantum number: It is denoted by 'm'. It describes the orientations of the subshells. It can have values from $-l$ to $+l$ including zero, *i.e.*, total $(2l+1)$ values. Each value corresponds to an orbital. s-subshell has one orbital, p-subshell has three orbitals (p_x, p_y and p_z), d-subshell has five orbitals ($d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ and d_{z^2}) and f-subshell has seven orbitals. One orbital can accommodate either one or two electrons but not more than two. s-orbital is spherically symmetrical and non-directional. p-orbitals have dumb-bell shape and are directional in nature. Four d-orbitals have double dumb-bell shape but d_{z^2} has a baby soother shape. The total number of orbitals present in a main energy level is 'n^2'.

(iv) Spin quantum number (s): It describes the spin of the electron. It has values +1/2 and −1/2. (+) signifies clockwise spinning and (−) signifies anticlockwise spinning.

$$\text{Spin angular momentum} = \sqrt{s(s+1)}\,\frac{h}{2\pi}$$

$$= \sqrt{s(s+1)}\,\hbar = \frac{\sqrt{3}}{2}\hbar \quad \left(\text{where, } s = \frac{1}{2}\right)$$

Total spin of an atom or an ion $= n \times \frac{1}{2}$; where, 'n' is the number of unpaired electrons.

Spin multiplicity of an atom $= (2\Sigma s + 1)$

Singlet state (Normal)

$\uparrow\downarrow$

Spin multiplicity

$= 2\Sigma s + 1$

$= 2 \times 0 + 1 = 1$

Singlet excited

$\downarrow$

$\uparrow$

Spin multiplicity = 1

Triplet excited state

$\uparrow$

$\uparrow$

Spin multiplicity

$= 2 \times \left(\frac{1}{2} + \frac{1}{2}\right) + 1 = 3$

19. (i) Number of subshells in a shell = n

(ii) Number of maximum orbitals in a shell = n^2

(iii) Number of maximum orbitals in a subshell = $2l + 1$

(iv) Maximum number of electrons in a shell = $2n^2$

(v) Maximum number of electrons in a subshell

$= 2(2l + 1)$

(vi) Z-component of the angular momentum depends upon magnetic quantum number and is given as:

$$L_Z = m\left(\frac{h}{2\pi}\right)$$

(vii) Number of radial/spherical nodes in any orbital

$= (n - l - 1)$

1s orbital has no node; 2s orbital has one spherical node; 2p orbital has no spherical node; 3p orbital has one spherical node.

(viii) Schrödinger wave equation does not give spin quantum number.

(ix) A plane passing through the nucleus at which the probability of finding the electron is zero, is called nodal plane.

The number of nodal plane in an orbital = l

s-orbitals have no nodal plane; p-orbitals have one nodal plane, d-orbitals have two nodal planes and so on.

20. Pauli's exclusion principle: No two electrons in an atom can have the same set of all the four quantum numbers, *i.e.*, an orbital cannot have more than 2 electrons because three quantum numbers (principal, azimuthal and magnetic) at the most may be same but the fourth must be different, *i.e.*, spins must be in opposite directions. It is possible to calculate the maximum number of electrons which can be accommodated on a main energy shell or subenergy shell on the basis of this principle.

21. Electronic configuration: The arrangement of electrons in various shells, subshells and orbitals in an atom is termed electronic configuration. It is written in terms of nl^x where n indicates the order of shell, l indicates the subshell and x the number of electrons present in the subshell.

22. Aufbau principle: Aufbau is a German word meaning building up. The electrons are filled in various orbitals in an order of their increasing energies. An orbital of lowest energy is filled first. The sequence of orbitals in the order of their increasing energy is:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, ...

The energy of the orbitals is governed by $(n + l)$ rule.

(i) Subshell with lower of $(n + l)$ has lower energy, hence filled first, *e.g.*,

$3p(n + l = 4)$ will be filled before $3d(n + l = 5)$.

(ii) When $(n + l)$ values are same, then the subshell with lower value of 'n' is filled first, *e.g.*,

$3p(n + l = 4)$ will be filled before $4s\,(n + l = 4)$ because 3p has lower value of n.

23. Hund's rule: No electron pairing takes place in the orbitals in a subenergy shell until each orbital is occupied by one electron with parallel spin. Exactly half-filled and fully-filled orbitals make the atoms more stable, *i.e.*, $p^3, p^6, d^5, d^{10}, f^7$ and f^{14} configurations are most stable.

All those atoms which consist of at least one orbital singly occupied behave as **paramagnetic** while all those atoms in which all the orbitals are doubly occupied are **diamagnetic** in nature.

Magnetic moment = $\sqrt{n(n+2)}$ BM

n = number of unpaired electrons

24. Half-filled and fully-filled subshells have extra stability due to **greater exchange energy** and **spherical symmetry** around the nucleus.

25. It is only dz^2 orbitals which do not have four lobes like other d-orbitals.

26. The d-orbital whose lobes lie along the axes is $d_{x^2-y^2}$.

27. Wave mechanical model of atom: It was Schrödinger who developed a new model known as wave mechanical model of atom by incorporating the conclusions of de Broglie and Heisenberg uncertainty principles. He derived an equation, known as Schrödinger equation.

$$\frac{d^2\psi}{dx^2} + \frac{d^2\psi}{dy^2} + \frac{d^2\psi}{dz^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$

The solution of the equation provides data which enables us to calculate the probability of finding an electron of specific energy. It is possible to determine the regions of space around the nucleus where there is maximum probability of locating an electron of specific energy. This region of space is termed orbital.

ψ is the amplitude of the wave at a point with coordinates x, y and z. 'E' is total energy called eigen value and V denotes the potential energy of the electron.

ψ^2 gives the probability of finding the electron at (x, y and z).

Operator form of the equation can be given as:

$$\hat{H}\psi = E\psi$$

$$\hat{H} = \left[-\frac{h^2}{8\pi^2 m} \Delta^2 + \hat{V} \right] \text{called Hamiltonian operator}$$

$$= \hat{T} + \hat{V}$$

$\hat{T}$ = Kinetic energy operator

$\hat{V}$ = Potential energy operator

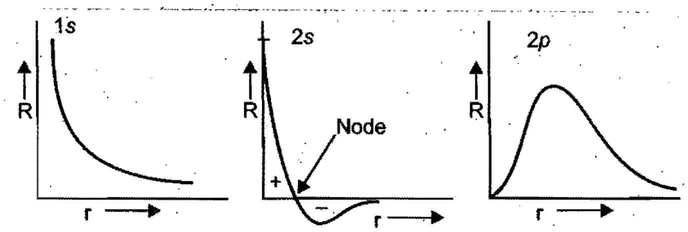
28. Complete wave function can be given as:

$$\psi(r,\theta,\phi) = \underbrace{R(r)}_{\text{Radial part}} \; ; \underbrace{\Theta(\theta)\,\Phi(\phi)}_{\text{Angular part}}$$

Dependence of the wave function on quantum number can be given as:

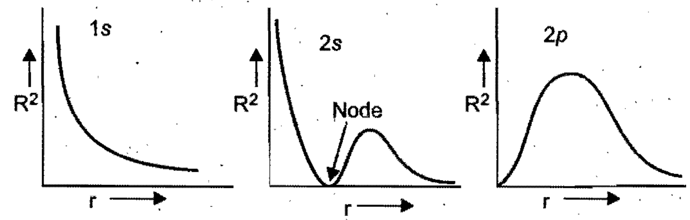
$$\psi_{nlm}(r,\theta,\phi) = R_{n,l}(r)\,\Theta_{lm}(\theta)\,\Phi_m(\phi)$$

29. Graph of radial wave function '*R*': At node, the value of '*R*' changes from positive to negative.

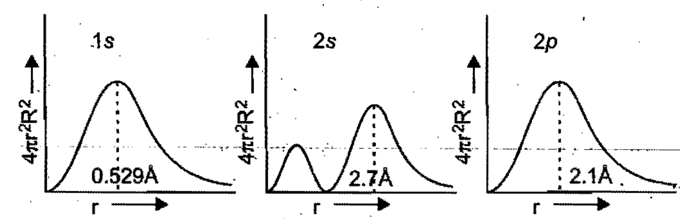


Number of radial nodes = $(n - l - 1)$.

30. Plot of radial probability density 'R^2':



31. Plot of radial probability function ($4\pi r^2 R^2$):



In the plot of radial probability against 'r', number peaks, *i.e.*, region of maximum probability = $n - l$.

Questions

1. Match the following:

[A]

(i) Aufbau principle	(a) Line spectrum in visible region
(ii) de Broglie	(b) Orientation of an electron in an orbital
(iii) Angular momentum	(c) Photon
(iv) Hund's rule	(d) $\lambda = h/(mv)$
(v) Balmer series	(e) Electronic configuration
(vi) Planck's law	(f) mvr

[B]

(i) Thomson	(a) Exclusion principle
(ii) Pauli	(b) Radioactivity
(iii) Becquerel	(c) Atomic model
(iv) Soddy	(d) Cathode rays
(v) Bohr	(e) Neutron
(vi) Chadwick	(f) Isotopes

[C]

(i) Cathode rays	(a) Helium nuclei
(ii) Dumb-bell	(b) Uncertainty principle
(iii) Alpha particles	(c) Electromagnetic radiation
(iv) Moseley	(d) p-orbital
(v) Heisenberg	(e) Atomic number
(vi) X-rays	(f) Electrons

2. **Matrix Matching Problems (For IIT Aspirants):**

[A] Match the Column-I and Column-II:

Column-I	Column-II
(a) X-rays	(p) Davisson and Germer experiment
(b) Atomic number determination	(q) Crystal structure determination
(c) Dual nature of matter	(r) Moseley's law
(d) Dual nature of radiation	(s) Bragg's law

[B] Match the Column-I and Column-II:

Column-I	Column-II
(a) Lyman series	(p) Visible region
(b) Balmer series	(q) UV region
(c) Pfund series	(r) IR region
(d) Light emitted by sodium lamp	(s) Line emission spectrum

[C] Match the List-I with List-II in hydrogen atom spectrum:

List-I	List-II
(a) Lyman series	(p) Visible region
(b) Balmer series	(q) Infrared region
(c) Paschen series	(r) Absorption spectrum
(d) Brackett series	(s) Ultraviolet region

[D] Match the List-I with List-II:

List-I	List-II
(a) K-shell	(p) Electrons in elliptical orbit
(b) L-shell	(q) Electrons in circular orbit
(c) Hydrogen atom	(r) Shell of lowest energy
(d) Boron atom in ground state	(s) Bohr's atomic model

[E] Match the ions of List-I with the properties of List-II:

List-I	List-II
(a) Mn^{2+}	(p) Diamagnetic
(b) V^{2+}	(q) Paramagnetic
(c) Zn^{2+}	(r) Coloured compounds
(d) Ti^{4+}	(s) Magnetic moment = 2.82 BM

[F] Match the List-I with List-II:

List-I	List-II
(a) Mg^{2+}	(p) Zero spin multiplicity
(b) Fe^{2+}	(q) Spin multiplicity = 3
(c) Co^{3+}	(r) Total spin = 0
(d) Ca^{2+}	(s) Total spin = 2

[G] Match the properties of List-I with the formulae in List-II:

List-I	List-II
(a) Angular momentum of electron	(p) $\sqrt{l(l+1)}\dfrac{h}{2\pi}$
(b) Orbital angular momentum	(q) $I\omega$
(c) Wavelength of matter wave	(r) $nh/2\pi$
(d) Quantised value(s)	(s) h/p

[H] Match the orbitals of List-I with the nodal properties of List-II:

List-I	List-II
(a) $2s$	(p) Angular node = 1
(b) $1s$	(q) Radial node = 0
(c) $2p$	(r) Radial node = 1
(d) $3p$	(s) Angular node = 0

[I] Match the electronic transitions of List-I with spectral properties of List-II:

List-I	List-II
(a) $n = 6 \longrightarrow n = 3$	(p) 10 lines in the spectrum
(b) $n = 7 \longrightarrow n = 3$	(q) Spectral lines in visible region
(c) $n = 5 \longrightarrow n = 2$	(r) 6 lines in the spectrum
(d) $n = 6 \longrightarrow n = 2$	(s) Spectral lines in infrared region

[J] Match the List-I with List-II:

List-I	List-II
(a) Radius of electron orbit	(p) Principal quantum number
(b) Energy of electron	(q) Azimuthal quantum number
(c) Energy of subshell	(r) Magnetic quantum number
(d) Orientation of the atomic orbitals	(s) Spin quantum number

[K] Match the List-I with List-II:

List-I	List-II
(a) Electron cannot exist in the nucleus	(p) de Broglie wave
(b) Microscopic particles in motion are associated with	(q) Electromagnetic wave
(c) No medium is required for propagation	(r) Uncertainty principle
(d) Concept of orbit was replaced by orbital	(s) Transverse wave

[L] According to Bohr theory: **(IIT 2006)**

E_n = Total energy
K_n = Kinetic energy
V_n = Potential energy
r_n = Radius of nth orbit

Match the following:

Column-I	Column-II
(a) $V_n / K_n = ?$	(p) 0
(b) If radius of nth orbit $\propto En^x$; $x = ?$	(q) -1
(c) Angular momentum in lowest orbital	(r) -2
(d) $\frac{1}{r^n} \propto Z^y$; $y = ?$	(s) 1

[M] Match the List-I with List-II:

List-I	List-II
(a) Radius of nth orbital	(p) Inversely proportional to Z
(b) Energy of nth shell	(q) Integral multiple of $h/2\pi$
(c) Angular momentum of electron	(r) Proportional to n^2
(d) Velocity of electron in nth orbit	(s) Inversely proportional to 'n'

[N] Match the entries in Column-I with the correctly related quantum number(s) in Column-II:

Column-I	Column-II
(a) Orbital angular momentum of the electron in a hydrogen-like atomic orbital	(p) Principal quantum number
(b) A hydrogen-like one electron wave function obeying Pauli principle	(q) Azimuthal quantum number
(c) Shape, size and orientation of hydrogen-like atomic orbitals	(r) Magnetic quantum number
(d) Probability density of electron at the nucleus in hydrogen-like atom	(s) Electron spin quantum number

[O] Match the List-I with List-II: **(IIT 2006)**

List-I	List-II
(a) Wave nature of radiation	(p) Photoelectric effect
(b) Photon nature of radiation	(q) Compton effect
(c) Interaction of a photon with an electron, such that quantum energy is slightly equal to or greater than the binding energy of electron, is more likely to result in:	(r) Diffraction
(d) Interaction of a photon with an electron, such that photon energy is much greater than the binding energy of electron, is more likely to result in:	(s) Interference

[P] Match the Column-I with Column-II:

Column-I	Column-II
(a) Orbital angular momentum of an electron	(p) $\sqrt{s(s+1)}\, h/2\pi$
(b) Angular momentum of electron	(q) $\sqrt{n(n+2)}$ BM
(c) Spin angular momentum of electron	(r) $nh/2\pi$
(d) Magnetic moment of atom	(s) $\sqrt{l(l+1)}\, h/2\pi$

[Q] Match the Column-I with Column-II:

Column-I	Column-II
(a) Scintillation	(p) Wave nature
(b) Photoelectric effect	(q) Particle nature
(c) Diffraction	(r) Particle nature dominates over wave nature
(d) Principle of electron microscope	(s) Wave nature dominates over particle nature

[R] Match the Column-I with Column-II:

Column-I	Column-II
(a) Radial function R	(p) Principal quantum number 'n'
(b) Angular function Θ	(q) Azimuthal quantum number 'l'
(c) Angular function Φ	(r) Magnetic quantum number 'm'
(d) Quantized angular momentum	(s) Spin quantum number 's'

Answers

1. [A] (i—e); (ii—d); (iii—f); (iv—b); (v—a); (vi—c)
   [B] (i—d); (ii—a); (iii—b); (iv—f); (v—c); (vi—e)
   [C] (i—f); (ii—d); (iii—a); (iv—e); (v—b); (vi—c)
2. [A] (a—q, r, s) (b—r) (c—p) (d—does not match)
   [B] (a—q, s) (b—p, s) (c—r, s) (d—p, s)
   [C] (a—r, s) (b—p) (c—q) (d—q)
   [D] (a—q, r) (b—p, q) (c—s) (d—p, q)
   [E] (a—q, r) (b—q, r, s) (c—p) (d—p)
   [F] (a—p, r) (b—q, s) (c—q, s) (d—p, r)
   [G] (a—q, r) (b—p) (c—s) (d—q, r)
   [H] (a—r, s) (b—q, s) (c—q, p) (d—p, r)
   [I] (a—r, s) (b—p, s) (c—q, r) (d—p, q)
   [J] (a—p) (b—p) (c—p, r) (d—r)
   [K] (a—r) (b—p) (c—q, s) (d—r)
   [L] (a—r) (b—q) (c—p) (d—s)
   [M] (a—r, p) (b—r) (c—q) (d—s)
   [N] (a—p) (b—s) (c—p, q, r) (d—p, q)
   [O] (a—r, s) (b—p, q) (c—p) (d—q)
   [P] (a—s) (b—r) (c—p) (d—q)
   [Q] (a—q) (b—r) (c—p) (d—p, s)
   [R] (a—p, q) (b—q, r) (c—r) (d—q, s)

• PRACTICE PROBLEMS •

1. An atom of an element contains 13 electrons. Its nucleus has 14 neutrons. Find out its atomic number and approximate atomic mass. An isotope has atomic mass 2 units higher. What will be the number of protons, neutrons and electrons in the isotope?
   [**Ans.** At. No. = 13, atomic mass = 27; the isotope will have same number of protons and electrons = 13 but neutrons will be 14 + 2 = 16]
2. From the following find out groups of isotopes, isobars and isotones:
   $$^{16}_{8}O,\ ^{39}_{19}K,\ ^{14}_{6}C,\ ^{239}_{92}U,\ ^{14}_{7}N,\ ^{40}_{20}Ca,\ ^{238}_{92}U,\ ^{77}_{32}Ge,$$
   $$^{77}_{33}As,\ ^{18}_{8}O,\ ^{76}_{32}Ge,\ ^{78}_{34}Se$$
   [**Ans.** Isotopes—same at. no. but different at. masses.
   $$^{16}_{8}O,\ ^{18}_{8}O;\ ^{239}_{92}U,\ ^{238}_{92}U;\ ^{77}_{32}Ge,\ ^{76}_{32}Ge$$
   Isobars—same atomic masses but different at. numbers
   $$^{14}_{6}C,\ ^{14}_{7}N;\ ^{77}_{32}Ge,\ ^{77}_{33}As$$
   Isotones—same number of neutrons.
   $$^{16}_{8}O,\ ^{14}_{6}C;\ ^{39}_{19}K,\ ^{40}_{20}Ca;\ ^{77}_{33}As,\ ^{78}_{34}Se]$$
3. An element has atomic number 30. Its cation has 2 units positive charge. How many protons and electrons are present in the cation?
   [**Ans.** Protons = 30, Electrons = 28]
4. Calculate the number of neutrons in 18 mL of water. (Density of water = 1)
   [**Ans.** 48.16×10^{23}]
   [**Hint:** One molecule of water contains = 8 neutrons]
5. Find (i) the total number of neutrons and (ii) the total mass of neutrons in 7 mg of ^{14}C (assuming that mass of neutron = mass of hydrogen atom).
   [**Ans.** (i) 24.08×10^{20} and (ii) 4 mg]
6. Calculate the wavelength of a photon in Angstroms having an energy of 1 electron volt.
   [**Hint:** $1\ eV = 1.602 \times 10^{-19}$ joule;
   $$h = 6.62 \times 10^{-34}\ J\text{-}s,\ c = 3 \times 10^{8}\ ms^{-1}$$
   $$\lambda = \frac{h \cdot c}{E} = 12.42 \times 10^{-7}\ m = 12.42 \times 10^{3}\ Å]$$
7. A photon of light with wavelength 6000 Å has an energy E. Calculate the wavelength of photon of a light which corresponds to an energy equal to $2E$.
   [**Ans.** 3000 Å]
8. Calculate the energy in kilocalorie per mol of the photons of an electromagnetic radiation of wavelength 5700 Å.
   [**Ans.** 56.3 kcal per mol]

9. Light of what frequency and wavelength is needed to ionise sodium atom. The ionisation potential of sodium is 8.2×10^{-19} J.

[**Ans.** $\nu = 1.238 \times 10^{15}$ Hz; $\lambda = 242$ nm]

10. Determine the energy of 1 mole photons of radiations whose frequency is 5×10^{10} s^{-1}. ($h = 6.62 \times 10^{-34}$ J-s)

[**Ans.** 19.9 J]

11. Find e/m for He^{2+} ion and compare with that for electron.
[**Ans.** 4.87×10^7 coulomb kg^{-1}]

12. A ball of mass 100 g is moving with a velocity of 100 m sec^{-1}. Find its wavelength.

[**Hint:** $\lambda = \dfrac{h}{mv} = \dfrac{6.626 \times 10^{-34}}{0.1 \times 100} = 6.626 \times 10^{-35}$ m]

13. Calculate the wavelength of radiation and energy per mol necessary to ionize a hydrogen atom in the ground state.
[**Ans.** $\lambda = 9.12 \times 10^{-8}$ m; 1313 kJ/mol]

14. Bond energy of F_2 is 150 kJ mol^{-1}. Calculate the minimum frequency of photon to break this bond.
[**Ans.** 3.759×10^{14} s^{-1}]

15. If an Einstein (E) is the total energy absorbed by 1 mole of a substance and each molecule absorbs one quantum of energy, then calculate the value of 'E' in terms of λ in cm.

[**Ans.** $\dfrac{1.198 \times 10^8}{\lambda}$ erg mol^{-1}]

16. How many chlorine atoms can you ionize in the process? $Cl \rightarrow Cl^+ + e$ by the energy liberated from the following process:

$$Cl + e \rightarrow Cl^- \text{ for } 6 \times 10^{23} \text{ atoms}$$

given that electron affinity of chlorine is 3.61 eV and ionization energy of chlorine is 17.422 eV.

[**Ans.** 1.24×10^{23} atoms]

17. Find the velocity (ms^{-1}) of electron in first Bohr orbit of radius a_0. Also find the de Broglie wavelength (in 'm'). Find the orbital angular momentum of $2p$ orbital of hydrogen atom in units of $h/2\pi$.

[**Hint:** $v = \dfrac{2.188 \times 10^6}{n}$ m sec^{-1}

$$v = \frac{2.188 \times 10^6}{1} = 2.188 \times 10^6 \text{ m sec}^{-1}$$

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}}{9.1 \times 10^{-31} \times 2.188 \times 10^6}$$

$$= 3.3 \times 10^{-10} \text{ m}$$

$$\text{Orbital angular momentum} = \sqrt{l(l+1)}\,\frac{h}{2\pi}$$

$$= \sqrt{1(1+1)}\,\frac{h}{2\pi} \qquad (\because l = 1 \text{ for } 2p)$$

$$= \sqrt{2}\,\frac{h}{2\pi}\text{]}$$

18. The energy of an α-particle is 6.8×10^{-18} J. What will be the wavelength associated with it? **[CBSE-PMT (Mains) 2005]**

[**Hint:** $\lambda = \dfrac{h}{\sqrt{2Em}} = \dfrac{6.626 \times 10^{-34}}{\sqrt{2 \times 6.8 \times 10^{-18} \times 4 \times 1.66 \times 10^{-27}}}$

$= 2.2 \times 10^{-12}$ m]

19. Determine the number of revolutions made by an electron in one second in the 2nd Bohr orbit of H-atom.

[**Ans.** $n = \dfrac{2\pi r}{v}$]

20. What is the speed of an electron whose de Broglie wavelength is 0.1 nm? By what potential difference, must have such an electron accelerated from an initial speed zero?
[**Ans.** 7.28×10^6 m/sec; 150 V]

21. A green ball weighs 75 g; it is travelling towards observer at a speed of 400 cm/sec. The ball emits light of wavelength 5×10^{-5} cm. Assuming that the error in the position of ball is the same as wavelength of itself, calculate error in the momentum of the green ball.

[**Hint:** $\Delta x \cdot \Delta p \geq \dfrac{h}{4\pi}$

$$\Delta p \geq \frac{h}{4\pi\,\Delta x}$$

$$\Delta p \geq \frac{h}{4\pi\lambda}$$

$$\Delta p \approx \frac{6.626 \times 10^{-27}}{4 \times 3.14 \times 5 \times 10^{-5}} \approx 1.055 \times 10^{-23}\text{]}$$

22. What is the relationship between the eV and the wavelength in metre of the energetically equivalent photons?
[**Ans.** $\lambda = 12.4237 \times 10^{-7}$ metre]

23. What is the velocity of an electron ($m = 9.11 \times 10^{-31}$ kg) in the innermost orbit of the hydrogen atom?
(Bohr radius $= 0.529 \times 10^{-10}$ m)

[**Ans.** 2.187×10^6 m/sec]

24. In a hydrogen atom, an electron jumps from the third orbit to the first orbit. Find out the frequency and wavelength of the spectral line. ($R_H = 1.09678 \times 10^7$ m^{-1})

[**Ans.** 2.925×10^{15} Hz, 1025.6 Å]

25. The energy of the electron in the second and third Bohr orbits of hydrogen atom is -5.42×10^{-12} erg and -2.41×10^{-12} erg respectively. Calculate the wavelength of the emitted radiation when the electron drops from third to second orbit.
[**Ans.** 6.6×10^3 Å]

26. Calculate the wavelength in angstroms of the photon that is emitted when an electron in Bohr orbit $n = 2$ returns to the orbit $n = 1$ in the hydrogen atom. The ionisation potential of the ground state of hydrogen atom is 2.17×10^{-11} erg per atom.

[**Hint:** Energy of the electron in the 1st orbit = – (ionisation potential), $\Delta E = (3/4) \times 2.17 \times 10^{-11}$ erg per atom]

[**Ans.** $\lambda = 1220$ Å]

27. Calculate the wave number for the shortest wavelength transition in Balmer series of atomic hydrogen. **(IIT 1996)**
[**Ans.** 27419.25 cm^{-1}]

28. The wavelength of the first member of the Balmer series of hydrogen is 6563×10^{-10} m. Calculate the wavelength of its second member.

[Hint: $\frac{1}{\lambda_1} = R_H \left[\frac{1}{2^2} - \frac{1}{3^2}\right]$ and $\frac{1}{\lambda_2} = R_H \left[\frac{1}{2^2} - \frac{1}{4^2}\right]$

$$\frac{\lambda_2}{\lambda_1} = \frac{5}{36} \times \frac{16}{3} = \frac{20}{27}$$

$$\lambda_2 = \frac{20}{27} \times 6563 \times 10^{-10} = 4861 \times 10^{-10} \text{ m}]$$

29. According to Bohr theory, the electronic energy of hydrogen atom in the nth Bohr orbit is given by,

$$E_n = -\frac{21.76 \times 10^{-19}}{n^2} \text{ J}$$

Calculate the longest wavelength of light that will be needed to remove an electron from the 2nd orbit of Li^{2+} ion.

[Ans. 4.059×10^{-8} m]

30. Calculate the frequency, energy and wavelength of the radiation corresponding to spectral line of lowest frequency in Lyman series in the spectra of hydrogen atom. Also calculate the energy of the corresponding line in the spectra of Li^{2+}.

(IIT 1991)

[Ans. $\lambda = 1.216 \times 10^{-7}$ m, $v = 2.47 \times 10^{15}$ cycle sec^{-1}, $E = 16.36 \times 10^{-19}$ J, $E_{Li^{2+}} = Z^2 \times E_H = 9 \times 16.36 \times 10^{-19}$ J $= 147.27 \times 10^{-19}$ J]

31. Calculate the ratio of the velocity of light and the velocity of electron in the 2nd orbit of a hydrogen atom. (Given $h = 6.624 \times 10^{-27}$ erg-sec; $m = 9.108 \times 10^{-28}$ g; $r = 2.11 \times 10^{-8}$ cm)

[Ans. 273.2]

32. What hydrogen-like ion has the wavelength difference between the first lines of Balmer and Lyman series equal to 59.3 nm ($R_H = 109678$ cm^{-1})?

[Hint: Wavelength of 1st line in Balmer series,

$$\frac{1}{\lambda_B} = Z^2 R_H \left[\frac{1}{2^2} - \frac{1}{3^2}\right] = \frac{5}{36} R_H Z^2$$

or $$\lambda_B = \frac{36}{5 R_H Z^2}$$

Wavelength of 1st line in Lyman series is,

$$\frac{1}{\lambda_L} = Z^2 R_B \left[\frac{1}{1^2} - \frac{1}{2^2}\right]$$

or $$\lambda_L = \frac{4}{3 \times R_H Z^2}$$

$$\text{Difference } \lambda_B - \lambda_L = 59.3 \times 10^{-7} = \frac{36}{5 R_H Z^2} - \frac{4}{3 R_H Z^2}$$

$$= \frac{1}{R_H Z^2}\left[\frac{36}{5} - \frac{4}{3}\right]$$

$$Z^2 = \frac{88}{59.3 \times 10^{-7} \times 109678 \times 15} = 9.0$$

or $Z = 3$

Hydrogen-like species is Li^{2+}.]

33. The velocity of an electron in certain Bohr orbit of H-atom bears the ratio 1 : 275 to the velocity of light. (a) What is the quantum number 'n' of the orbit? (b) Calculate the wave number of the radiation when the electron jumps from $(n + 1)$ state to ground state.

[Ans. $\bar{v} = 9.75 \times 10^4$ cm^{-1}]

[Hint: (a) $\frac{v}{c} = \frac{1}{275}$ or $v = \frac{3 \times 10^{10}}{275} = 1.09 \times 10^8$ cm sec^{-1}

$$v = \frac{nh}{2\pi mr} = \frac{nh}{2\pi m \times 0.529 \times 10^{-8} \times n^2}$$

or $$n = \frac{h}{2\pi m \times 0.529 \times 10^{-8} \times v}$$

$$= \frac{6.625 \times 10^{-27}}{2 \times 3.14 \times 9.1 \times 10^{-28} \times 0.529 \times 10^{-8} \times 1.09 \times 10^8}$$

$$= 2$$

(b) Thus, $n + 1 = 2 + 1 = 3$. The electron jumps from 3rd orbit to 1st orbit.]

34. Find out the wavelength of the next line in the series having lines of spectrum of H-atom of wavelengths 6565 Å, 4863 Å, 4342 Å and 4103 Å.

[Ans. 3972 Å]

[Hint: All these lines are in visible region and thus, belong to Balmer series. Next line is, therefore, from 7th orbit.]

35. Which jump is responsible for the wave number of emitted radiations equal to 9.7490×10^6 m^{-1} in Lyman series of hydrogen spectrum? ($R = 1.09678 \times 10^7$ m^{-1})

[Ans. 3]

36. Calculate the ionisation energy of the hydrogen atom. How much energy will be required to ionise 1 mole of hydrogen atoms? Given, that the Rydberg constant is 1.0974×10^7 m^{-1}.

[Ans. IE per hydrogen atom $= 2.182 \times 10^{-18}$ J
IE per mole $= 1314$ kJ mol^{-1}]

37. Calculate the ionisation energy of (a) one Li^{2+} ion and (b) one mole of Li^{2+} ion. (Given, $R = 1.0974 \times 10^{-7}$ m^{-1})

[Ans. (a) 19.638×10^{-18} J (b) 1.118×10^4 kJ mol^{-1}]

38. A series of lines in the spectrum of atomic hydrogen lies at 656.46 nm, 486.27 nm, 439.17 nm and 410.29 nm. What is the wavelength of the next line in this series? What is the ionisation energy of the atom when it is in the lower state of transition?

[Ans. $\lambda_{next} = 397.15$ nm; IE = 3.40 eV]

39. A certain line of the Lyman series of hydrogen and a certain line of the Balmer series of He^+ ion have nearly the same wavelength. To what transition do they belong? Small differences between their Rydberg constant may be neglected.

[Ans.

Hydrogen	Helium
$2 \to 1$	$4 \to 2$
$3 \to 1$	$6 \to 2$
$4 \to 1$	$8 \to 2$]

40. What element has a hydrogen-like spectrum whose lines have wavelengths four times shorter than those of atomic hydrogen?

[Ans. He^+]

41. What lines of atomic hydrogen absorption spectrum fall within the wavelength ranges from 94.5 to 130 nm?
[**Ans.** 97.3; 102.6; 121.6 nm]

42. The binding energy of an electron in the ground state of an atom is equal to 24.6 eV. Find the energy required to remove both the electrons from the atom.
[**Ans.** 79 eV]

43. What is the ratio of the speeds of an electron in the first and second orbits of a hydrogen atom?
[**Ans.** 2 : 1]

44. Find out the number of waves made by a Bohr electron in one complete revolution in its third orbit. **(IIT 1994)**
[**Ans.** 3]

45. The wave number of first line in Balmer series of hydrogen is $15200\ cm^{-1}$. What is the wave number of first line in Balmer series of Be^{3+}?
[**Ans.** $2.43 \times 10^5\ cm^{-1}$]

46. Calculate the speed of an electron in the ground state of hydrogen atom. What fraction of the speed of light is this value? How long does it take for the electron to complete one revolution around the nucleus? How many times does the electron travel around the nucleus in one second?
[**Ans.** $2.186 \times 10^6\ ms^{-1}; 7.29 \times 10^{-3}$]

47. An electron, in a hydrogen atom, in its ground state absorbs 1.5 times as much energy as the minimum required for its escape (*i.e.*, 13.6 eV) from the atom. Calculate the value of λ for the emitted electron.
[**Ans.** 4.69 Å]

48. The radius of the fourth orbit of hydrogen is 0.85 nm. Calculate the velocity of an electron in this orbit ($m_e = 9.1 \times 10^{-31}$ kg).
[**Ans.** $5.44 \times 10^5\ m\ sec^{-1}$]

49. A beam of electrons accelerated with 4.64 V was passed through a tube having mercury vapours. As a result of absorption, electronic changes occurred with mercury atoms and light was emitted. If the full energy of single electron was converted into light, what was the wave number of emitted light?
[**Ans.** $3.75 \times 10^4\ cm^{-1}$]

50. An electron jumps from an outer orbit to an inner orbit with the energy difference of 3.0 eV. What will be the wavelength of the line and in what region does the emission take place?
[**Ans.** $\lambda = 4140$ Å ; visible region]
[**Hint:** $1\ eV = 1.6 \times 10^{-12}$ erg]

51. The first ionisation energy of a certain atom took place with an absorption of radiation of frequency 1.5×10^{18} cycle per second. Calculate its ionisation energy in calorie per gram atom.
[**Ans.** 1.43×10^8 cal]
[**Hint:** 1 calorie $= 4.18 \times 10^7$ erg
Apply $E = h \times \nu \times$ Avogadro's number]

52. Find the wavelength associated with an electron which has mass 9.1×10^{-28} g and is moving with a velocity of 10^5 cm sec^{-1}. (Given $h = 6.625 \times 10^{-27}$ erg-sec)
[**Ans.** $\lambda = 7.28 \times 10^{-5}$ cm]

53. Calculate the momentum of the particle which has de Broglie wavelength 1 Å (10^{-10} m) and $h = 6.6 \times 10^{-34}$ J-sec.
[**Ans.** 6.6×10^{-24} kg m sec^{-1}]

54. The uncertainty of a particle in momentum is 3.3×10^{-2} kg ms^{-1}. Calculate the uncertainty in its position. ($h = 6.6 \times 10^{-34}$ J-sec)
[**Ans.** 3.1×10^{-14} m]

55. Calculate the product of uncertainties of displacement and velocity of a moving electron having a mass 9.1×10^{-28} g.
[**Ans.** $5.77 \times 10^{-5}\ m^2\ s^{-1}$]
[**Hint:** $\Delta x \cdot \Delta v \geq \dfrac{h}{4\pi m}$]

56. (a) A transition metal cation x^{3+} has magnetic moment $\sqrt{35}$ BM. What is the atomic number of x^{3+}?
(b) Select the coloured ion and the ion having maximum $\boxed{1}\boxed{1}\boxed{1}\boxed{1}\boxed{1}$ $\boxed{}$ magnetic moment (i) Fe^{2+}, (ii) Cu^{+}, (iii) Sc^{3+} and (iv) Mn^{2+}

(b) $Fe^{2+} \rightarrow$ $\boxed{1\!\downarrow}\boxed{1}\boxed{1}\boxed{1}\boxed{1}$

$Mn^{2+} \rightarrow$ $\boxed{1}\boxed{1}\boxed{1}\boxed{}\boxed{}$

[**Hint:** (a) 26, $_{26}Fe \rightarrow 3d^6 4s^2$
$Fe^{3+} \rightarrow 3d^5 4s^0$
$\mu = \sqrt{n(n+2)} = \sqrt{5 \times 7} = \sqrt{35}$
Both these ions will be coloured and magnetic moment of Fe^{2+} will be greater.]

57. A photon of wavelength 4000 Å strikes a metal surface, the work function of the metal being 2.13 eV. Calculate (i) energy of the photon in eV; (ii) kinetic energy of the emitted photoelectron and (iii) velocity of the photoelectron.
[**Ans.** $E = 3.10$ eV; KE = 0.97 eV; Velocity $= 5.85 \times 10^5\ ms^{-1}$]
[**Hint:** $1\ eV = 1.602 \times 10^{-19}$ J]

58. Calculate the ratio between the wavelengths of an electron and a proton, if the proton is moving at half the velocity of the electron (mass of the proton $= 1.67 \times 10^{-27}$ kg; mass of the electron $= 9.11 \times 10^{-28}$ g).
[**Ans.** 9.2455×10^{-2} m]
[**Hint:** Apply de Broglie equation, $\lambda = \dfrac{h}{mv}$

$$\text{Wavelength of electron} = \frac{6.625 \times 10^{-34}}{9.11 \times 10^{-31} v}$$

$$\text{Wavelength of proton} = \frac{6.625 \times 10^{-34}}{1.67 \times 10^{-27} \times 0.5v}$$]

59. A moving electron has 2.8×10^{-25} J of kinetic energy. Calculate its wavelength.
(Mass of electron $= 9.1 \times 10^{-31}$ kg)
[**Ans.** 9.2455×10^{-7} m]

[**Hint:** $v = \sqrt{\frac{2 \times KE}{m}} = 784.46\ ms^{-1}$; $\lambda = \frac{h}{mv}$]

60. Helium has mass number 4 and atomic number 2. Calculate the nuclear binding energy per nucleon (mass of neutron = 1.00893 amu and proton = 1.00814 amu, He = 4.0039 amu and mass of electron is negligible).

[**Ans.** 7.038 MeV]

61. Calculate the mass defect and binding energy per nucleon of $^{16}_{8}O$ which has a mass 15.99491 amu.

Mass of neutron = 1.008655 amu
Mass of proton = 1.007277 amu
Mass of electron = 0.0005486 amu
1 amu = 931.5 MeV

[**Ans.** 7.976 MeV/nucleon]

62. The circumference of the second Bohr orbit of electron in the hydrogen atom is 600 nm. Calculate the potential difference to which the electron has to be subjected so that the electron stops. The electron had the de Broglie wavelength corresponding to the circumference.

[**Hint:** Number of waves 'n' = $\frac{\text{Circumference}}{\text{Wavelength}}$

$$n\lambda = 2\pi r$$
$$2\lambda = 600$$
$$\lambda = 300 \text{ nm}$$

Let stopping potential is V_0.

$$eV_0 = \frac{1}{2}mv^2 \quad \ldots (i)$$
$$\lambda = \frac{h}{mv}$$
$$v = \frac{h}{\lambda m} \quad \ldots (ii)$$

From equations (i) and (ii),

$$eV_0 = \frac{1}{2}m\left(\frac{h}{\lambda m}\right)^2$$
$$V_0 = \frac{h^2}{2m\lambda^2 e}$$
$$= \frac{(6.626 \times 10^{-34})^2}{2 \times (9.1 \times 10^{-31}) \times (300 \times 10^{-9})^2 \times 1.6 \times 10^{-19}}$$
$$= 1.675 \times 10^{-5} \text{ V}]$$

63. The velocity of an electron of mass 9.1×10^{-31} kg moving round the nucleus in the Bohr orbit (diameter of the orbit is 1.058 Å) is 2.2×10^{-6} m sec^{-1}. If momentum can be measured within the accuracy of 1%, then calculate uncertainty in position (Δx) of the electron.

[**Ans.** 2.64×10^3 metre]

64. An electron wave has wavelength 1 Å. Calculate the potential with which the electron is accelerated.

[**Ans.** 0.0826 volt]

65. Calculate the de Broglie wavelength associated with an α-particle having an energy of 7.7×10^{-13} J and a mass of 6.6×10^{-24} g. ($h = 6.6 \times 10^{-34}$ J-s)

[**Ans.** 6.56×10^{-13} cm]

66. An electron has mass 9.1×10^{-28} g and is moving with a velocity of 10^5 cm/sec. Calculate its kinetic energy and wavelength when $h = 6.626 \times 10^{-27}$ erg-sec.

[**Ans.** 4.55×10^{-8} erg; $\lambda = 7.28 \times 10^{-5}$ cm]

67. Calculate the de Broglie wavelengths of an electron and a proton having same kinetic energy of 100 eV.

[**Ans.** $\lambda_e = 123$ pm; $\lambda_p = 2.86$ pm]

68. Work function of sodium is 2.5 eV. Predict whether the wavelength 6500 Å is suitable for a photoelectron or not?

[**Ans.** No ejection]

69. Calculate the de Broglie wavelength associated with a helium atom in a helium gas sample at 27°C and 1 atm pressure.

[**Ans.** 7.3×10^{-11} metre]

70. The threshold frequency for a certain metal is 3.3×10^{14} cycle/sec. If incident light on the metal has a cut-off frequency 8.2×10^{14} cycle/sec, calculate the cut-off potential for the photoelectron.

[**Ans.** 2 volt]

71. Can you locate the electron within 0.005 nm?

[**Ans.** No.]

[**Hint:** Use uncertainty principle to determine uncertainty in velocity.

$$\Delta v \geq \frac{h}{4\pi m\, \Delta x}$$

On substitution, you will get,

$$\Delta v \geq 1.16 \times 10^7\ ms^{-1}$$

Velocity of electron is therefore expected to be as high as velocity of light. We may say that the velocity of electron is uncertain within 0.005 nm.]

72. The photoelectric cut-off voltage in a certain experiment is 1.5 volt. What is the maximum kinetic energy of the photoelectrons emitted?

[**Ans.** 2.4×10^{-19} joule]

73. A proton is accelerated to one-tenth the velocity of light. If its velocity can be measured with a precision of ±1%, what must be its uncertainty in position?
($h = 6.6 \times 10^{-34}$ J-s; mass of proton = 1.66×10^{-27} kg)

[**Ans.** 1.05×10^{-14} m]

74. In a photoelectric effect experiment, irradiation of a metal with light of frequency 5.2×10^{14} sec^{-1} yields electrons with maximum kinetic energy 1.3×10^{-19} J. Calculate the ν_0 of the metal.

[**Ans.** 3.2×10^{14} sec^{-1}]

75. Calculate the wavelength of a CO_2 molecule moving with a velocity of 440 m sec^{-1}.

[**Ans.** 2.06×10^{-11} metre]

76. The predominant yellow line in the spectrum of a sodium vapour lamp has a wavelength of 590 nm. What minimum accelerating potential is needed to excite this line in an electron tube having sodium vapours?

[**Ans.** 2.11 volt]

77. Find out the wavelength of a track star running a 100 metre dash in 10.1 sec, if its weight is 75 kg.

[Ans. 8.92×10^{-37} m]

78. At what velocity ratio are the wavelengths of an electron and a proton equal?

($m_e = 9.1 \times 10^{-28}$ g and $m_p = 1.6725 \times 10^{-24}$ g)

[Ans. $\frac{v_e}{v_p} = 1.8 \times 10^3$]

79. Through what potential difference must an electron pass to have a wavelength of 500 Å?

[Ans. 6.03×10^{-4} eV]

[Hint: Use $\lambda = \frac{h}{\sqrt{2\,eV\,m}}$]

80. Calculate the velocity of an α-particle which begins to reverse its direction at a distance of 2×10^{-14} m from a scattering gold nucleus ($Z = 79$).

[Ans. 2.346×10^7 m / sec]

81. Two hydrogen atoms collide head-on and end up with zero kinetic energy. Each then emits a photon with a wavelength 121.6 nm. Which transition leads to this wavelength? How fast were the hydrogen atoms travelling before the collision? (Given, $R_H = 1.097 \times 10^7 \text{ m}^{-1}$ and $m_H = 1.67 \times 10^{-27}$ kg)

[Ans. $n_1 = 1$; $n_2 = 2$; 4.43×10^4 m sec^{-1}]

[Hint: Wavelength is in UV region; thus n_1 will be 1.

$$\frac{1}{121.6 \times 10^{-9}} = 1.097 \times 10^7 \times 1^2 \times \left(\frac{1}{1^2} - \frac{1}{n_2^2}\right)$$

$$n_2 = 2$$

$$\frac{1}{2} mv^2 = \frac{hc}{\lambda}$$

$$\frac{1}{2} \times 1.67 \times 10^{-27} \times v^2 = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{121.6 \times 10^{-9}}$$

$$v = 4.43 \times 10^4 \text{ m sec}^{-1}]$$

82. Show that the wavelength of electrons moving at a velocity very small compared to that of light and with a kinetic energy of V electron volt can be written as,

$$\lambda = \frac{12.268}{\sqrt{V}} \times 10^{-8} \text{ cm}$$

[Hint: Use the relation, $\lambda = \frac{h}{\sqrt{2Em}}$

Here, h = Planck's constant

$m = 9.1 \times 10^{-28}$ g (mass of e^-)

E = Kinetic energy of electron

$= V$ eV $= V \times 1.6 \times 10^{-12}$ erg]

83. What is the distance of closest approach to the nucleus of an α-particle which undergoes scattering by 180° in Geiger-Marsden experiment?

[Ans. $r_0 = 4.13$ fm]

[Hint: For closest approach,

$$\frac{1}{2} mv^2 = K \frac{Ze \times e}{r_0}$$

For Rutherford experiment,

$$\frac{1}{2} mv^2 = 5.5 \text{ MeV} = 5.5 \times 10^6 \times 1.6 \times 10^{-19} \text{ J} = 8.8 \times 10^{-13} \text{ J}$$

$$8.8 \times 10^{-13} = \frac{9 \times 10^9 \times 2 \times 79 \times (1.6 \times 10^{-19})^2}{r_0}$$

$$r_0 = 4.136 \times 10^{-15} \text{ m}$$

$$r_0 = 4.13 \text{ fm}]$$

84. Photoelectrons are liberated by ultraviolet light of wavelength 3000 Å from a metallic surface for which the photoelectric threshold is 4000 Å. Calculate de Broglie wavelength of electrons emitted with maximum kinetic energy.

[Ans. $\lambda = 1.2 \times 10^{-9}$ m]

[Hint:

KE = Quantum energy − Threshold energy

$$= \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{3000 \times 10^{-10}} - \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{4000 \times 10^{-10}}$$

$$= 6.626 \times 10^{-19} - 4.9695 \times 10^{-19}$$

$$= 1.6565 \times 10^{-19} \text{ joule}$$

$$\frac{1}{2} mv^2 = 1.6565 \times 10^{-19}$$

$$m^2v^2 = 2 \times 1.6565 \times 10^{-19} \times 9.1 \times 10^{-31}$$

$$mv = 5.49 \times 10^{-25}$$

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}}{5.49 \times 10^{-25}} = 1.2 \times 10^{-9} \text{ m}]$$

85. Show that de Broglie wavelength of electrons accelerated V volt is very nearly given by:

$$\lambda \text{ (in Å)} = \left(\frac{150}{V}\right)^{1/2}$$

[Hint: $\lambda = \frac{h}{\sqrt{2eVm}}$

$$\lambda = \left[\frac{h^2}{2eVm} \times 10^{20}\right]^{1/2} \text{ Å}$$

$$= \left[\frac{(6.626 \times 10^{-34})^2 \times 10^{20}}{2 \times 1.6 \times 10^{-19} \times V \times 9.1 \times 10^{-31}}\right]^{1/2} = \left[\frac{150}{V}\right]^{1/2}]$$

86. A 1 MeV proton is sent against a gold leaf ($Z = 79$). Calculate the distance of closest approach for head-on collision.

[Ans. 1.137×10^{-13} m]

[Hint: $d = \frac{Ze^2}{4\pi\varepsilon_0(\frac{1}{2}mv^2)}$. Do like Q.No. 83]

87. What is the energy, momentum and wavelength of the photon emitted by a hydrogen atom when an electron makes a transition from $n = 2$ to $n = 1$? Given that ionization potential is 13.6 eV.

[**Ans.** 16.32×10^{-19} J, 5.44×10^{-27} kg m / sec, 1218 Å]

[**Hint:** $E_1 = -13.6$ eV

$$E_2 = \frac{-13.6}{4} \text{ eV}$$

$$\Delta E = \frac{3}{4} \times 13.6 \text{ eV}$$

$$= 0.75 \times 13.6 \times 1.6 \times 10^{-19} \text{ J} = 1.632 \times 10^{-18} \text{ J}$$

$$\frac{hc}{\lambda} = 1.632 \times 10^{-18}$$

$$\lambda = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{1.632 \times 10^{-18}} = 1218 \times 10^{-10} \text{ m} = 1218 \text{ Å}$$

$$\lambda = \frac{h}{p}$$

$$\therefore \quad p = \frac{h}{\lambda} = \frac{6.626 \times 10^{-34}}{1218 \times 10^{-10}} = 5.44 \times 10^{-27} \text{ kg-m/sec}]$$

88. Calculate the orbital angular momentum of the following orbitals:
(a) $3p$ (b) $3d$ (c) $3s$
[**Ans.** (a) $\sqrt{2}\hbar$ (b) $\sqrt{6}\hbar$ (c) 0]
[**Hint:** (a) $\mu_l = \sqrt{l(l+1)}\,\frac{h}{2\pi}$ for $3p$, $l = 1 = \sqrt{2}\,\hbar$
(b) $\mu_l = \sqrt{6}\,\hbar$ for $3d$, $l = 2$
(c) $\mu_l = 0$ for $3s$, $l = 0$]

89. A single electron system has ionization energy 11180 kJ mol^{-1}. Find the number of protons in the nucleus of the system.
[**Ans.** $Z = 3$]

[**Hint:** $$\text{IE} = \frac{Z^2}{n^2} \times 21.69 \times 10^{-19} \text{ J}$$

$$\frac{11180 \times 10^3}{6.023 \times 10^{23}} = \frac{Z^2}{1^2} \times 21.69 \times 10^{-19}$$

$$Z \approx 3]$$

90. Suppose 10^{-17} J of light energy is needed by the interior of the human eye to see an object. How many photons of green light ($\lambda = 550$ nm) are needed to generate this minimum amount of energy?
[**Ans.** 28]

91. How many hydrogen atoms in the ground state are excited by means of monochromatic radiation of wavelength 970.6 Å. How many different lines are possible in the resulting emission spectrum? Find the longest wavelength among these.
[**Ans.** Six different lines, $\lambda = 1215.6$ Å]

[**Hint:** $$E_n - E_1 = \frac{hc}{\lambda}$$

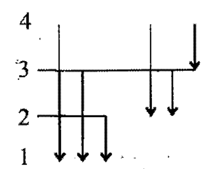
$$\frac{-21.69 \times 10^{-19}}{n^2} + \frac{21.69 \times 10^{-19}}{1} = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{970.6 \times 10^{-10}}$$

$n \approx 4$

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\frac{1}{\lambda} = 109677.77 \times 1^2\left(\frac{1}{1^2} - \frac{1}{4}\right)$$

$\lambda = 1215.68$ Å]



six different lines

OBJECTIVE QUESTIONS

Set-1: Questions with single correct answer

1. The ratio of e/m for a cathode ray:
   (a) varies with a gas in a discharge tube
   (b) is fixed
   (c) varies with different electrodes
   (d) is maximum if hydrogen is taken
2. Which of the following statements is wrong about cathode rays?
   (a) They travel in straight lines towards cathode
   (b) They produce heating effect
   (c) They carry negative charge
   (d) They produce X-rays when strike with material having high atomic masses
3. Cathode rays are:
   (a) electromagnetic waves (b) stream of α-particles
   (c) stream of electrons (d) radiations
4. Cathode rays have:
   (a) mass only (b) charge only
   (c) no mass and no charge (d) mass and charge both
5. Which is the correct statement about proton?
   (a) It is a nucleus of deuterium
   (b) It is an ionised hydrogen molecule
   (c) It is an ionised hydrogen atom
   (d) It is an α-particle
6. Neutron was discovered by:
   (a) J.J. Thomson (b) Chadwick
   (c) Rutherford (d) Priestley
7. The discovery of neutron came very late because:
   (a) it is present in nucleus
   (b) it is a fundamental particle
   (c) it does not move
   (d) it does not carry any charge
8. The fundamental particles present in equal numbers in neutral atoms (atomic number 71) are:
   (a) protons and electrons (b) neutrons and electrons
   (c) protons and neutrons (d) protons and positrons
9. The nucleus of the atom consists of:
   (a) protons and neutrons
   (b) protons and electrons
   (c) neutrons and electrons
   (d) protons, neutrons and electrons
10. The absolute value of charge on the electron was determined by:
    (a) J.J. Thomson (b) R.A. Millikan
    (c) Rutherford (d) Chadwick
11. Atomic number of an element represents: **(CBSE 1990)**
    (a) number of neutrons in the nucleus
    (b) atomic mass of an element
    (c) valency of an element
    (d) number of protons in the nucleus
12. Rutherford's experiment on scattering of α-particles showed for the first time that the atom has: **[CMC (Vellore) 1991]**
    (a) electrons (b) protons (c) neutrons (d) nucleus
13. Rutherford's scattering experiment is related to the size of the:
    (a) nucleus (b) atom (c) electron (d) neutron
14. When alpha particles are sent through a thin metal foil, most of them go straight through the foil because:
    (a) alpha particles are much heavier than electrons
    (b) alpha particles are positively charged
    (c) most part of the atom is empty space
    (d) alpha particles move with very high velocity
15. The radius of an atomic nucleus is of the order of: **[PMT (MP) 1991]**
    (a) 10^{-10} cm (b) 10^{-13} cm
    (c) 10^{-15} cm (d) 10^{-8} cm
16. Atomic size is of the order of:
    (a) 10^{-8} cm (b) 10^{-10} cm (c) 10^{-13} cm (d) 10^{-6} cm
17. Atoms may be regarded as comprising of protons, neutrons and electrons. If the mass attributed by electrons was doubled and that attributed by neutrons was halved, the atomic mass of ^{12}C would be:
    (a) approximately the same (b) doubled
    (c) reduced approx. 25% (d) approx. halved
18. Positive ions are formed from neutral atoms by the loss of:
    (a) neutrons (b) protons
    (c) nuclear charge (d) electrons
19. The nitrogen atom has 7 protons and 7 electrons. The nitride ion will have:
    (a) 10 protons and 7 electrons
    (b) 7 protons and 10 electrons
    (c) 4 protons and 7 electrons
    (d) 4 protons and 10 electrons
20. A light whose frequency is equal to 6×10^{14} Hz is incident on a metal whose work function is 2eV ($h = 6.63\times10^{-34}$ Js, $1\text{eV} = 1.6\times10^{-19}$ J). The maximum energy of electrons emitted will be: **(VITEEE 2008)**
    (a) 2.49 eV (b) 4.49 eV (c) 0.49 eV (d) 5.49 eV
    [**Hint :** Absorbed energy = Threshold energy + Kinetic energy of photoelectrons

    Absorbed energy = $h\nu$
    $$= 6.626\times10^{-34}\times6\times10^{14}$$
    $$= 3.9756\times10^{-19}\text{ J}$$
    $$= \frac{3.9756\times10^{-19}}{1.6\times10^{-19}} = 2.49\text{ eV}$$
    $$2.49 = 2\text{ eV} + \text{Kinetic energy of photoelectron}$$
    Kinetic energy of photoelectron = 0.49 eV]
21. The size of the nucleus is measured in:
    (a) amu (b) angstrom
    (c) cm (d) fermi
22. The highest value of e/m of anode rays has been observed when the discharge tube is filled with:
    (a) nitrogen (b) oxygen (c) hydrogen (d) helium

23. The particle with 13 protons and 10 electrons is:
(a) Al atom (b) Al^{3+} ion
(c) nitrogen isotope (d) none of these

24. Which of the following atoms contains the least number of neutrons?
(a) $^{235}_{92}U$ (b) $^{238}_{92}U$
(c) $^{239}_{93}Np$ (d) $^{240}_{93}Np$

25. The number of neutrons in dipositive zinc ion (Zn^{2+}, with mass number 70) is:
(a) 34 (b) 36 (c) 38 (d) 40

26. Which of the properties of the elements is a whole number?
(a) Atomic mass (b) Atomic number
(c) Atomic radius (d) Atomic volume

27. Increasing order (lowest first) for the values of e/m (charge/mass) for electron (e), proton (p), neutron (n) and α-particle (α) is:
(a) e, p, n, α (b) n, p, e, α
(c) n, p, α, e (d) n, α, p, e

28. The mass of neutron is of the order of:
(a) 10^{-27} kg (b) 10^{-26} kg
(c) 10^{-25} kg (d) 10^{-24} kg

29. The atoms of various isotopes of a particular element differ from each other in the number of:
(a) electrons in the outer shell only
(b) protons in the nucleus
(c) electrons in the inner shell only
(d) neutrons in the nucleus

30. Isotopes of the same element have:
(a) same number of neutrons
(b) same number of protons
(c) same atomic mass
(d) different chemical properties

31. Which of the following conditions is incorrect for a well behaved wave function (ψ)? **[EAMCET (Engg.) 2010]**
(a) ψ must be finite (b) ψ must be single valued
(c) ψ must be infinite (d) ψ must be continuous

32. Atomic mass of an element is not a whole number because:
(a) it contains electrons, protons and neutrons
(b) it contains isotopes
(c) it contains allotropes
(d) all of the above

33. Nucleons are:
(a) protons and neutrons
(b) neutrons and electrons
(c) protons and electrons
(d) protons, neutrons and electrons

34. Isotopes of an element have:
(a) different chemical and physical properties
(b) similar chemical and physical properties
(c) similar chemical but different physical properties
(d) similar physical and different chemical properties

35. Isotopes are identified by:
(a) positive ray analysis
(b) Astons' mass spectrograph
(c) Dempster's mass spectrograph
(d) all of the above

36. Mass spectrograph helps in the detection of isotopes because they:
(a) have different atomic masses
(b) have same number of electrons
(c) have same atomic number
(d) have same atomic masses

37. Which of the following statements is incorrect?
(a) The charge on an electron and proton are equal and opposite
(b) Neutrons have no charge
(c) Electrons and protons have the same mass
(d) The mass of a proton and a neutron are nearly the same

38. The charge on positron is equal to the charge on:
(a) proton (b) electron
(c) α-particle (d) neutron

39. Discovery of the nucleus of an atom was due to the experiment carried out by:
(a) Bohr (b) Rutherford
(c) Moseley (d) Thomson

40. Isobars are the atoms of: **(CBSE 1991)**
(a) same elements having same atomic number
(b) same elements having same atomic mass
(c) different elements having same atomic mass
(d) none of the above

41. Which of the following pairs represents isobars?
(a) $^{3}_{2}He$ and $^{4}_{2}He$ (b) $^{24}_{12}Mg$ and $^{25}_{12}Mg$
(c) $^{40}_{19}K$ and $^{40}_{20}Ca$ (d) $^{40}_{19}K$ and $^{39}_{19}K$

42. Na^+ ion is isoelectronic with: **(CPMT 1990)**
(a) Li^+ (b) Mg^{2+} (c) Ca^{2+} (d) Ba^{2+}

43. The triad of nuclei that is isotonic is:
(a) $^{14}_{6}C$, $^{14}_{7}N$, $^{19}_{9}F$ (b) $^{12}_{6}C$, $^{14}_{7}N$, $^{19}_{9}F$
(c) $^{14}_{6}C$, $^{14}_{7}N$, $^{17}_{9}F$ (d) $^{14}_{6}C$, $^{15}_{7}N$, $^{17}_{9}F$

44. Sodium atoms and sodium ions:
(a) are chemically similar
(b) both react vigorously with water
(c) have same number of electrons
(d) have same number of protons

45. In $^{35}_{17}Cl$ and $^{37}_{17}Cl$, which of the following is false?
(a) Both have 17 protons
(b) Both have 17 electrons
(c) Both have 18 neutrons
(d) Both show same chemical properties

46. Which of the following is isoelectronic with neon?
(a) O^{2-} (b) F^+ (c) Mg (d) Na

47. Neutrino has:
(a) charge +1, mass 1 (b) charge 0, mass 0
(c) charge −1, mass 1 (d) charge 0, mass 1

48. Positronium is the name given to an atom-like combination formed between: **(JIPMER 1991)**
(a) a positron and a proton
(b) a positron and a neutron

(c) a positron and an α-particle
(d) a positron and an electron

49. An isotone of $^{76}_{32}Ge$ is:
(a) $^{77}_{32}Ge$ (b) $^{78}_{33}As$ (c) $^{77}_{34}Se$ (d) $^{78}_{34}Se$

50. Which of the following does not characterise X-rays? **(IIT 1992)**
(a) The radiations can ionise gases
(b) It causes ZnS to fluorescence
(c) Deflected by electric and magnetic fields
(d) Have wavelengths shorter than ultraviolet rays

51. X-rays are produced when a stream of electrons in an X-ray tube:
(a) hits the glass wall of the tube
(b) strikes the metal target
(c) passes through a strong magnetic field
(d) none of the above

52. Radius of a nucleus is proportional to:
(a) A (b) $A^{1/3}$ (c) A^2 (d) $A^{2/3}$

53. The nature of positive rays produced in a vacuum discharge tube depends upon:
(a) the nature of the gas filled
(b) nature of the material of cathode
(c) nature of the material of anode
(d) the potential applied across the electrodes

54. Electromagnetic radiation with maximum wavelength is: **(MLNR 1991)**
(a) ultraviolet (b) radiowaves
(c) X-rays (d) infrared

55. The ratio of energy of radiations of wavelengths 2000 Å and 4000 Å is: **(CBSE 1994)**
(a) 2 (b) 4 (c) 1/2 (d) 1/4

56. The ratio of the diameter of the atom and the diameter of the nucleus is:
(a) 10^5 (b) 10^3 (c) 10 (d) 10^{-1}

57. The ratio of the volume of the atom and the volume of the nucleus is:
(a) 10^{10} (b) 10^{12} (c) 10^{15} (d) 10^{20}

58. Which of the following statements is incorrect?
(a) The frequency of radiation is inversely proportional to its wavelength
(b) Energy of radiation increases with increase in frequency
(c) Energy of radiation decreases with increase in wavelength
(d) The frequency of radiation is directly proportional to its wavelength

59. Visible light consists of rays with wavelengths in the approximate range of:
(a) 4000 Å to 7500 Å
(b) 4×10^{-3} cm to 7.5×10^{-4} cm
(c) 4000 nm to 7500 nm
(d) 4×10^{-5} m to 7.5×10^{-6} m

60. Which of the following statements concerning light is false?
(a) It is a part of the electromagnetic spectrum
(b) It travels with same velocity, *i.e.*, 3×10^{10} cm/s
(c) It cannot be deflected by a magnet
(d) It consists of photons of same energy

61. A 600 W mercury lamp emits monochromatic radiation of wavelength 331.3 nm. How many photons are emitted from the lamp per second? **[PET (Kerala) 2010]**
($h = 6.626 \times 10^{-34}$ Js, velocity of light $= 3 \times 10^8$ ms^{-1})
(a) 1×10^{19} (b) 1×10^{20}
(c) 1×10^{21} (d) 1×10^{23}
(e) 1×10^{22}

[**Hint :** Power $= \frac{\text{Energy}}{\text{Time}}$

$$600 = \frac{nhc}{\lambda \times 1 \text{ sec}}$$

$$600 = \frac{n \times 6.626 \times 10^{-34} \times 3 \times 10^8}{331.3 \times 10^{-9}}$$

$n = 1 \times 10^{21}$]

62. Out of X-rays, visible, ultraviolet, radiowaves, the largest frequency is of:
(a) X-rays (b) visible
(c) ultraviolet (d) radiowaves

63. The wave number which corresponds to electromagnetic radiations of 600 nm is equal to:
(a) 1.6×10^4 cm^{-1} (b) 0.16×10^4 cm^{-1}
(c) 16×10^4 cm^{-1} (d) 160×10^4 cm^{-1}

64. Line spectrum is characteristic of:
(a) molecules (b) atoms
(c) radicals (d) none of these

65. Which one of the following is not the characteristic of Planck's quantum theory of radiation? **(AIIMS 1991)**
(a) The energy is not absorbed or emitted in whole number multiple of quantum
(b) Radiation is associated with energy
(c) Radiation energy is not emitted or absorbed continuously but in the form of small packets called quanta
(d) This magnitude of energy associated with a quantum is proportional to the frequency

66. Which of the following among the visible colours has the minimum wavelength?
(a) Red (b) Blue
(c) Green (d) Violet

67. The spectrum of helium is expected to be similar to that of:
(a) H (b) Na (c) He^+ (d) Li^+

68. According to classical theory if an electron is moving in a circular orbit around the nucleus:
(a) it will continue to do so for sometime
(b) its orbit will continuously shrink
(c) its orbit will continuously enlarge
(d) it will continue to do so for all the time

69. Bohr advanced the idea of:
(a) stationary electrons (b) stationary nucleus
(c) stationary orbits (d) elliptical orbits

70. On Bohr stationary orbits:
(a) electrons do not move
(b) electrons move emitting radiations
(c) energy of the electron remains constant
(d) angular momentum of the electron is $\frac{h}{2\pi}$

71. Energy of Bohr orbit: **(DPMT 1991)**
(a) increases as we move away from the nucleus
(b) decreases as we move away from the nucleus
(c) remains the same as we move away from the nucleus
(d) none of the above

72. Which of the following statements does not form part of Bohr's model of the hydrogen atom?
(a) Energy of the electron in the orbit is quantized
(b) The electron in the orbit nearest to the nucleus has the lowest energy
(c) Electrons revolve in different orbit nucleus
(d) The position and velocity of the electron in the orbit cannot be determined simultaneously

73. Which of the following statements does not form a part of Bohr's model of hydrogen atom? **(DCE 2005)**
(a) Energy of the electrons in the orbit is quantised
(b) The electron in the orbit nearest to the nucleus has the lowest energy
(c) Electrons revolve in different orbits around the nucleus
(d) The position and velocity of the electrons in the orbit cannot be determined simultaneously

74. The radius of the first orbit of H-atom is r. Then the radius of the first orbit of Li^{2+} will be: **[AMU-PMT 2009]**
(a) $\frac{r}{9}$ (b) $\frac{r}{3}$ (c) $3r$ (d) $9r$

[**Hint :** $r = \frac{n^2}{z} \times 0.529\text{Å}$]

75. The energy liberated when an excited electron returns to its ground state can have:
(a) any value from zero to infinity
(b) only negative values
(c) only specified positive values
(d) none of the above

76. On the basis of Bohr's model, the radius of the 3rd orbit is:
(a) equal to the radius of first orbit
(b) three times the radius of first orbit
(c) five times the radius of first orbit
(d) nine times the radius of first orbit

77. The ratio of 2nd, 4th and 6th orbits of hydrogen atom is:
(a) 2 : 4 : 6 (b) 1 : 4 : 9
(c) 1 : 4 : 6 (d) 1 : 2 : 3

78. Which point does not pertain to Bohr's model of atom?
(a) Angular momentum is an integral multiple of $h/(2\pi)$
(b) The path of the electron is circular
(c) Force of attraction towards nucleus = centrifugal force
(d) The energy changes are taking place continuously

79. The distance between 3rd and 2nd orbits in the hydrogen atom is:
(a) 2.646×10^{-8} cm (b) 2.116×10^{-8} cm
(c) 1.058×10^{-8} cm (d) 0.529×10^{-8} cm

80. The correct expression derived for the energy of an electron in the nth energy level in hydrogen atom is:
(a) $E_n = \frac{2\pi^2 me^4}{n^2h^2}$ (b) $E_n = -\frac{2\pi^2 me^4}{nh^2}$
(c) $E_n = -\frac{2\pi me^2}{n^2h^2}$ (d) $E_n = -\frac{2\pi^2 me^4}{n^2h^2}$

81. According to Bohr theory, the angular momentum for an electron of 5th orbit is:
(a) $5h/\pi$ (b) $2.5h/\pi$ (c) $5\pi/h$ (d) $25h/\pi$

82. The value of Bohr radius of hydrogen atom is: **(CBSE 1991)**
(a) 0.529×10^{-7} cm (b) 0.529×10^{-8} cm
(c) 0.529×10^{-9} cm (d) 0.529×10^{-10} cm

83. The energy of an electron in the nth Bohr orbit of hydrogen atom is: **(CBSE 1992)**
(a) $-\frac{13.6}{n^4}$ eV (b) $-\frac{13.6}{n^3}$ eV (c) $-\frac{13.6}{n^2}$ eV (d) $-\frac{13.6}{n}$ eV

84. Which of the following electron transitions in hydrogen atom will require largest amount of energy? **(MLNR 1992)**
(a) from $n = 1$ to $n = 2$ (b) from $n = 2$ to $n = 3$
(c) from $n = \infty$ to $n = 1$ (d) from $n = 3$ to $n = 5$

85. For a hydrogen atom, the energies that an electron can have are given by the expression, $E = -13.58/n^2$ eV, where n is an integer. The smallest amount of energy that a hydrogen atom in the ground state can absorb is:
(a) 1.00 eV (b) 3.39 eV (c) 6.79 eV (d) 10.19 eV

86. The energy of hydrogen atom in its ground state is −13.6 eV. The energy of the level corresponding to $n = 5$ is: **(CBSE 1990)**
(a) − 0.54 eV (b) − 5.40 eV (c) − 0.85 eV (d) − 2.72 eV

87. $E_n = -313.6/n^2$ kcal/mol. If the value of $E = -34.84$ kcal/mol, to which value does 'n' correspond?
(a) 4 (b) 3 (c) 2 (d) 1

88. The ratio of the difference between 1st and 2nd Bohr orbits energy to that between 2nd and 3rd orbits energy is:
(a) 1/2 (b) 1/3 (c) 27/5 (d) 5/27

89. Bohr's model can explain:
(a) spectrum of hydrogen atom only
(b) spectrum of any atom or ion having one electron only
(c) spectrum of hydrogen molecule
(d) solar spectrum

90. The energy difference between two electronic states is 43.56 kcal/mol. The frequency of light emitted when the electron drops from higher orbit to lower orbit, is:
(Planck's constant = 9.52×10^{-14} kcal/mol)
(a) 9.14×10^{14} cycle/sec (b) 45.7×10^{14} cycle/sec
(c) 91.4×10^{14} cycle/sec (d) 4.57×10^{14} cycle/sec

91. Which of the following transitions of an electron in hydrogen atom emits radiation of the lowest wavelength? **[EAMCET (Engg.) 2010]**
(a) $n_2 = \infty$ to $n_1 = 2$ (b) $n_2 = 4$ to $n_1 = 3$
(c) $n_2 = 2$ to $n_1 = 1$ (d) $n_2 = 5$ to $n_1 = 3$

92. The wavelength of a spectral line for an electronic transition is inversely related to:
(a) number of electrons undergoing transition
(b) the nuclear charge of the atom
(c) the velocity of an electron undergoing transition
(d) the difference in the energy levels involved in the transition

93. The ionisation energy of the electron in the 1s-orbital of the hydrogen atom is 13.6 eV. The energy of the electron after promotion to 2s-orbital is: **[ISC (Bihar) 1993]**
(a) −3.4 eV (b) −13.6 eV
(c) −27.2 eV (d) 0.0 eV

94. Which electronic level would allow the hydrogen atom to absorb a photon but not to emit it?
(a) 1s (b) 2s (c) 3s (d) 4s

95. The spectral lines corresponding to the radiation emitted by an electron jumping from 6th, 5th and 4th orbits to second orbit belong to:
(a) Lyman series (b) Balmer series
(c) Paschen series (d) Pfund series

96. The spectral lines corresponding to the radiation emitted by an electron jumping from higher orbits to first orbit belong to:
(a) Paschen series (b) Balmer series
(c) Lyman series (d) None of these

97. In a hydrogen atom, the transition takes place from $n = 3$ to $n = 2$. If Rydberg constant is 1.097×10^7 m^{-1}, the wavelength of the emitted radiation is:
(a) 6564 Å (b) 6064 Å
(c) 6664 Å (d) 5664 Å
[**Hint:** Apply $\frac{1}{\lambda} = R\left[\frac{1}{x^2} - \frac{1}{y^2}\right]$]

98. The speed of the electron in the 1st orbit of the hydrogen atom in the ground state is (c is the velocity of light):
(a) $\frac{c}{1.37}$ (b) $\frac{c}{1370}$ (c) $\frac{c}{13.7}$ (d) $\frac{c}{137}$
[**Hint:** Velocity of electron in the 1st orbit, $v = h/(2\pi mr)$ $= 2.189 \times 10^8$ cm/sec.; velocity of light, $c = 3 \times 10^{10}$ cm/sec. Ratio $c/v = 137$]

99. Find the value of wave number $\bar{\nu}$ in terms of Rydberg's constant, when transition of electron takes place between two levels of He^+ ion whose sum is 4 and difference is 2.
(a) $\frac{8R}{9}$ (b) $\frac{32R}{9}$
(c) $\frac{3R}{4}$ (d) None of these
[**Hint :** $n_1 + n_2 = 4$, $n_2 - n_1 = 2$ $\therefore$ $n_1 = 1, n_2 = 3$

$$\bar{\nu} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
$$= R \times 2^2\left[\frac{1}{1} - \frac{1}{3^2}\right] = \frac{32R}{9}$$]

100. With the increasing principal quantum number, the energy difference between adjacent energy levels in hydrogen atom:
(a) increases (b) decreases
(c) is the same (d) none of these

101. An electron in an atom: **[CEET (Bihar) 1992]**
(a) moves randomly around the nucleus
(b) has fixed space around the nucleus
(c) is stationary in various energy levels
(d) moves around its nucleus in definite energy levels

102. The wave number of first line of Balmer series of hydrogen is 15200 cm^{-1}. The wave number of the first Balmer line of Li^{2+} ion is: **[IIT (Screening) 1992]**
(a) 15,200 cm^{-1} (b) 60,800 cm^{-1}
(c) 76,000 cm^{-1} (d) 1,36,800 cm^{-1}

103. "The position and the velocity of a small particle like electron cannot be simultaneously determined." This statement is:
(a) Heisenberg uncertainty principle
(b) Pauli's exclusion principle
(c) aufbau's principle
(d) de Broglie's wave nature of the electron

104. de Broglie equation describes the relationship of wavelength associated with the motion of an electron and its:
(a) mass (b) energy (c) momentum (d) charge

105. If the magnetic quantum number of a given atom is represented by −3, then what will be its principal quantum number? **[BHU (Pre.) 2005]**
(a) 2 (b) 3 (c) 4 (d) 5

106. Which of the following relates to photons both as wave motion and as a stream of particles? **(IIT 1992)**
(a) Interference (b) Diffraction
(c) $E = h\nu$ (d) $E = mc^2$

107. If uncertainty in the position of an electron is zero, the uncertainty in its momentum would be:
(a) zero (b) $< h/(4\pi)$
(c) $> h/(4\pi)$ (d) infinite

108. Which one of the following explains light both as a stream of particles and as wave motion?
(a) Diffraction (b) $\lambda = h/p$
(c) Interference (d) Photoelectric effect

109. A body of mass x kg is moving with velocity of 100 m sec^{-1}. Its de Broglie wavelength is 6.62×10^{-35} m. Hence x is:
($h = 6.62 \times 10^{-34}$ J sec) **[CET (Karnataka) 2009]**
(a) 0.25 kg (b) 0.15 kg
(c) 0.2 kg (d) 0.1 kg

110. A 200 g cricket ball is thrown with a speed of 3.0×10^3 cm sec^{-1}. What will be its de Broglie's wavelength?
($h = 6.6 \times 10^{-27}$ g cm^2 sec^{-1}) **[CET (Gujarat) 2008]**
(a) 1.1×10^{-32} cm (b) 2.2×10^{-32} cm
(c) 0.55×10^{-32} cm (d) 11.0×10^{-32} cm

111. The electronic configuration of a dipositive ion M^{2+} is 2, 8, 14 and its atomic mass is 56. The number of neutrons in the nucleus would be:
(a) 30 (b) 32 (c) 34 (d) 42

112. An element with atomic number 20 will be placed in which period of the periodic table?
(a) 5th (b) 4th (c) 3rd (d) 2nd

113. The frequency of radiation emitted when the electron falls from $n = 4$ to $n = 1$ in a hydrogen atom will be (Given ionisation energy of $H = 2.18 \times 10^{-18}$ J $atom^{-1}$ and $h = 6.626 \times 10^{-34}$ Js): **[Manipal (Med.) 2007]**
(a) 1.54×10^{15} s^{-1} (b) 1.03×10^{15} s^{-1}
(c) 3.08×10^{15} s^{-1} (d) 2×10^{15} s^{-1}

114. In a multi-electron atom, which of the following orbitals described by the three quantum numbers will have the same energy in the absence of magnetic and electric fields? **(AIEEE 2005)**
(i) $n = 1, l = 0, m = 0$ (ii) $n = 2, l = 0, m = 0$

(iii) $n = 2, l = 1, m = 1$ (iv) $n = 3, l = 2, m = 1$
(v) $n = 3, l = 2, m = 0$
(a) (i) and (ii) (b) (ii) and (iii)
(c) (iii) and (iv) (d) (iv) and (v)

115. Which of the ions is not having the configuration of Ne?
(a) Cl^- (b) F^- (c) Na^+ (d) Mg^{2+}

116. Which of the following has the maximum number of unpaired d-electrons? **(KCET 2008)**
(a) Ni^{3+} (b) Cu^+ (c) Zn^{2+} (d) Fe^{2+}

117. Which of the following expressions gives the de Broglie relationship? **[JEE (WB) 2008]**
(a) $p = \frac{h}{mv}$ (b) $\lambda = \frac{h}{mv}$
(c) $\lambda = \frac{h}{mp}$ (d) $\lambda m = \frac{h}{p}$

118. The principal quantum number of an atom is related to the: **(MLNR 1990)**
(a) size of the orbital
(b) orbital angular momentum
(c) spin angular momentum
(d) orientation of the orbital in space

119. The magnetic quantum is a number related to:
(a) size (b) shape
(c) orientation (d) spin

120. The principal quantum number represents: **(CPMT 1991)**
(a) shape of an orbital
(b) number of electrons in an orbit
(c) distance of electron from nucleus
(d) number of orbitals in an orbit

121. The quantum number not obtained from the Schrödinger's wave equation is: **(IIT 1990)**
(a) n (b) l (c) m (d) s

122. In a given atom, no two electrons can have the same values for all the four quantum numbers. This is called: **(CPMT 1990)**
(a) Hund's rule (b) Pauli's exclusion principle
(c) Uncertainty principle (d) aufbau principle

123. The atomic orbital is:
(a) the circular path of the electron
(b) elliptical shaped orbit
(c) three-dimensional field around nucleus
(d) the region in which there is maximum probability of finding an electron

124. If the ionization energy for hydrogen atom is 13.6 eV, then the ionization energy for He^+ ion should be: **[PMT (Haryana) 2004]**
(a) 13.6 eV (b) 6.8 eV
(c) 54.4 eV (d) 72.2 eV

125. Principal, azimuthal and magnetic quantum numbers are respectively related to:
(a) size, shape and orientation
(b) shape, size and orientation
(c) size, orientation and shape
(d) none of the above

126. Energy of electron in the H-atom is determined by :
(a) only n (b) both n and l
(c) n, l and m (d) all the four quantum numbers

127. Any p-orbital can accommodate up to: **(MLNR 1990)**
(a) 4 electrons
(b) 2 electrons with parallel spins
(c) 6 electrons
(d) 2 electrons with opposite spins

128. How many electrons can fit into the orbitals that comprise the 3rd quantum shell $n = 3$?
(a) 2 (b) 8 (c) 18 (d) 32

129. The total number of orbitals in a principal shell is:
(a) n (b) n^2 (c) $2n^2$ (d) $3n^2$

130. Two electrons in K-shell will differ in:
(a) principal quantum number
(b) spin quantum number
(c) azimuthal quantum number
(d) magnetic quantum number

131. Which one of the following orbitals has the shape of a baby-boother ?
(a) d_{xy} (b) $d_{x^2-y^2}$ (c) d_{z^2} (d) p_y

132. Which one of the following represents an impossible arrangement? **(AIEEE 2009)**

	n	l	m	s		n	l	m	s
(a)	3	2	−2	1/2	(b)	4	0	0	1/2
(c)	3	2	−3	1/2	(d)	5	3	0	1/2

133. Which of the following sets of quantum numbers is correct for an electron in $4f$-orbital? **(AIEEE 2004)**
(a) $n = 4, l = 3, m = +4, s = +1/2$
(b) $n = 4, l = 4, m = -4, s = -1/2$
(c) $n = 4, l = 3, m = +1, s = +1/2$
(d) $n = 3, l = 2, m = -2, s = +1/2$

134. The correct quantum numbers of $3p$-electrons are: **[PMT (Raj.) 2004]**
(a) $n = 3, l = 2, m = +2, s = +1/2$
(b) $n = 3, l = 1, m = -1, s = -1/2$
(c) $n = 3, l = -2, m = -2, s = +1/2$
(d) none of the above

135. In any subshell, the maximum number of electrons having same values of spin quantum number is :
(a) $\sqrt{l(l+1)}$ (b) $l + 2$
(c) $2l + 1$ (d) $4l + 2$
[**Hint :** Number of electrons with same spin
$= \frac{1}{2} \times$ Total no. of electrons
$= \frac{1}{2} \times 2(2l + 1) = (2l + 1)$]

136. Which of the following represents the correct set of four quantum numbers of a $4d$-electron? **(MLNR 1992)**
(a) 4, 3, 2, +1/2 (b) 4, 2, 1, 0
(c) 4, 3, −2, +1/2 (d) 4, 2, 1, −1/2

137. Values of magnetic orbital quantum number for an electron of M-shell can be : **[PET (Raj.) 2008]**
(a) 0, 1, 2 (b) −2, −1, 0, +1, +2
(c) 0, 1, 2, 3 (d) −1, 0, +1

138. Correct set of four quantum numbers for the outermost electron of rubidium ($Z = 37$) is:

(a) 5, 0, 0, 1/2 (b) 5, 1, 0, 1/2
(c) 5, 1, 1, 1/2 (d) 6, 0, 0, 1/2

139. Which one of the following subshells is spherical in shape?
(a) $4s$ (b) $4p$ (c) $4d$ (d) $4f$

140. In hydrogen atom, the electron is at a distance of 4.768 Å from the nucleus. The angular momentum of the electron is : **[EAMCET (Med.) 2010]**
(a) $\frac{3h}{2\pi}$ (b) $\frac{h}{2\pi}$
(c) $\frac{h}{\pi}$ (d) $\frac{3h}{\pi}$

[**Hint :** $r = \frac{n^2}{z} \times 0.529$ Å
$4.768 = \frac{n^2}{1} \times 0.529$
$n = 3$
$\therefore$ Angular momentum $(mvr) = \frac{nh}{2\pi} = \frac{3h}{2\pi}$]

141. Total number of m values for $n = 4$ is:
(a) 8 (b) 16 (c) 12 (d) 20

142. What is the total number of orbitals in the shell to which the g-subshell first arise?
(a) 9 (b) 16 (c) 25 (d) 36
[**Hint :** For g-subshell, $l = 4$
$\therefore$ It will arise in 5th shell.
Total number of orbitals in 5th shell = $n^2 = 25$]

143. In Bohr's model, if the atomic radius of the first orbit r_1, then radius of fourth orbit will be : **[BHU (Screening) 2010]**
(a) $4r_1$ (b) $6r_1$ (c) $16r_1$ (d) $\frac{r_1}{16}$

144. Which of the following statements is not correct for an electron that has quantum numbers $n = 4$ and $m = 2$? **(MLNR 1993)**
(a) The electron may have the q. no. $l = +1/2$
(b) The electron may have the q. no. $l = 2$
(c) The electron may have the q. no. $l = 3$
(d) The electron may have the q. no. $l = 0, 1, 2, 3$

145. The angular momentum of an electron depends on:
(a) principal quantum number
(b) azimuthal quantum number
(c) magnetic quantum number
(d) all of the above

146. The correct set of quantum numbers for the unpaired electron of a chlorine atom is: **(DPMT 2009)**
(a) $2, 0, 0, +\frac{1}{2}$ (b) $2, 1, -1, +\frac{1}{2}$
(c) $3, 1, -1, \pm\frac{1}{2}$ (d) $3, 0, 0, \pm\frac{1}{2}$

147. The magnetic quantum number for valency electron of sodium atom is:
(a) 3 (b) 2 (c) 1 (d) zero

148. The shape of the orbital is given by: **[PET (Raj.) 2008]**
(a) spin quantum number
(b) magnetic quantum number
(c) azimuthal quantum number
(d) principal quantum number

149. The energy of an electron of $2p_y$ orbital is:
(a) greater than $2p_x$ orbital
(b) less than $2p_z$ orbital
(c) equal to $2s$ orbital
(d) same as that of $2p_x$ and $2p_z$ orbitals

150. The two electrons occupying the same orbital are distinguished by:
(a) principal quantum number
(b) azimuthal quantum number
(c) magnetic quantum number
(d) spin quantum number

151. The maximum number of electrons in a subshell is given by the expression: **(AIEEE 2009)**
(a) $4l + 2$ (b) $4l - 2$
(c) $2l + 1$ (d) $2n^2$

152. The electronic configuration of an atom/ion can be defined by which of the following?
(a) Aufbau principle
(b) Pauli's exclusion principle
(c) Hund's rule of maximum multiplicity
(d) All of the above

153. An electron has a spin quantum number $+1/2$ and a magnetic quantum number -1. It cannot be present in:
(a) d-orbital (b) f-orbital (c) s-orbital (d) p-orbital

154. The value of azimuthal quantum number for electrons present in $4p$-orbitals is:
(a) 1
(b) 2
(c) any value between 0 and 3 except 1
(d) zero

155. For the energy levels in an atom which one of the following statements is correct?
(a) The $4s$ sub-energy level is at a higher energy than the $3d$ sub-energy level
(b) The M-energy level can have maximum of 32 electrons
(c) The second principal energy level can have four orbitals and contain a maximum of 8 electrons
(d) The 5th main energy level can have maximum of 50 electrons

156. A new electron enters the orbital when:
(a) $(n + l)$ is minimum (b) $(n + l)$ is maximum
(c) $(n + m)$ is minimum (d) $(n + m)$ is maximum

157. For a given value of n (principal quantum number), the energy of different subshells can be arranged in the order of:
(a) $f > d > p > s$ (b) $s > p > d > f$
(c) $f > p > d > s$ (d) $s > f > p > d$

158. After filling the $4d$-orbitals, an electron will enter in:
(a) $4p$ (b) $4s$ (c) $5p$ (d) $4f$

159. According to Aufbau principle, the correct order of energy of $3d$, $4s$ and $4p$-orbitals is: **[CET (J&K) 2006]**
(a) $4p < 3d < 4s$ (b) $4s < 4p < 3d$
(c) $4s < 3d < 4p$ (d) $3d < 4s < 4p$

160. Number of p-electrons in bromine atom is: **[PMT (Haryana) 2004]**
(a) 12 (b) 15
(c) 7 (d) 17

161. [Ar] $3d^{10}4s^1$ electronic configuration belongs to: **[PET (MP) 2008]**
(a) Ti (b) Tl (c) Cu (d) V

162. How many unpaired electrons are there in Ni^{2+} ? ($Z = 28$)
(a) Zero (b) 8 (c) 2 (d) 4

163. The electronic configuration of chromium (Z = 24) is: **[PMT (MP) 1993; BHU (Pre.) 2005]**
(a) [Ne] $3s^2 3p^6 3d^4 4s^2$ (b) [Ne] $3s^2 3p^6 3d^5 4s^1$
(c) [Ne] $3s^2 3p^6 3d^1 4s^2$ (d) [Ne] $3s^2 3p^6 4s^2 4p^4$

164. The number of d-electrons in Fe^{2+} (At. No. 26) is not equal to that of the: **(MLNR 1993)**
(a) p-electrons in Ne (At. No. 10)
(b) s-electrons in Mg (At. No. 12)
(c) d-electrons in Fe atom
(d) p-electrons in Cl^- ion (At. No. 17)

165. If the electronic structure of oxygen atom is written as

$1s^2, 2s^2$ $\overbrace{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{}}^{\longleftarrow 2p \longrightarrow}$; it would violate: **[ISC (Bihar) 1993]**

(a) Hund's rule
(b) Pauli's exclusion principle
(c) both Hund's and Pauli's principles
(d) none of the above

166. The orbital diagram in which 'aufbau principle' is violated, is:

	2s	←— 2p —→		
(a)	↑↓	↑↓	↑	
(b)	↑	↑↓	↑	↑
(c)	↑↓	↑	↑	↑
(d)	↑↓	↑↓	↑↓	↑

167. The manganese ($Z = 25$) has the outer configuration:

	4s	←— 3p —→				
(a)	↑↓	↑↓	↑	↑	↑	↑
(b)	↑↓	↑	↑	↑	↑	↑
(c)	↑↓	↑↓	↑↓	↑		
(d)		↑↓	↑↓	↑	↑	↑

168. Which of the following elements is represented by the electronic configuration?

←— 2p —→
2s ↑↓ | ↑↓ | ↑
1s ↑↓
↑↓

(a) Nitrogen (b) Fluorine
(c) Oxygen (d) Neon

169. The radial probability distribution curve obtained for an orbital wave function (ψ) has 3 peaks and 2 radial nodes. The valence electron of which one of the following metals does this wave function (ψ) correspond to ? **[EAMCET (Med.) 2010]**
(a) Co (b) Li (c) K (d) Na
[**Hint :** $Na_{11} \longrightarrow 1s^2, 2s^2 2p^6, \underbrace{3s^1}_{\text{Valence electron}}$
Number of radial node $= n - l - 1$ ($n = 3$)
$= 3 - 0 - 1 = 2$]

170. Krypton (At. No. 36) has the electronic configuration [A] $4s^2 3d^{10} 4p^6$. The 37th electron will go into which one of the following sub-levels?
(a) $4f$ (b) $4d$
(c) $3p$ (d) $5s$

171. An ion which has 18 electrons in the outermost shell is: **(CBSE 1990)**
(a) K^+ (b) Cu^+ (c) Cs^+ (d) Th^{4+}

172. Which of the following has non-spherical shell of electrons? **(IIT 1993)**
(a) He (b) B (c) Be (d) Li

173. Which one of the following sets of quantum numbers is not possible for an electron in the ground state of an atom with atomic number 19? **[PET (Kerala) 2006; CET (Karnataka) 2009]**
(a) $n = 2, l = 0, m = 0$ (b) $n = 2, l = 1, m = 0$
(c) $n = 3, l = 1, m = -1$ (d) $n = 3, l = 2, m = \pm 2$
(e) $n = 4, l = 0, m = 0$

174. Helium nucleus is composed of two protons and two neutrons. If the atomic mass is 4.00388, how much energy is released when the nucleus is constituted?
(Mass of proton = 1.00757, Mass of neutron = 1.00893)
(a) 283 MeV (b) 28.3 MeV
(c) 2830 MeV (d) 2.83 MeV

175. Binding energy per nucleon of three nuclei A, B and C are 5.5, 8.5 and 7.5 respectively. Which one of the following nuclei is most stable?
(a) A (b) C
(c) B (d) Cannot be predicted

176. The mass of ${}_3^7Li$ is 0.042 less than the mass of 3 protons and 4 neutrons. The binding energy per nucleon in ${}_3^7Li$ is: **(BHU 1992)**
(a) 5.6 MeV (b) 56 MeV (c) 0.56 MeV (d) 560 MeV

177. Meson was discovered by:
(a) Powell (b) Seaborg
(c) Anderson (d) Yukawa

178. In most stable elements, the number of protons and neutrons are:
(a) odd-odd (b) even-even
(c) odd-even (d) even-odd

179. Nuclear particles responsible for holding all nucleons together are:
(a) electrons (b) neutrons
(c) positrons (d) mesons

180. The introduction of a neutron into the nuclear composition of an atom would lead to a change in: **(MLNR 1995)**

(a) its atomic mass
(b) its atomic number
(c) the chemical nature of the atom
(d) number of the electron also

181. Which of the following has highest orbital angular momentum?
(a) 4s (b) 4p (c) 4d (d) 4f

182. Which of the following has maximum number of unpaired electrons? **[PMT (Raj.) 2004; BHU (Pre.) 2005]**
(a) Fe^{3+} (b) Fe^{2+} (c) Co^{2+} (d) Co^{3+}

183. An electron is not deflected on passing through a certain region, because:
(a) there is no magnetic field in that region
(b) there is a magnetic field but velocity of the electron is parallel to the direction of magnetic field
(c) the electron is a chargeless particle
(d) none of the above

184. In Millikan's oil drop experiment, we make use of:
(a) Ohm's law (b) Ampere's law
(c) Stoke's law (d) Faraday's law

185. A strong argument for the particle nature of cathode rays is:
(a) they can propagate in vacuum
(b) they produce fluorescence
(c) they cast shadows
(d) they are deflected by electric and magnetic fields

186. As the speed of the electrons increases, the measured value of charge to mass ratio (in the relativistic units):
(a) increases
(b) remains unchanged
(c) decreases
(d) first increases and then decreases

187. Which of the following are true for cathode rays?
(a) It travels along a straight line
(b) It emits X-rays when strikes a metal
(c) It is an electromagnetic wave
(d) It is not deflected by magnetic field

188. Three isotopes of an element have mass numbers, M, $(M+1)$ and $(M+2)$. If the mean mass number is $(M+0.5)$ then which of the following ratios may be accepted for M, $(M+1)$, $(M+2)$ in that order?
(a) 1 : 1 : 1 (b) 4 : 1 : 1
(c) 3 : 2 : 1 (d) 2 : 1 : 1

189. The radii of two of the first four Bohr orbits of the hydrogen atom are in the ratio 1 : 4. The energy difference between them may be:
(a) either 12.09 eV or 3.4 eV (b) either 2.55 eV or 10.2 eV
(c) either 13.6 eV or 3.4 eV (d) either 3.4 eV or 0.85 eV

190. Photoelectric emission is observed from a surface for frequencies ν_1 and ν_2 of the incident radiation $(\nu_1 > \nu_2)$. If the maximum kinetic energies of the photoelectrons in the two cases are in the ratio 1 : k then the threshold frequency ν_0 is given by:
(a) $\dfrac{\nu_2 - \nu_1}{k-1}$ (b) $\dfrac{k\nu_1 - \nu_2}{k-1}$ (c) $\dfrac{k\nu_2 - \nu_1}{k-1}$ (d) $\dfrac{\nu_2 - \nu_1}{k}$

191. The number of waves made by a Bohr electron in an orbit of maximum magnetic quantum number +2 is:
(a) 3 (b) 4 (c) 2 (d) 1

192. A certain negative ion X^{2-} has in its nucleus 18 neutrons and 18 electrons in its extranuclear structure. What is the mass number of the most abundant isotope of X?
(a) 36 (b) 35.46 (c) 32 (d) 39

193. Which of the following statements is not correct?
(a) The shape of an atomic orbital depends on the azimuthal quantum number
(b) The orientation of an atomic orbital depends on the magnetic quantum number
(c) The energy of an electron in an atomic orbital of multielectron atom depends on the principal quantum number
(d) The number of degenerate atomic orbitals of one type depends on the values of azimuthal and magnetic quantum numbers

194. Gases begin to conduct electricity at low pressure because: **(CBSE 1994)**
(a) at low pressures gases turn to plasma
(b) colliding electrons can acquire higher kinetic energy due to increased mean free path leading to ionisation of atoms
(c) atoms break up into electrons and protons
(d) the electrons in atoms can move freely at low pressure

195. An electron of mass m and charge e, is accelerated from rest through a potential difference V in vacuum. Its final speed will be: **(CBSE 1994)**
(a) $\sqrt{(eV/m)}$ (b) $2eV/m$
(c) $\sqrt{(eV/2m)}$ (d) $\sqrt{(2eV/m)}$

196. The difference in angular momentum associated with the electron in the two successive orbits of hydrogen atom is:
(a) h/π (b) $h/2\pi$ (c) $h/2$ (d) $(n-1)h/2\pi$

197. Photoelectric effect can be explained by assuming that light:
(a) is a form of transverse waves
(b) is a form of longitudinal waves
(c) can be polarised
(d) consists of quanta

198. The photoelectric effect supports quantum nature of light because:
(a) there is a minimum frequency of light below which no photoelectrons are emitted
(b) the maximum kinetic energy of photoelectrons depends only on the frequency of light and not on its intensity
(c) even when metal surface is faintly illuminated the photoelectrons leave the surface immediately
(d) electric charge of photoelectrons is quantised

199. The mass of a proton at rest is: **(CBSE 1991)**
(a) zero (b) 1.67×10^{-35} kg
(c) one amu (d) 9×10^{-31} kg

200. Momentum of a photon of wavelength λ is: **(CBSE 1993)**
(a) h/λ (b) zero (c) $h\lambda/c^2$ (d) $h\lambda/c$

201. When X-rays pass through air they:
(a) produce light track in the air
(b) ionise the gas

(c) produce fumes in the air
(d) accelerate gas atoms

202. X-rays: **(CPMT 1991)**
(a) are deflected in a magnetic field
(b) are deflected in an electric field
(c) remain undeflected by both the fields
(d) are deflected in both the fields

203. Find the frequency of light that corresponds to photons of energy 5.0×10^{-5} erg: **(AIIMS 2010)**
(a) $7.5 \times 10^{-21} \text{sec}^{-1}$ (b) 7.5×10^{-21} sec
(c) 7.5×10^{21} sec^{-1} (d) 7.5×10^{21} sec

[**Hint :** $\nu = \frac{E}{h} = \frac{5 \times 10^{-5} \text{ erg}}{6.63 \times 10^{-27} \text{ erg sec}}$
$= 7.54 \times 10^{21} \text{ sec}^{-1}$]

204. The energy of an electron in the first Bohr orbit of H-atom is -13.6 eV. The possible energy value(s) of the excited state(s) for electrons in Bohr orbits of hydrogen is/are: **(IIT 1998)**
(a) -3.4 eV (b) -4.2 eV
(c) -6.8 eV (d) $+6.8$ eV

205. The electrons identified by quantum numbers n and l, (i) $n = 4$, $l = 1$ (ii) $n = 4, l = 0$ (iii) $n = 3, l = 2$ (iv) $n = 3, l = 1$ can be placed in order of increasing energy, from the lowest to highest as: **(IIT 1999)**
(a) (iv) < (ii) < (iii) < (i) (b) (ii) < (iv) < (i) < (iii)
(c) (i) < (iii) < (ii) < (iv) (d) (iii) < (i) < (iv) < (ii)

206. The wavelength of the radiation emitted when an electron falls from Bohr orbit 4 to 2 in hydrogen atom is: **(IIT 1999)**
(a) 243 nm (b) 972 nm
(c) 486 nm (d) 182 nm

207. The energy of the electron in the first orbit of He^+ is -871.6×10^{-20} J. The energy of the electron in the first orbit of hydrogen would be: **(IIT 1998)**
(a) -871.6×10^{-20} J (b) -435×10^{-20} J
(c) -217.9×10^{-20} J (d) -108.9×10^{-20} J

208. The wavelength associated with a golf ball weighing 200 g and moving with a speed of 5 m/h is of the order of: **(IIT 2000)**
(a) 10^{-10} m (b) 10^{-20} m (c) 10^{-30} m (d) 10^{-40} m

209. Who modified Bohr theory by introducing elliptical orbits for electron path? **(CBSE 1999)**
(a) Hund (b) Thomson
(c) Rutherford (d) Sommerfeld

210. The uncertainty in momentum of an electron is 1×10^{-5} kg ms^{-1}. The uncertainty in its position will be:
($h = 6.62 \times 10^{-34}$ kg-m^2-s) **(CBSE 1999; BHU 2010)**
(a) 1.05×10^{-28} m (b) 1.05×10^{-26} m
(c) 5.27×10^{-30} m (d) 5.25×10^{-28} m

211. The Bohr orbit radius for the hydrogen atom ($n = 1$) is approximately 0.530 Å. The radius for the first excited state ($n = 2$) orbits is: **(CBSE 1998)**
(a) 0.13 Å (b) 1.06 Å (c) 4.77 Å (d) 2.12 Å

212. The number of nodal planes in p_x-orbital is: **(IIT 2000)**
(a) one (b) two (c) three (d) zero

213. The angular momentum (L) of an electron in a Bohr orbit is given as: **(IIT 1997)**
(a) $L = \frac{nh}{2\pi}$ (b) $L = \sqrt{l(l+1)}\frac{h}{2\pi}$
(c) $L = \frac{mg}{2\pi}$ (d) $L = \frac{h}{4\pi}$

214. Ground state electronic configuration of nitrogen atom can be represented by: **(IIT 1999)**

1. $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \uparrow \mid \uparrow}$ 2. $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \downarrow \mid \uparrow}$

3. $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \downarrow \mid \downarrow}$ 4. $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\downarrow \mid \downarrow \mid \downarrow}$

(a) 1 only (b) 1,2 only (c) 1,4 only (d) 2,3 only

215. Which of the following statement(s) are correct?
1. Electronic configuration of Cr is [Ar] $3d^5 4s^1$ (At. No. of Cr = 24)
2. The magnetic quantum number may have negative value
3. In silver atom, 23 electrons have a spin of one type and 24 of the opposite type (At. No. of Ag = 47)
4. The oxidation state of nitrogen in HN_3 is -3 **(IIT 1998)**

(a) 1,2,3 (b) 2,3,4 (c) 3,4 (d) 1,2,4

216. The electronic configuration of an element is $1s^2 2s^2 2p^6, 3s^2 3p^6 3d^5, 4s^1$. This represents: **(IIT 2000)**
(a) excited state (b) ground state
(c) cationic state (d) anionic state

217. The quantum numbers $+\frac{1}{2}$ and $-\frac{1}{2}$ for the electron spin represents: **(IIT 2000)**
(a) rotation of the electron in clockwise and anticlockwise directions respectively
(b) rotation of the electron in anticlockwise and clockwise directions respectively
(c) magnetic moment of electron pointing up and down respectively
(d) two quantum mechanical spin states which have no classical analogues

218. Rutherford's experiment, which established the nuclear model of the atom, used a beam of: **(IIT 2002)**
(a) β-particles, which impinged on a metal foil and got absorbed
(b) γ-rays, which impinged on a metal foil and ejected electrons
(c) helium atoms, which impinged on a metal foil and got scattered
(d) helium nuclei, which impinged on a metal foil and got scattered

219. How many moles of electrons weigh one kilogram? (Mass of electron $= 9.108 \times 10^{-31}$ kg, Avogadro's number $= 6.023 \times 10^{23}$) **(IIT 2002)**
(a) 6.023×10^{23} (b) $\frac{1}{9.108} \times 10^{31}$
(c) $\frac{6.023}{9.108} \times 10^{54}$ (d) $\frac{1}{9.108 \times 6.023} \times 10^{8}$

220. If the electronic configuration of nitrogen had $1s^7$, it would have energy lower than that of the normal ground state configuration $1s^2 2s^2 2p^3$ because the electrons would be closer to the nucleus. Yet $1s^7$ is not observed because it violates: **(IIT 2002)**

(a) Heisenberg uncertainty principle
(b) Hund's rule
(c) Pauli's exclusion principle
(d) Bohr postulates of stationary orbits

221. The orbital angular momentum of an electron in 2s-orbital is: **[IIT 1996; AIEEE 2003; PMT (MP) 2004]**

(a) $+\frac{1}{2}\frac{h}{2\pi}$ (b) zero
(c) $\frac{h}{2\pi}$ (d) $\sqrt{2}\,\frac{h}{2\pi}$

222. Calculate the wavelength (in nanometre) associated with a proton moving at 1×10^3 m sec^{-1}.
(mass of proton = 1.67×10^{-27} kg, $h = 6.63\times10^{-34}$ J sec) **(AIEEE 2009)**

(a) 0.032 nm (b) 0.40 nm
(c) 2.5 nm (d) 14 nm

[**Hint :** $\lambda = \frac{h}{mv} = \frac{6.63\times10^{-34}}{1.67\times10^{-27}\times10^{3}}$
$= 0.397\times10^{-9}$ m = 0.4 nm]

223. The value of Planck's constant is 6.63×10^{-34} J-s. The velocity of light is 3×10^8 m/sec. Which value is closest to the wavelength in nanometer of a quantum of light with frequency of 8×10^{15} sec^{-1}? **[CBSE (PMT) 2003]**

(a) 5×10^{-18} (b) 4×10^{1}
(c) 3×10^{7} (d) 2×10^{-25}

224. Which of the following statements in relation to the hydrogen atom is correct? **(AIEEE 2005)**

(a) 3s-orbital is lower in energy than 3p-orbital
(b) 3p-orbital is lower in energy than 3d-orbital
(c) 3s-and 3p-orbitals are of lower energy than 3d-orbital
(d) 3s, 3p-and 3d-orbitals all have the same energy

225. The number of d-electrons in Ni (At. No. = 28) is equal to that of the: **[CPMT (UP) 2004]**

(a) s and p-electrons in F^-
(b) p-electrons in Ar (At. No. = 18)
(c) d-electrons in Ni^{2+}
(d) total number of electrons in N (At. No. = 7)

226. The number of radial nodes of 3s- and 2p-orbitals are respectively: **[IIT (Screening) 2005]**

(a) 2, 0 (b) 0, 2 (c) 1, 2 (d) 2, 1

227. Which of the following is not permissible? **(DCE 2005)**

(a) $n = 4, l = 3, m = 0$ (b) $n = 4, l = 2, m = 1$
(c) $n = 4, l = 4, m = 1$ (d) $n = 4, l = 0, m = 0$

228. According to Bohr theory, the angular momentum of electron in 5th orbit is: **(AIEEE 2006)**

(a) $25\frac{h}{\pi}$ (b) $1\frac{h}{\pi}$ (c) $10\frac{h}{\pi}$ (d) $2.5\frac{h}{\pi}$

229. Which of the following sets of quantum numbers represents the highest energy of an atom? **(AIEEE 2007)**

(a) $n = 3, l = 0, m = 0, s = +\frac{1}{2}$
(b) $n = 3, l = 1, m = 1, s = +\frac{1}{2}$
(c) $n = 3, l = 2, m = 1, s = +\frac{1}{2}$
(d) $n = 4, l = 0, m = 0, s = +\frac{1}{2}$

230. In ground state, the radius of hydrogen atom is 0.53 Å. The radius of Li^{2+} ion ($Z = 3$) in the same state is: **[PET (Raj.) 2007]**

(a) 0.17 Å (b) 1.06 Å (c) 0.53 Å (d) 0.265 Å

231. How many d-electrons in Cu^+ (At. No. = 29) can have the spin quantum number $(-\frac{1}{2})$? **(SCRA 2007)**

(a) 3 (b) 7 (c) 5 (d) 9

232. Which of the following electronic configurations, an atom has the lowest ionisation enthalpy? **[CBSE (Med.) 2007]**

(a) $1s^2 2s^2 2p^3$ (b) $1s^2 2s^2 2p^6 3s^1$
(c) $1s^2 2s^2 2p^6$ (d) $1s^2 2s^2 2p^5$

233. The measurement of the electron position is associated with an uncertainty in momentum, which is equal to 1×10^{-18} g cm s^{-1}. The uncertainty in electron velocity is: (mass of an electron is 9×10^{-28} g) **[CBSE-PMT (Pre.) 2008]**

(a) 1×10^{5} cm s^{-1} (b) 1×10^{11} cm s^{-1}
(c) 1×10^{9} cm s^{-1} (d) 1×10^{6} cm s^{-1}

234. The ionization enthalpy of hydrogen atom is 1.312×10^6 J mol^{-1}. The energy required to excite the electron in the atom from $n = 1$ to $n = 2$ is : **(AIEEE 2008)**

(a) 9.84×10^5 J mol^{-1} (b) 8.51×10^5 J mol^{-1}
(c) 6.56×10^5 J mol^{-1} (d) 7.56×10^5 J mol^{-1}

[**Hint :** $E_1 = -1.312\times10^6$ J mol^{-1}

$$E_2 = \frac{E_1}{2^2} = -\frac{1.312\times10^6}{4}\text{ J mol}^{-1}$$

$$\Delta E = (E_2 - E_1) = 1.312\times10^6\left(1-\frac{1}{4}\right)$$

$$= \frac{3}{4}\times1.312\times10^6 = 9.84\times10^5\text{ J mol}^{-1}]$$

235. The wavelengths of electron waves in two orbits is 3 : 5. The ratio of kinetic energy of electrons will be: **(EAMCET 2009)**

(a) 25 : 9 (b) 5 : 3
(c) 9 : 25 (d) 3 : 5

[**Hint :** We know, $\lambda = \frac{h}{\sqrt{2Em}}$

$$\frac{\lambda_1}{\lambda_2} = \sqrt{\frac{E_2}{E_1}}$$

$$\frac{3}{5} = \sqrt{\frac{E_2}{E_1}}$$

$\therefore$ $E_1 : E_2 = 25 : 9$]

236. Electrons with a kinetic energy of 6.023×10^4 J/mol are evolved from the surface of a metal, when it is exposed to radiation of wavelength of 600 nm. The minimum amount of energy required to remove an electron from the metal atom is : **(EAMCET 2009)**

(a) 2.3125×10^{-19} J (b) 3×10^{-19} J
(c) 6.02×10^{-19} J (d) 6.62×10^{-34} J

[**Hint :** Absorbed energy = Threshold energy + kinetic energy of photoelectron

$$\frac{hc}{\lambda} = E_0 + KE$$

$$\frac{6.62\times10^{-34}\times3\times10^8}{600\times10^{-9}} = E_0 + \frac{6.023\times10^4}{6.023\times10^{23}}\text{ J/atom}$$

$$3.31\times10^{-19} = E_0 + 1\times10^{-19}$$
$$E_0 = 2.31\times10^{-19}\text{ J}]$$

237. For the Paschen series the value of n_1 and n_2 in the expression $\Delta E = R_H \times c\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$ is: **[JEE (WB) 2009]**

(a) $n_1 = 1, n_2 = 2, 3, 4 \ldots$
(b) $n_1 = 2, n_2 = 3, 4, 5 \ldots$
(c) $n_1 = 3, n_2 = 4, 5, 6 \ldots$
(d) $n_1 = 4, n_2 = 5, 6, 7 \ldots$

238. Ionization energy of He^+ is 19.6×10^{-18} J atom^{-1}. The energy of the first stationary state ($n = 1$) of Li^{2+} is: **(AIEEE 2010)**

(a) -2.2×10^{-15} J atom^{-1} (b) 8.82×10^{-17} J atom^{-1}
(c) 4.41×10^{-16} J atom^{-1} (d) -4.41×10^{-17} J atom^{-1}

[**Hint :** $\frac{I_{He^+}}{I_{Li^{2+}}} = \frac{Z_1^2}{Z_2^2}$

$$\frac{19.6\times10^{-18}}{I_{Li^{2+}}} = \frac{4}{9}$$
$$I_{Li^{2+}} = \frac{9}{4}\times19.6\times10^{-18}$$
$$= 44.1\times10^{-18}$$
$$= 4.41\times10^{-17}\text{ J atom}^{-1}$$
$$E_{Li^{2+}} = -4.41\times10^{-17}\text{ J atom}^{-1}]$$

239. The energy required to break one mole of Cl—Cl bonds in Cl_2 is 242 kJ mol^{-1}. The longest wavelength of light capable of breaking single Cl—Cl bond is: **(AIEEE 2010)**
($c = 3\times10^8$ m sec^{-1}, $N_A = 6.023\times10^{23}$ mol^{-1})

(a) 700 nm (b) 494 nm (c) 594 nm (d) 640 nm

[**Hint :** Bond energy of single bond $= \frac{242}{6.023\times10^{23}}$
$$= 4.017\times10^{-22}\text{ kJ}$$
$$= 4.017\times10^{-19}\text{ J}$$
$$E = \frac{hc}{\lambda}$$
$$4.017\times10^{-19} = \frac{6.626\times10^{-34}\times3\times10^8}{\lambda}$$
$$\lambda = 4.94\times10^{-7}\text{ m} = 494\text{ nm}]$$

240. In Sommerfeld's modification of Bohr's theory, the trajectory of an electron in a hydrogen atom is: **[JEE (WB) 2010]**

(a) perfect ellipse
(b) a closed ellipse like curve, narrower at the perihelion position and flatter at the aphelion position
(c) a closed loop on spherical surface
(d) a rosette

Set-2: The questions given below may have more than one correct answers

1. Correct order of radius of the 1st orbit of H, He^+, Li^{2+} and Be^{3+} is:

(a) H $> He^+ > Li^{2+} > Be^{3+}$
(b) $Be^{3+} > Li^{2+} > He^+ >$ H
(c) $He^+ > Be^{3+} > Li^{2+} >$ H
(d) $He^+ >$ H $> Li^{2+} > Be^{3+}$

2. Which is the correct relationship?

(a) E_1 of H $= 1/2\, E_2$ of $He^+ = 1/3\, E_3$ of $Li^{2+} = 1/4\, E_4$ of Be^{3+}
(b) $E_1(H) = E_2(He^+) = E_3(Li^{2+}) = E_4(Be^{3+})$
(c) $E_1(H) = 2E_2(He^+) = 3E_3(Li^{2+}) = 4E_4(Be^{3+})$
(d) No relation

3. Which is correct for any kind of species?

(a) $(E_2 - E_1) > (E_3 - E_2) > (E_4 - E_3)$
(b) $(E_2 - E_1) < (E_3 - E_2) < (E_4 - E_3)$
(c) $(E_2 - E_1) = (E_3 - E_2) = (E_4 - E_3)$
(d) $(E_2 - E_1) = 1/4(E_3 - E_2) = 1/9(E_4 - E_3)$

4. No. of visible lines when an electron returns from 5th orbit to ground state in H spectrum is:

(a) 5 (b) 4 (c) 3 (d) 10

5. Quantum numbers $l = 2$ and $m = 0$ represent which orbital?

(a) d_{xy} (b) $d_{x^2-y^2}$ (c) d_{z^2} (d) d_{zx}

6. If n and l are principal and azimuthal quantum numbers respectively, then the expression for calculating the total numbers of electrons in any energy level is:

(a) $\sum_{l=0}^{l=n} 2(2l+1)$ (b) $\sum_{l=1}^{l=n-1} 2(2l+1)$
(c) $\sum_{l=0}^{l=n+1} 2(2l+1)$ (d) $\sum_{l=0}^{l=n-1} 2(2l+1)$

7. Order of no. of revolution/sec $\gamma_1, \gamma_2, \gamma_3$ and γ_4 for I, II, III and IV orbits is:

(a) $\gamma_1 > \gamma_2 > \gamma_3 > \gamma_4$ (b) $\gamma_4 > \gamma_3 > \gamma_2 > \gamma_1$
(c) $\gamma_1 > \gamma_2 > \gamma_4 > \gamma_3$ (d) $\gamma_2 > \gamma_3 > \gamma_4 > \gamma_1$

8. Consider the following statements:

(A) Electron density in the xy-plane in $3d_{x^2-y^2}$ orbital is zero
(B) Electron density in the xy-plane in $3d_{z^2}$ orbital is zero
(C) $2s$-orbital has one nodal surface
(D) For $2p_z$-orbital yz is the nodal plane,

Which are the correct statements?

(a) (A) and (C) (b) (B) and (C)
(c) Only (B) (d) (A), (B), (C) and (D)

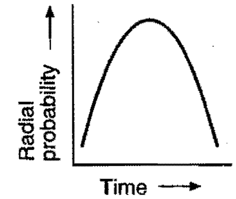
9. The first emission line in the H-atom spectrum in the Balmer series appears at:

(a) $\frac{5R}{36}$ cm^{-1} (b) $\frac{3R}{4}$ cm^{-1} (c) $\frac{7R}{144}$ cm^{-1} (d) $\frac{9R}{400}$ m^{-1}

10. 1 BM is equal to:

(a) $\frac{hc}{m\pi e^4}$ (b) $\frac{hc}{4\pi m}$ (c) $\frac{e^2hc}{4m}$ (d) $\frac{ehc}{\pi m}$

11. Radial probability distribution curve is shown for s-orbital. The curve is:



(a) $1s$
(b) $2s$
(c) $3s$
(d) $4s$

12. dz^2 orbital has:

(a) a lobe along z-axis and a ring along xy-plane
(b) a lobe along z-axis and a lobe along xy-plane
(c) a lobe along z-axis and a ring along yz-plane
(d) a lobe and ring along z-axis

13. When a light of frequency ν_1 is incident on a metal surface the photoelectrons emitted have twice the kinetic energy as did the photoelectron emitted when the same metal has irradiated with light of frequency ν_2. What will be the value of threshold frequency?
(a) $\nu_0 = \nu_1 - \nu_2$ (b) $\nu_0 = \nu_1 - 2\nu_2$
(c) $\nu_0 = 2\nu_1 - \nu_2$ (d) $\nu_0 = \nu_1 + \nu_2$

14. Heisenberg's uncertainty principle is not valid for:
(a) moving electrons (b) motor car
(c) stationary particles (d) all of these

15. Consider these electronic configurations for neutral atoms;
(i) $1s^2 2s^2 2p^6 3s^1$ (ii) $1s^2 2s^2 2p^6 4s^1$
Which of the following statements is/are false?
(a) Energy is required to change (i) to (ii)
(b) (i) represents 'Na' atom
(c) (i) and (ii) represent different elements
(d) More energy is required to remove one electron from (i) than (ii)

16. For the energy levels in an atom which one of the following statements is/are correct?
(a) There are seven principal electron energy levels
(b) The second principal energy level can have 4 subenergy levels and contain a maximum of 8 electrons
(c) The M energy level can have a maximum of 32 electrons
(d) The 4s subenergy level is at a lower energy than the 3d subenergy level

17. Which of the following statements are correct for an electron that has $n = 4$ and $m = -2$?
(a) The electron may be in a d-orbital
(b) The electron is in the fourth principal electronic shell
(c) The electron may be in a p-orbital
(d) The electron must have the spin quantum number $= +1/2$

18. The angular momentum of electron can have the value(s):
(a) $\frac{h}{2\pi}$ (b) $\frac{h}{\pi}$
(c) $\frac{2h}{\pi}$ (d) $\frac{5}{2}\frac{h}{2\pi}$

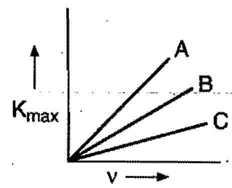
19. Which of the following statements is/are wrong?
(a) If the value of $l = 0$, the electron distribution is spherical
(b) The shape of the orbital is given by magnetic quantum no.
(c) Angular moment of 1s, 2s, 3s electrons are equal
(d) In an atom, all electrons travel with the same velocity

20. Consider the following sets of quantum numbers:

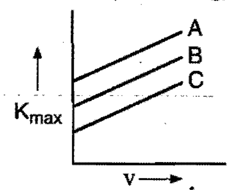
	n	l	m	s
(A)	3	0	0	$+\frac{1}{2}$
(B)	2	2	1	$+\frac{1}{2}$
(C)	4	3	-2	$-\frac{1}{2}$
(D)	1	0	-1	$-\frac{1}{2}$
(E)	3	2	3	$+\frac{1}{2}$

Which of the following sets of quantum numbers is not possible? **[CBSE (Med.) 2007]**
(a) (A), (B), (C) and (D) (b) (B), (D) and (E)
(c) (A) and (C) (d) (B), (C) and (D)

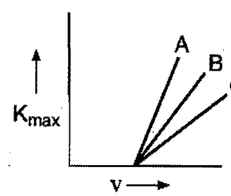
21. For three different metals A, B, C photo-emission is observed one by one. The graph of maximum kinetic energy versus frequency of incident radiation are sketched as : **[BHU (Screening) 2010]**



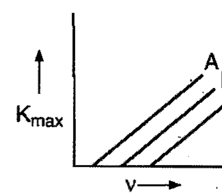
(a)



(b)



(c)



(d)

22. For which of the following species, the expression for the energy of electron in the n^{th} $\left[E_n = -\frac{13.6 Z^2}{n^2} \text{eV atom}^{-1}\right]$ has the validity? **[BHU (Mains) 2010]**
(a) Tritium (b) Li^{2+}
(c) Deuterium (d) He^{2+}

Assertion-Reason TYPE QUESTIONS

Set–1

The questions given below consist of an **'Assertion' (A)** and the **'Reason' (R).** Use the following keys for the appropriate answer:

(a) If both (A) and (R) are correct and (R) is the correct reason for (A).
(b) If both (A) and (R) are correct but (R) is not the correct explanation for (A).
(c) If (A) is correct but (R) is incorrect.
(d) If (A) is incorrect but (R) is correct.

1. (A) F-atom has less electron affinity than Cl^- atom.
(R) Additional electrons are repelled more effectively by $3p$ electrons in Cl atom than by $2p$ electrons in F-atom.

2. (A) Nuclide ${}^{30}_{13}Al$ is less stable than ${}^{40}_{20}Ca$.
(R) Nuclide having odd number of protons and neutrons are generally unstable. **(IIT 1998)**

3. (A) The first IE of Be is greater than that of B.
(R) $2p$-orbital is lower in energy than $2s$.

4. (A) The electronic configuration of nitrogen atom is represented as:

[⇅] [⇅] [↑ | ↑ | ↑]

and not as:

[⇅] [⇅] [⇅ | ↑ |]

(R) The electronic configuration of the ground state of an atom is the one which has the greatest multiplicity.

5. (A) The atomic radii of the elements of oxygen family are smaller than the atomic radii of corresponding elements of the nitrogen family.
(R) The members of oxygen family are all more electronegative and thus have lower value of nuclear charge than those of the nitrogen family.

6. (A) For $n = 3$, l may be 0, 1 and 2 and may be 0, ±1 and 0, ±1 and ±2.
(R) For each value of n, there are 0 to $(n - 1)$ possible values of l; for each value of l, there are 0 to $\pm l$ values of m.

7. (A) An orbital cannot have more than two electrons.
(R) The two electrons in an orbital create opposite magnetic field.

8. (A) The configuration of B-atom cannot be $1s^2 2s^2$.
(R) Hund's rule demands that the configuration should display maximum multiplicity.

9. (A) The ionization energy of N is more than that of O.
(R) Electronic configuration of N is more stable due to half-filled $2p$-orbitals.

10. (A) p-orbital is dumb-bell shaped.
(R) Electron present in p-orbital can have any one of the three values of magnetic quantum number, *i.e.*, 0, +1 or −1.

Set–2

The questions given below consist of two statements as **'Assertion' (A)** and **'Reason' (R);** while answering these choose any one of them:

(a) If (A) and (R) are both correct and (R) is the correct reason for (A).
(b) If (A) and (R) are both correct but (R) is not the correct reason for (A).
(c) If (A) is true but (R) is false.
(d) If both (A) and (R) are false.

11. (A) A special line will be seen for $2p_x - 2p_y$ transition.
(R) Energy is released in the form of wave of light when the electron drops from $2p_x$ to $-2p_y$ orbital. **(AIIMS 1996)**

12. (A) Ionization potential of Be (At. No. = 4) is less than B (At. No. = 5).
(R) The first electron released from Be is of p-orbital but that from B is of s-orbital. **(AIIMS 1997)**

13. (A) In Rutherford's gold foil experiment, very few α-particles are deflected back.
(R) Nucleus present inside the atom is heavy.

14. (A) Limiting line in the Balmer series has a wavelength of 364.4 mm.
(R) Limiting line is obtained for a jump of electron from $n = \infty$.

15. (A) Each electron in an atom has two spin quantum numbers.
(R) Spin quantum numbers are obtained by solving Schrödinger wave equation.

16. (A) There are two spherical nodes in $3s$-orbital.
(R) There is no planar node in $3s$-orbital.

17. (A) In an atom, the velocity of electron in the higher orbits keeps on decreasing.
(R) Velocity of electrons is inversely proportional to radius of the orbit.

18. (A) If the potential difference applied to an electron is made 4 times, the de Broglie wavelength associated is halved.
(R) On making potential difference 4 times, velocity is doubled and hence d is halved.

19. (A) Angular momentum of $1s$, $2s$, $3s$, etc., all have spherical shape.
(R) $1s$, $2s$, $3s$, etc., all have spherical shape.

20. (A) The radial probability of $1s$ electron first increases, till it is maximum at 53 Å and then decreases to zero.
(R) Bohr radius for the first orbit is 53 Å.

21. (A) On increasing the intensity of incident radiation, the number of photoelectrons ejected and their KE increases.
(R) Greater the intensity means greater the energy which in turn means greater the frequency of the radiation.

22. (A) A spectral line will be seen for a $2p_x - 2p_y$ transition.
(R) Energy is released in the form of wave of light when the electron drops from $2p_x$ to $2p_y$ orbital. **(VMMC 2007)**

Answers : OBJECTIVE QUESTIONS

- **Set–1**

1. (b)	2. (a)	3. (c)	4. (d)	5. (c)	6. (b)	7. (d)	8. (a)
9. (a)	10. (b)	11. (d)	12. (d)	13. (a)	14. (c)	15. (b)	16. (a)
17. (c)	18. (d)	19. (b)	20. (c)	21. (d)	22. (c)	23. (b)	24. (a)
25. (d)	26. (b)	27. (d)	28. (a)	29. (d)	30. (b)	31. (c)	32. (b)
33. (a)	34. (c)	35. (d)	36. (a)	37. (c)	38. (a)	39. (b)	40. (c)
41. (c)	42. (b)	43. (d)	44. (d)	45. (c)	46. (a)	47. (b)	48. (d)
49. (d)	50. (c)	51. (b)	52. (b)	53. (a)	54. (b)	55. (a)	56. (a)
57. (c)	58. (d)	59. (a)	60. (d)	61. (c)	62. (a)	63. (a)	64. (b)
65. (a)	66. (d)	67. (d)	68. (b)	69. (c)	70. (c)	71. (a)	72. (d)
73. (d)	74. (b)	75. (c)	76. (d)	77. (b)	78. (d)	79. (a)	80. (d)
81. (b)	82. (b)	83. (c)	84. (a)	85. (b)	86. (a)	87. (b)	88. (c)
89. (b)	90. (d)	91. (a)	92. (d)	93. (a)	94. (a)	95. (b)	96. (c)
97. (a)	98. (d)	99. (b)	100. (b)	101. (d)	102. (d)	103. (a)	104. (c)
105. (c)	106. (c)	107. (d)	108. (b)	109. (d)	110. (a)	111. (a)	112. (b)
113. (c)	114. (d)	115. (a)	116. (d)	117. (b)	118. (a)	119. (c)	120. (c)
121. (d)	122. (b)	123. (d)	124. (c)	125. (a)	126. (a)	127. (c)	128. (c)
129. (b)	130. (b)	131. (c)	132. (c)	133. (c)	134. (b)	135. (c)	136. (d)
137. (b)	138. (a)	139. (a)	140. (a)	141. (b)	142. (c)	143. (c)	144. (a)
145. (b)	146. (c)	147. (d)	148. (c)	149. (d)	150. (d)	151. (a)	152. (d)
153. (c)	154. (a)	155. (c)	156. (a)	157. (a)	158. (c)	159. (c)	160. (d)
161. (c)	162. (c)	163. (b)	164. (b)	165. (a)	166. (b)	167. (b)	168. (b)
169. (d)	170. (d)	171. (b)	172. (b)	173. (d)	174. (b)	175. (c)	176. (a)
177. (d)	178. (b)	179. (d)	180. (a)	181. (d)	182. (a)	183. (a, b)	184. (c)
185. (a)	186. (a)	187. (b)	188. (b)	189. (b)	190. (b)	191. (a)	192. (c)
193. (c)	194. (b)	195. (a)	196. (a)	197. (d)	198. (a)	199. (c)	200. (a)
201. (a)	202. (c)	203. (c)	204. (a)	205. (a)	206. (b)	207. (c)	208. (c)
209. (d)	210. (c)	211. (d)	212. (a)	213. (a)	214. (c)	215. (a)	216. (b)
217. (d)	218. (d)	219. (d)	220. (c)	221. (b)	222. (b)	223. (b)	224. (d)
225. (c)	226. (a)	227. (c)	228. (d)	229. (c)	230. (a)	231. (c)	232. (b)
233. (c)	234. (a)	235. (a)	236. (a)	237. (c)	238. (d)	239. (b)	240. (c)

- **Set–2**

1. (a)	2. (b)	3. (a)	4. (c)	5. (c)	6. (d)	7. (a)	8. (a)
9. (a)	10. (a)	11. (a)	12. (a)	13. (c)	14. (b, c)	15. (c, d)	16. (a, d)
17. (b, c)	18. (a, b, c)	19. (c)	20. (b)	21. (d)	22. (a, b, c)		

Answers : ASSERTION-REASON TYPE QUESTIONS

1. (c)	2. (a)	3. (c)	4. (a)	5. (c)	6. (a)	7. (b)	8. (c)
9. (c)	10. (a)	11. (d)	12. (d)	13. (b)	14. (a)	15. (d)	16. (b)
17. (c)	18. (a)	19. (b)	20. (b)	21. (d)	22. (d)		

BRAIN STORMING PROBLEMS

OBJECTIVE QUESTIONS *for* IIT ASPIRANTS

The following questions contain a single correct option:

1. The configuration of Cr atom is $3d^5 4s^1$ but not $3d^4 4s^2$ due to reason R_1 and the configuration of Cu atom is $3d^{10} 4s^1$ but not $3d^9 4s^2$ due to reason R_2, R_1 and R_2 are:

(a) R_1: The exchange energy of $3d^5 4s^1$ is greater than that of $3d^4 4s^2$.
R_2: The exchange energy of $3d^{10} 4s^1$ is greater than that of $3d^9 4s^2$.

(b) R_1: $3d^5 4s^1$ and $3d^4 4s^2$ have same exchange energy but $3d^5 4s^1$ is spherically symmetrical.
R_2: $3d^{10} 4s^1$ is also spherically symmetrical.

(c) R_1: $3d^5 4s^1$ has greater exchange energy than $3d^4 4s^2$.
R_2: $3d^{10} 4s^1$ has spherical symmetry.

(d) R_1: $3d^5 4s^1$ has greater energy than $3d^4 4s^2$.
R_2: $3d^{10} 4s^1$ has greater energy than $3d^9 4s^2$.

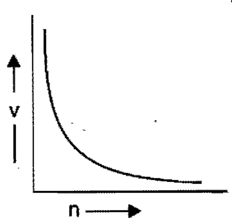
[**Hint:** $3d^5 4s^1$ is correct because it has greater exchange possibilities of unpaired electrons.

↑	↑	↑	↑	↑
1	2	3	4	5

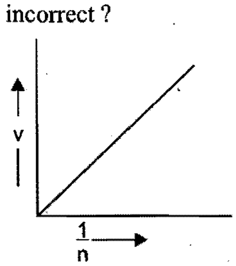
Exchange possibilities: (10) 1–2, 1–3, 1–4, 1–5
2–3, 2–4, 2–5
3–4, 3–5
4–5

$3d^{10} 4s^1$ is correct because $3d^{10}$-orbitals are spherically symmetrical.]

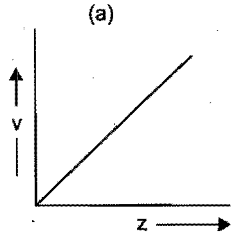
2. Which of the following graphs is incorrect?



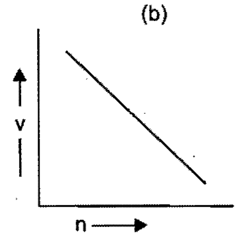
(a)



(b)



(c)



(d)

3. Which among the following is correct of $_5B$ in normal state?

	2s	2p	
(a)	[↑↑]	[| ↑ |]	: Against Hund's rule
(b)	[↑]	[↑ | ↑ |]	: Against aufbau principle as well as Hund's rule
(c)	[↑↑]	[↑ | |]	: Violation of Pauli's exclusion principle and not Hund's rule
(d)	[↑↓]	[↑ | |]	: Against aufbau principle

4. Maximum value $(n + l + m)$ for unpaired electrons in second excited state of chlorine $_{17}Cl$ is:

(a) 28 (b) 25
(c) 20 (d) none of these

[**Hint:** Configuration in second excited state may be given as:

	3p			3d				
	↑	↑	↑	↑	↑			
n	3	3	3	3	3			
l	1	1	1	2	2			
m	-1	0	$+1$	$+2$	$+1$			

Total $n + l + m = 25$]

5. Which of the following is correctly matched?

(a) Momentum of H-atom when electrons return from $n = 2$ to $n = 1$: $\frac{3Rh}{4}$

(b) Momentum of photon : Independent of wavelength of light

(c) e/m ratio of anode rays : Independent of gas in the discharge tube

(d) Radius of nucleus : $(\text{Mass no.})^{1/2}$

[**Hint:** $\frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{2^2}\right] = \frac{3R}{4}$

$\lambda = \frac{h}{p}$

$p = \frac{h}{\lambda} = h \times \frac{3R}{4} = \frac{3Rh}{4}$]

6. In hydrogen spectrum, the third line from the red end corresponds to which one of the following inter-orbit jumps of the electrons from Bohr orbit of hydrogen?

(a) $4 \rightarrow 1$ (b) $2 \rightarrow 5$ (c) $3 \rightarrow 2$ (d) $5 \rightarrow 2$

7. In which of the following pairs is the probability of finding the electron in xy-plane zero for both orbitals?

(a) $3d_{yz}, 4d_{x^2-y^2}$ (b) $2p_z, d_{z^2}$
(c) $4d_{zx}, 3p_z$ (d) None of these

8. In which of the following orbital diagrams are both Pauli's exclusion principle and Hund's rule violated?

(a) [↑↓] [↑ | ↑↓ | ↑] (b) [↑↓] [↑ | ↑↓ |]

(c) [↑↓] [↑↓ | ↑↑ | ↓] (d) [↓↓] [↑↓ | ↑ |]

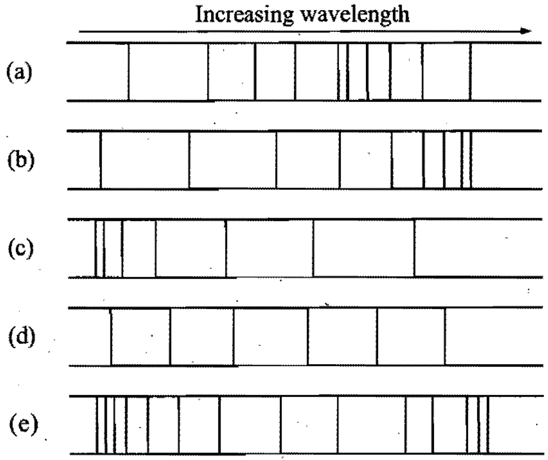
9. The distance between 3rd and 2nd Bohr orbits of hydrogen atom is:

(a) 0.529×10^{-8} cm (b) 2.645×10^{-8} cm
(c) 2.116×10^{-8} cm (d) 1.058×10^{-8} cm

[**Hint:** $r_3 - r_2 = (3^2 - 2^2) \times 0.529 \times 10^{-8}$ cm]

10. Which diagram represents the best appearance of the line spectrum of atomic hydrogen in the visible region?

[PET (Kerala) 2007]



11. The 'm' value for an electron in an atom is equal to the number of m values for $l = 1$. The electron may be present in:

(a) $3d_{x^2-y^2}$ (b) $5f_{x(x^2-y^2)}$
(c) $4f_{x^3/z}$ (d) none of these

[**Hint:** Total values of $m = (2l + 1) = 3$ for $l = 1$

$m = 3$ is for f-subshell orbitals.]

12. If m = magnetic quantum number, l = azimuthal quantum number, then:

(a) $m = l + 2$ (b) $m = 2l^2 + 1$
(c) $l = \frac{m-1}{2}$ (d) $l = 2m + 1$

[**Hint:** Magnetic quantum number 'm' lies between $(-l, 0, +l)$; thus total possible values of 'm' will be $(2l + 1)$.

$$m = 2l + 1, \quad i.e., \quad l = \frac{m-1}{2}\ ]$$

13. What are the values of the orbital angular momentum of an electron in the orbitals 1s, 3s, 3d and 2p?

(a) $0, 0, \sqrt{6}\,\hbar, \sqrt{2}\,\hbar$ (b) $1, 1, \sqrt{4}\,\hbar, \sqrt{2}\,\hbar$
(c) $0, 1, \sqrt{6}\,\hbar, \sqrt{3}\,\hbar$ (d) $0, 0, \sqrt{20}\,\hbar, \sqrt{6}\,\hbar$

[**Hint:** Orbital angular momentum $= \sqrt{l(l+1)}\,\frac{h}{2\pi} = \sqrt{l(l+1)}\,\hbar$]

14. After np-orbitals are filled, the next orbital filled will be:

(a) $(n+1)s$ (b) $(n+2)p$ (c) $(n+1)d$ (d) $(n+2)s$

15. The ratio of $(E_2 - E_1)$ to $(E_4 - E_3)$ for the hydrogen atom is approximately equal to:

(a) 10 (b) 15 (c) 17 (d) 12

[**Hint:**

$$\frac{E_4 - E_3}{E_2 - E_1} = \frac{\left(-\frac{1}{16}\right) - \left(-\frac{1}{9}\right)}{\left(-\frac{1}{4}\right) - (-1)} = \frac{\frac{1}{9} - \frac{1}{16}}{\frac{3}{4}} = \frac{7}{144} \times \frac{4}{3} = \frac{7}{108} \approx \frac{1}{15}\ ]$$

16. Which of the following electronic configurations has zero spin multiplicity?

(a)

↑	↑	↑

(b)

↑	↑	↓

(c)

↑	↓	↓

(d)

↓	↓	↓

[**Hint:** Spin multiplicity $= (2\Sigma s + 1)$]

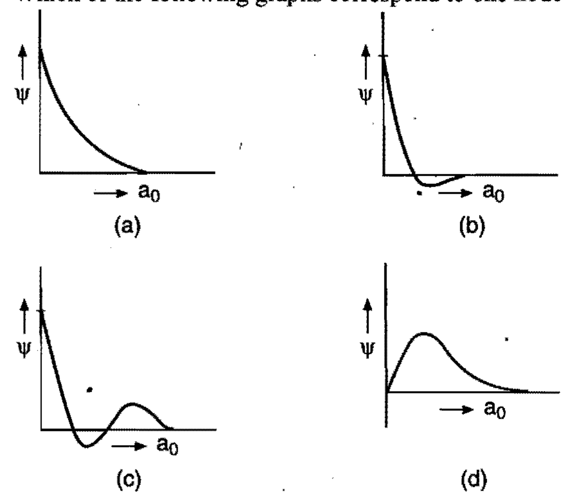
17. A photosensitive material would emit electrons if excited by photons beyond a threshold. To overcome the threshold, one would increase: **(VITEEE 2007)**

(a) the voltage applied to the light source
(b) the intensity of light
(c) the wavelength of light
(d) the frequency of light

18. Which of the following electronic configurations have the highest exchange energy?

	3d					4s
(a)	↑	↑	↑			↑
(b)	↑	↑	↑	↑	↑	↑
(c)	↑↓	↑↓	↑↓	↑↓	↑	↑
(d)	↑↓	↑↓	↑↓	↑↓	↑↓	↑

19. Which of the following graphs correspond to one node?



20. Angular distribution functions of all orbitals have:

(a) l nodal surfaces (b) $(l-1)$ nodal surfaces
(c) $(n+l)$ nodal surfaces (d) $(n-l-1)$ nodal surfaces

21. If uncertainty in position and momentum are equal then uncertainty in velocity is: **[CBSE-PMT (Pre.) 2008]**

(a) $\sqrt{\frac{h}{\pi}}$ (b) $\sqrt{\frac{h}{2\pi}}$ (c) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$ (d) $\frac{1}{m}\sqrt{\frac{h}{\pi}}$

[**Hint:** $\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$

$\Delta p \geq \sqrt{\frac{h}{4\pi}}$ when $\Delta x = \Delta p$

$m\Delta v \geq \sqrt{\frac{h}{4\pi}}$

$\Delta v \geq \frac{1}{2m}\sqrt{\frac{h}{\pi}}$]

22. The number of waves made by a Bohr electron in an orbit of maximum magnetic quantum number 3 is:

(a) 3 (b) 4 (c) 2 (d) 1

[**Hint:** $m = 3, l = 3, n = 4$

For, $n = 4$, number of waves will be 4.]

23. The number of elliptical orbits excluding circular orbits in the N-shell of an atom is:
(a) 3 (b) 4 (c) 2 (d) 1
[**Hint:** For, N-shell, $n = 4$. This shell will have one circular and three elliptical orbits.]

24. From the electronic configuration of the given elements K, L, M and N, which one has the highest ionization potential?
(a) $M = [Ne]\,3s^2 3p^2$ (b) $L = [Ne]\,3s^2 3p^3$
(c) $K = [Ne]\,3s^2 3p^1$ (d) $N = [Ar]\,3d^{10}, 4s^2 4p^3$
[**Hint:** L has half-filled p-subshell and it is smaller than N, hence, L will have the highest ionization potential.]

25. Which of the following pairs of electrons is excluded from an atom?
(a) $n = 2, l = 0, m = 0, s = +\frac{1}{2}$ and $n = 2, l = 0, m = 0, s = +\frac{1}{2}$
(b) $n = 2, l = 1, m = +1, s = +\frac{1}{2}$
and $n = 2, l = 1, m = -1, s = +\frac{1}{2}$
(c) $n = 1, l = 0, m = 0, s = +\frac{1}{2}$ and $n = 1, l = 0, m = 0, s = -\frac{1}{2}$
(d) $n = 3, l = 2, m = -2, s = +\frac{1}{2}$
and $n = 3, l = 0, m = 0, s = +\frac{1}{2}$
[**Hint:** Both $2s$ electrons have same spin, hence excluded from the atom.]

26. Given set of quantum numbers for a multielectron atom is:

n	l	m	s
2	0	0	+1/2
2	0	0	−1/2

What is the next higher allowed set of 'n' and 'l' quantum numbers for this atom in the ground state?
(a) $n = 2, l = 0$ (b) $n = 2, l = 1$
(c) $n = 3, l = 0$ (d) $n = 3, l = 1$

27. In how many elements does the last electron have the quantum numbers of $n = 4$ and $l = 1$?
(a) 4 (b) 6 (c) 8 (d) 10
[**Hint:** $n = 4, l = 1$ represent $4p$-subshell containing six electrons. Thus, there will be six elements having $4p^1$ to $4p^6$ electronic configuration.]

28. If there are three possible values $(-1/2, 0, +1/2)$ for the spin quantum, then electronic configuration of K (19) will be:
(a) $1s^3, 2s^3 2p^9, 3s^3 3p^1$ (b) $1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^1$
(c) $1s^2, 2s^2 2p^9, 3s^2 3p^4$ (d) none of these

29. If the radius of first Bohr orbit of hydrogen atom is 'x' then de Broglie wavelength of electron in 3rd orbit is nearly:
(a) $2\pi x$ (b) $6\pi x$ (c) $9x$ (d) $\frac{x}{3}$
[**Hint:**
$$r_n = n^2 r_1$$
$$r_3 = 9r_1 = 9x$$
$$mvr = n\frac{h}{2\pi}$$
$$mv9x = 3\frac{h}{2\pi}$$
$$\frac{h}{mv} = 6\pi x$$
$$\lambda = 6\pi x]$$

30. How many times does light travel faster in vacuum than an electron in Bohr first orbit of hydrogen atom?
(a) 13.7 times (b) 67 times (c) 137 times (d) 97 times
[**Hint:**
$$v = \frac{Z}{n} \times 2.188 \times 10^8 \text{ cm/sec}$$
$$v_1 = \frac{1}{1} \times 2.188 \times 10^8 \text{ cm/sec}$$
$$\frac{\text{Velocity of light}}{\text{Velocity of electron}} = \frac{3 \times 10^{10}}{2.188 \times 10^8} = 137 \text{ times}]$$

31. A compound of vanadium has a magnetic moment of 1.73 BM. The electronic configuration of vanadium ion in the compound is:
(a) $[Ar]\,3d^2$ (b) $[Ar]3d^1 4s^0$ (c) $[Ar]\,3d^3$ (d) $[Ar]\,3d^0 4s^1$
[**Hint:** Magnetic moment $= \sqrt{n(n+2)}$ BM
$$1.73 = \sqrt{n(n+2)}$$
$$\sqrt{3} = \sqrt{n(n+2)}$$
$n = 1$ (number of unpaired electrons)
$V_{22} \rightarrow 3d^2 4s^2$
$V^{3+} \rightarrow 3d^1 4s^0$ (has one unpaired electron)]

32. The orbital angular momentum of an electron in p-orbital is: **[PET (Kerala) 2006]**
(a) zero (b) $\frac{h}{\sqrt{2}\pi}$ (c) $\frac{h}{2\pi}$ (d) $\frac{1}{2}\frac{h}{2\pi}$
(e) $\frac{h}{2\sqrt{2}\pi}$

33. When a hydrogen atom emits a photon of energy 12.1 eV, the orbital angular momentum changes by:
(a) 1.05×10^{-34} J sec (b) 2.11×10^{-34} J sec
(c) 3.16×10^{-34} J sec (d) 4.22×10^{-34} J sec
[**Hint:** Emission of photon of 12.1 eV corresponds to the transition from $n = 3$ to $n = 1$.
$\therefore$ Change in angular momentum
$$= (n_2 - n_1)\frac{h}{2\pi}$$
$$= (3 - 1)\frac{h}{2\pi} = \frac{h}{\pi}$$
$$= \frac{6.626 \times 10^{-34}}{3.14}$$
$$= 2.11 \times 10^{-34} \text{ J sec}]$$

34. The total energy of the electron of H-atom in the second quantum state is $-E_2$. The total energy of the He^+ atom in the third quantum state is:
(a) $-\left(\frac{3}{2}\right)E_2$ (b) $-\left(\frac{2}{3}\right)E_2$ (c) $-\left(\frac{4}{9}\right)E_2$ (d) $-\left(\frac{16}{9}\right)E_2$
[**Hint:** Energy of electrons in nth state
$$= -\frac{Z^2}{n^2} \times 13.6 \text{ eV}$$

$$E_2(\text{H}) = -\frac{13.6}{1} \text{ eV}$$

$$E_3 (\text{He}^+) = -\frac{13.6 \times 4}{9} \text{ eV}$$

$$\frac{E_2}{E_3} = \frac{9}{4} \quad \text{or} \quad E_3 = \frac{4}{9} E_2$$

For negative value of E_2, E_3 will also be negative.]

35. What is the ratio of the Rydberg constant for helium to hydrogen atom?
(a) 1/2 (b) 1/4 (c) 1/8 (d) 1/16

[**Hint:** $R = \frac{-2\pi^2 mZ^2 e^4}{ch^3}$

$$\frac{R_{He}}{R_H} = \frac{2 \times 2^2}{1 \times 1^2} = 8$$

$$\therefore \quad \frac{R_H}{R_{He}} = \frac{1}{8} \text{]}$$

36. If the kinetic energy of a particle is doubled, de Broglie wavelength becomes:
(a) 2 times (b) 4 times (c) $\sqrt{2}$ times (d) $\frac{1}{\sqrt{2}}$ times

[**Hint:** $\lambda = \frac{h}{\sqrt{2Em}}$, where, E = Kinetic energy of the particle

$$\therefore \quad \lambda_1 = \frac{h}{\sqrt{2Em}}; \quad \lambda_2 = \frac{h}{\sqrt{2 \times 2Em}}$$

$$\therefore \quad \frac{\lambda_1}{\lambda_2} = \sqrt{2}, \quad i.e., \ \lambda_2 = \frac{\lambda_1}{\sqrt{2}} \text{]}$$

37. Imagine an atom made up of a proton and a hypothetical particle of double the mass of the electron but having the same charge as the electron. Apply the Bohr's atomic model and consider all possible transitions of this hypothetical particle to the first excited level. The largest wavelength photon that will be emitted has wavelength λ (given in terms of the Rydberg constant R for the hydrogen atom) equal to:
(a) $\frac{9}{5R}$ (b) $\frac{36}{5R}$ (c) $\frac{18}{5R}$ (d) $\frac{4}{R}$

[**Hint:** Energy is related to mass:

$$E_n \propto m$$

The longest wavelength λ_{max} photon will correspond to the transition of particle from $n = 3$ to $n = 2$

$$\frac{1}{\lambda_{max}} = 2R\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$\lambda_{max} = \frac{18}{5R} \text{]}$$

38. What is ratio of time periods (T_1 / T_2) in second orbit of hydrogen atom to third orbit of He^+ ion?
(a) $\frac{8}{27}$ (b) $\frac{32}{27}$ (c) $\frac{27}{32}$ (d) $\frac{27}{8}$

[**Hint :** $T \propto \frac{n^3}{Z^2}$

$$\frac{T_1}{T_2} = \frac{n_1^3 \times Z_2^2}{Z_1^2 \times n_2^3} = \frac{2^3 \times 2^2}{1^2 \times 3^3} = \frac{32}{27} \text{]}$$

39. The de Broglie wavelength of an electron accelerated by an electric field of V volt is given by :
(a) $\lambda = \frac{1.23}{\sqrt{m}}$ (b) $\lambda = \frac{1.23m}{\sqrt{h}}$ (c) $\frac{1.23}{\sqrt{V}} nm$ (d) $\lambda = \frac{1.23}{V}$

40. An excited electron of H-atoms emits of photon of wavelength λ and returns in the ground state, the principal quantum number of excited state is given by :
(a) $\sqrt{\lambda R (\lambda R - 1)}$ (b) $\sqrt{\frac{\lambda R}{(\lambda R - 1)}}$
(c) $\frac{1}{\sqrt{\lambda R (\lambda R - 1)}}$ (d) $\sqrt{\frac{(\lambda R - 1)}{\lambda R}}$

[**Hint :** $\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = R\left[\frac{1}{1} - \frac{1}{n_2^2}\right]$

$$n_2 = \sqrt{\frac{\lambda R}{\lambda R - 1}} \text{]}$$

41. A dye absorbs a photon of wavelength λ and re-emits the same energy into two photons of wavelength λ_1 and λ_2 respectively. The wavelength λ is related to λ_1 and λ_2 as :
(a) $\lambda = \frac{\lambda_1 \lambda_2}{(\lambda_1 + \lambda_2)^2}$ (b) $\lambda = \frac{\lambda_1 + \lambda_2}{\lambda_1 \lambda_2}$
(c) $\lambda = \frac{\lambda_1 \lambda_2}{\lambda_1 + \lambda_2}$ (d) $\frac{\lambda_1^2 \lambda_2^2}{\lambda_1 + \lambda_2}$

42. The radii of maximum probability for $3s$, $3p$ and $3d$-electrons are in the order :
(a) $(r_{max})_{3s} > (r_{max})_{3p} > (r_{max})_{3d}$
(b) $(r_{max})_{3s} = (r_{max})_{3p} = (r_{max})_{3d}$
(c) $(r_{max})_{3d} > (r_{max})_{3p} > (r_{max})_{3s}$
(d) $(r_{max})_{3d} > (r_{max})_{3s} > (r_{max})_{3p}$

Following questions may have more than one correct options:

1. Select the correct relations on the basis of Bohr theory:
(a) velocity of electron $\propto \frac{1}{n}$ (b) frequency of revolution $\propto \frac{1}{n^3}$
(c) radius of orbit $\propto n^2 Z$ (d) force on electron $\propto \frac{1}{n^4}$

2. To which of the following species, the Bohr theory is not applicable?
(a) He (b) Li^{2+} (c) He^{2+} (d) H-atom

3. The magnitude of spin angular momentum of an electron is given by:
(a) $S = \sqrt{s(s+1)} \frac{h}{2\pi}$ (b) $S = s\frac{h}{2\pi}$
(c) $S = \frac{\sqrt{3}}{2} \times \frac{h}{2\pi}$ (d) $S = \pm \frac{1}{2} \times \frac{h}{2\pi}$

[**Hint:** Spin angular momentum $= \sqrt{s(s+1)} \frac{h}{2\pi}$

$$S = \sqrt{\frac{1}{2}\left(\frac{1}{2} + 1\right)} \frac{h}{2\pi} = \frac{\sqrt{3}}{2} \times \frac{h}{2\pi} \text{]}$$

4. Select the correct configurations among the following:
(a) Cr $(Z = 24)$: [Ar] $3d^5, 4s^1$
(b) Cu $(Z = 29)$: [Ar] $3d^{10}, 4s^1$
(c) Pd $(Z = 46)$: [Kr] $4d^{10}, 5s^0$
(d) Pt $(Z = 78)$: [Xe] $4d^{10} 4s^2$

5. Which among the following statements is/are correct?
   (a) ψ^2 represents the atomic orbitals
   (b) The number of peaks in radial distribution is $(n - l)$
   (c) Radial probability density $\rho_{nl}(r) = 4\pi r^2 R_{nl}^2(r)$
   (d) A node is a point in space where the wave function (ψ) has zero amplitude
6. Select the correct statement(s) among the following:
   (i) Total number of orbitals in a shell with principal quantum number 'n' is n^2
   (ii) Total number of subshells in the n th energy level is n
   (iii) The maximum number of electrons in a subshell is given by the expression $(4l + 2)$
   (iv) $m = l + 2$, where l and m are azimuthal and magnetic quantum numbers
   (a) (i), (iii) and (iv) are correct
   (b) (i), (ii) and (iii) are correct
   (c) (ii), (iii) and (iv) are correct
   (d) (i), (ii) and (iv) are correct
7. Which among the following are correct about angular momentum of electron?
   (a) $2\hbar$ (b) $1.5\frac{h}{\pi}$ (c) $2.5\hbar$ (d) $0.5\frac{h}{\pi}$
8. Which of the following is/are incorrect for Humphrey lines of hydrogen spectrum?
   (a) $n_2 = 7 \rightarrow n_1 = 2$ (b) $n_2 = 10 \rightarrow n_1 = 6$
   (c) $n_2 = 5 \rightarrow n_1 = 1$ (d) $n_2 = 11 \rightarrow n_1 = 3$
9. In the Bohr's model of the atom:
   (a) the radius of n th orbit is proportional to n^2
   (b) the total energy of the electron in the n th orbit is inversely proportional to 'n'
   (c) the angular momentum of the electron is integral multiple of $h/2\pi$
   (d) the magnitude of potential energy of an electron in an orbit is greater than kinetic energy
10. Which among the following series is obtained in both absorption and emission spectrums?
   (a) Lyman series (b) Balmer series
   (c) Paschen series (d) Brackett series
11. The maximum kinetic energy of photoelectrons is directly proportional to . . . of the incident radiation. The missing word can be:
   (a) intensity (b) wavelength
   (c) wave number (d) frequency
12. Rutherford's experiment established that:
   (a) inside the atom there is a heavy positive centre
   (b) nucleus contains protons and neutrons
   (c) most of the space in an atom is empty
   (d) size of nucleus is very small
13. Which of the following orbital(s) lie in the xy-plane?
   (a) $d_{x^2-y^2}$ (b) d_{xy} (c) d_{xz} (d) d_{yz}
14. In which of the following sets of orbitals, electrons have equal orbital angular momentum?
   (a) $1s$ and $2s$ (b) $2s$ and $2p$ (c) $2p$ and $3p$ (d) $3p$ and $3d$
15. Which of the following orbitals have no spherical nodes?
   (a) $1s$ (b) $2s$ (c) $2p$ (d) $3p$
16. For a shell of principal quantum number $n = 4$, there are:
   (a) 16 orbitals (b) 4 subshells
   (c) 32 electrons (maximum) (d) 4 electrons with $l = 3$
17. The isotopes contain the same number of:
   (a) neutrons (b) protons
   (c) protons + neutrons (d) electrons
18. Which of the following species has less number of protons than the number of neutrons?
   (a) ${}^{12}_{6}C$ (b) ${}^{19}_{9}F$ (c) ${}^{23}_{11}Na$ (d) ${}^{24}_{12}Mg$
19. The angular part of the wave function depends on the quantum numbers are:
   (a) n (b) l (c) m (d) s
20. Which of the following species are expected to have spectrum similar to hydrogen?
   (a) He^+ (b) He^{2+} (c) Li^{2+} (d) Li^+
21. Which of the following statements is/are correct regarding a hydrogen atom?
   (a) Kinetic energy of the electron is maximum in the first orbit
   (b) Potential energy of the electron is maximum in the first orbit
   (c) Radius of the second orbit is four times the radius of the first orbit
   (d) Various energy levels are equally spaced

Answers

• Single correct option

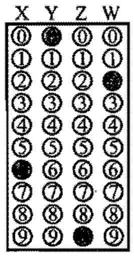
1. (c)	2. (d)	3. (c)	4. (b)	5. (a)	6. (d)	7. (d)	8. (d)
9. (b)	10. (c)	11. (b)	12. (c)	13. (a)	14. (a)	15. (b)	16. (c)
17. (d)	18. (b)	19. (b)	20. (a)	21. (c)	22. (b)	23. (a)	24. (b)
25. (a)	26. (b)	27. (b)	28. (a)	29. (b)	30. (c)	31. (b)	32. (b)
33. (b)	34. (c)	35. (c)	36. (d)	37. (c)	38. (b)	39. (c)	40. (b)
41. (c)	42. (a)						

• One or more than one correct options

1. (a, b, d)	2. (a, c)	3. (a, c)	4. (a, b, c)	5. (a, b, c, d)	6. (b)	7. (a, b, d)	8. (a, c, d)
9. (a, c, d)	10. (a)	11. (c, d)	12. (a, c, d)	13. (a, b)	14. (a, c)	15. (a, c)	16. (a, b, c)
17. (b, d)	18. (b, c)	19. (b, c)	20. (a, c)	21. (a, c)			

Integer Answer TYPE QUESTIONS

This section contains 10 questions. The answer to each of the questions is a single digit integer, ranging from 0 to 9. If the correct answers to question numbers X, Y, Z and W (say) are 6, 0, 9 and 2 respectively, then the correct darkening of bubbles will look like the figure :



1. For Li^{2+}, when an electron falls from a higher orbit to nth orbit, all the three types of lines, *i.e.*, Lyman, Balmer and Paschen was found in the spectrum. Here, the value of 'n' will be:
2. The emission lines of hydrogen contains ten lines. The highest orbit in which the electron is expected to be found is :

   [**Hint :** Number of lines $= \frac{n(n-1)}{2} = 10$

   $\therefore \qquad n = 5$]
3. Total number of nodes present in $4d$ orbitals will be :
4. Spin multiplicity of nitrogen in ground state will be :
5. Orbital frequency of electron in nth orbit of hydrogen is twice that of 2nd orbit. The value of n is :
6. If kinetic energy of an electron is reduce by (1/9) then how many times its de Broglie wavelength will increase.
7. If electrons in hydrogen sample return from 7th shell to 4th shell then how many maximum number of lines can be observed in the spectrum of hydrogen.
8. An electron in Li^{2+} ion is in excited state (n_2) . The wavelength corresponding to a transition to second orbit is 48.24 nm. From the same orbit, wavelength corresponding to a transition to third orbit is 142.46 nm. The value of n_2 is :
9. The energy corresponding to one of the lines in the Paschen series for H-atom is 18.16×10^{-20} J. Find the quantum numbers for the transition which produce this line.

   [**Hint :** $\quad \Delta E = 2.18 \times 10^{-18} \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

   $$18.16 \times 10^{-20} = 2.18 \times 10^{-18} \left[\frac{1}{9} - \frac{1}{n^2}\right]$$

   On solving, $n = 6$]
10. The angular momentum of electron in the shell in which the g-subshell first appears is $x \times \frac{h}{2\pi}$. The value of x will be :

    [**Hint :** $l = 4$ for g-subshell

    Thus, the subshell will first appear in $(n = l + 1 = 5)$ 5th shell.

    Angular momentum $(mvr) = n\frac{h}{2\pi}$

    $= 5\frac{h}{2\pi}$

    $\therefore \qquad n = 5$]

Answers

1. (1)	2. (5)	3. (3)	4. (4)	5. (1)	6. (3)	7. (6)	8. (5)
9. (6)	10. (5)						

LINKED COMPREHENSION TYPE QUESTIONS

• Passage 1

The observed wavelengths in the line spectrum of hydrogen atom were first expressed in terms of a series by Johann Jakob Balmer, a Swiss teacher.

Balmer's empirical formula is:

$$\frac{1}{\lambda}=R_H\left[\frac{1}{2^2}-\frac{1}{n^2}\right] n=3,4,5,\ldots$$

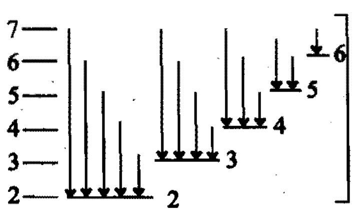
$R_H = 109678\ \text{cm}^{-1}$ *is the Rydberg constant.*

Niels Bohr derived this expression theoretically in 1913. *The formula is generalised to any one electron atom/ion.*

Answer the following questions:

1. Calculate the longest wavelength in Å ($1\ \text{Å} = 10^{-10}$ m) in the Balmer series of singly ionized helium He^+. Select the correct answer. Ignore the nuclear motion in your calculation.
   (a) 2651 Å (b) 1641.1 Å
   (c) 6569 Å (d) 3249 Å
   [**Hint:** $\frac{1}{\lambda_{He^+}}=R_H Z^2\left[\frac{1}{2^2}-\frac{1}{3^2}\right]$
   $=109678\times 4\left[\frac{5}{36}\right]$
   $\lambda_{He^+}=1641.1\ \text{Å}$]

2. How many lines in the spectrum will be observed when electrons return from 7th shell to 2nd shell?
   (a) 13 (b) 14 (c) 15 (d) 16
   [**Hint:** Number of lines in the spectrum
   $=\frac{(n_2-n_1)(n_2-n_1+1)}{2}$
   $=\frac{(7-2)(7-2+1)}{2}=15$



15 lines in the spectrum.]

3. The wavelength of first line of Balmer spectrum of hydrogen will be:
   (a) 4340 Å (b) 4101 Å (c) 6569 Å (d) 4861 Å
   [**Hint:** $\frac{1}{\lambda}=R_H\left[\frac{1}{2^2}-\frac{1}{n^2}\right]$
   for first line $n = 3$,
   $\therefore \frac{1}{\lambda}=109678\left[\frac{1}{2^2}-\frac{1}{3^2}\right]$
   $\lambda = 6569\ \text{Å}$]

4. In which region of electromagnetic spectrum does the Balmer series lie?
   (a) UV (b) Visible
   (c) Infrared (d) Far infrared

5. Which of the following is not correctly matched?
   (a) H_α— 6569 Å (Red) (b) H_β — 4861 Å (Blue)
   (c) H_γ— 4340 Å (Orange) (d) H_δ — 4101 Å (Violet)

• Passage 2

A formula analogous to the Rydberg formula applies to the series of spectral lines which arise from transitions from higher energy level to the lower energy level of hydrogen atom.

A muonic hydrogen atom is like a hydrogen atom in which the electron is replaced by a heavier particle, the 'muon'. The mass of the muon is about 207 *times the mass of an electron, while the charge remains same as that of the electron. Rydberg formula for hydrogen atom is:*

$$\frac{1}{\lambda}=R_H\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right](R_H = 109678\ \text{cm}^{-1})$$

Answer the following questions:

1. Radius of first Bohr orbit of muonic hydrogen atom is:
   (a) $\frac{0.259}{207}$ Å (b) $\frac{0.529}{207}$ Å
   (c) 0.529×207 Å (d) 0.259×207 Å

2. Energy of first Bohr orbit of muonic hydrogen atom is:
   (a) $-\frac{13.6}{207}$ eV (b) -13.6×207 eV
   (c) $+\frac{13.6}{207}$ eV (d) $+13.6\times 207$ eV

3. Ionization energy of muonic hydrogen atom is:
   (a) $+\frac{13.6}{207}$ eV (b) $+13.6\times 207$ eV
   (c) $-\frac{13.6}{207}$ eV (d) -13.6×207 eV

4. Angular momentum of 'muon' in muonic hydrogen atom may be given as:
   (a) $\frac{h}{\pi}$ (b) $\frac{h}{2\pi}$ (c) $\frac{h}{4\pi}$ (d) $\frac{h}{6\pi}$

5. Distance between first and third Bohr orbits of muonic hydrogen atom will be:
   (a) $\frac{0.529}{207}\times 2$ Å (b) $\frac{0.529}{207}\times 7$ Å
   (c) $\frac{0.529}{207}\times 8$ Å (d) $\frac{0.529}{207}$ Å

• Passage 3

Nuclei that have 2, 8, 20, 28, 50, 82 *and* 126 *neutrons or protons are more abundant and more stable than other nuclei of similar mass. It is suggested that in the nuclear structure of the numbers* 2, 8, 20, 28, 50, 82 *and* 126, *which have become known as magic numbers, the nuclei possessing magic numbers are spherical and have zero quadruple moment and hence they are highly stable. Nuclear shells are filled when there are* 2, 8, 20, 28, 50, 82 *and* 126 *neutrons or protons in a nucleus. In even-even nuclei all the neutrons and protons are paired and cancel out spin and orbital angular momenta.*

Answer the following questions regarding the stability of nucleus:

1. Which of the following element(s) is/are stable though having odd number of neutrons and protons?
   (a) ${}_3^6Li$ (b) ${}_5^{11}B$ (c) ${}_2^4He$ (d) ${}_7^{14}N$
2. Stable nuclei having number of neutrons less than number of protons are:
   (a) ${}_1^1H$ (b) ${}_2^3He$ (c) ${}_5^{11}B$ (d) ${}_6^{11}C$
3. Doubly magic nucleus is . . . .
   (a) ${}_{82}^{207}Pb$ (b) ${}_{82}^{206}Pb$ (c) ${}_{82}^{208}Pb$ (d) ${}_{83}^{209}Bi$
4. Which among the following has unstable nucleus?
   (a) ${}_7^{14}N$ (b) ${}_7^{15}N$ (c) ${}_7^{13}N$ (d) ${}_8^{16}O$
5. Which of the following has zero spin and angular momentum?
   (a) ${}_{20}^{40}Ca$ (b) ${}_1^3H$ (c) ${}_6^{14}C$ (d) ${}_{17}^{37}Cl$

• Passage 4

The substances which contain species with unpaired electrons in their orbitals behave as paramagnetic substances. Such substances are weakly attracted by the magnetic field. The paramagnetism is expressed in terms of magnetic moment. The magnetic moment is related to the number of unpaired electrons according to the following relation:

$$\text{Magnetic moment, } \mu = \sqrt{n(n+2)}\ BM$$

where, n = number of unpaired electrons.

BM stands for Bohr magneton, a unit of magnetic moment.

$$1\,BM = \frac{eh}{4\pi mc} = 9.27 \times 10^{-24}\ Am^2 \text{ or } J\,T^{-1}$$

Answer the following questions:

1. Which of the following ions has the highest magnetic moment?
   (a) Fe^{2+} (b) Mn^{2+} (c) Cr^{3+} (d) V^{3+}
2. Which of the following ions has magnetic moment equal to that of Ti^{3+}:
   (a) Cu^{2+} (b) Ni^{2+} (c) Co^{2+} (d) Fe^{2+}
3. An ion of a d-block element has magnetic moment 5.92 BM Select the ion among the following:
   (a) Zn^{2+} (b) Sc^{3+} (c) Mn^{2+} (d) Cr^{3+}
4. In which of these options do both constituents of the pair have the same magnetic moment?
   (a) Zn^{2+} and Cu^{+} (b) Co^{2+} and Ni^{2+}
   (c) Mn^{4+} and Co^{2+} (d) Mg^{2+} and Sc^{+}
5. Which of the following ions are diamagnetic?
   (a) He^{2+} (b) Sc^{3+} (c) Mg^{2+} (d) O^{2-}

• Passage 5

At the suggestion of Ernest Rutherford, Hans Geiger and Ernest Marsden bombarded a thin gold foil by α-particles from a polonium source. It was expected that α-particles would go right through the foil with hardly any deflection. Although, most of the alpha particles indeed were not deviated by much, a few were scattered through very large angles. Some were even scattered in the backward direction. The only way to explain the results, Rutherford found, was to picture an atom as being composed of a tiny nucleus in which its positive charge and nearly all its mass are concentrated. Scattering of α-particles is proportional to target thickness and is inversely proportional to the fourth power of sin $\frac{\theta}{2}$, where, θ is scattering angle. Distance of closest approach may be calculated as:

$$r_{min} = \frac{Z_1 Z_2 e^2}{4\pi\varepsilon_0 K}$$

where, K = kinetic energy of α-particles.

Answer the following questions:

1. Rutherford's α-particle scattering experiment led to the conclusion that:
   (a) mass and energy are related
   (b) mass and positive charge of an atom are concentrated in the nucleus
   (c) neutrons are present in the nucleus
   (d) atoms are electrically neutral
2. From the α-particle scattering experiment, Rutherford concluded that:
   (a) α-particles can approach within a distance of the order of 10^{-14} m of the nucleus
   (b) the radius of the nucleus is less than 10^{-14} m
   (c) scattering follows Coulomb's law
   (d) the positively charged parts of the atom move with extremely high velocities
3. Rutherford's scattering formula fails for very small scattering angles because:
   (a) the gold foil is very thin
   (b) the kinetic energy of α-particles is very high
   (c) the full nuclear charge of the target atom is partially screened by its electron
   (d) there is strong repulsive force between the α-particles and nucleus of the target
4. Alpha particles that come closer to the nuclei:
   (a) are deflected more (b) are deflected less
   (c) make more collision (d) are slowed down more
5. Which of the following quantities will be zero for alpha particles at the point of closest approach to the gold atom, in Rutherford's scattering of alpha particles?
   (a) Acceleration (b) Kinetic energy
   (c) Potential energy (d) Electrical energy

• Passage 6

The splitting of spectral lines by a magnetic field is called the Zeeman effect after the Dutch physicist Pieter Zeeman. The Zeeman effect is a vivid confirmation of space quantization. Magnetic quantum number 'm' was introduced during the study of Zeeman effect. 'm' can have the $(2l+1)$ values $(-l, 0, +l)$. Magnetic quantum number represents the orientation of atomic orbitals in three-dimensional space. The normal Zeeman effect consists of the splitting of a spectral line of frequency ν_0 into three components, i.e.,

$$\nu_1 = \nu_0 - \frac{e}{4\pi m}B;\ \nu_2 = \nu_0;\ \nu_3 = \nu_0 + \frac{e}{4\pi m}B$$

Here, B is magnetic field.

Answer the following questions:

1. Which of the following statements is incorrect with reference to the Zeeman effect?
   (a) In a magnetic field, the energy of a particular atomic state depends on the values of 'm' and 'n'
   (b) Zeeman effect is used to calculate the e/m ratio for an electron
   (c) Individual spectral lines split into separate lines. The distance between them is independent of the magnitude of the magnetic field
   (d) The Zeeman effect involves splitting of a spectral line of frequency ν_0 into three components
2. A d-subshell in an atom in the presence and absence of magnetic field is:
   (a) five-fold degenerate, non-degenerate
   (b) seven-fold degenerate, non-degenerate
   (c) five-fold degenerate, five-fold degenerate
   (d) non-degenerate, five-fold degenerate
3. Which among the following is/are correct about the orientation of atomic orbitals in space?
   (a) s-orbitals has single orientation
   (b) d-subshell orbitals have three orientations along x, y and z directions
   (c) f-subshell have seven orientations in their orbitals
   (d) None of the above
4. Zeeman effect explains splitting of spectral lines in:
   (a) magnetic field (b) electric field
   (c) both (a) and (b) (d) none of these
5. In presence of magnetic field, d-suborbit is:
   (a) five-fold degenerate (b) three-fold degenerate
   (c) seven-fold degenerate (d) non-degenerate

Passage 7

Spin angular momentum of an electron has no analogue in classical mechanics. However, it turns out that the treatment of spin angular momentum is closely analogous to the treatment of orbital angular momentum.

$$\text{Spin angular momentum} = \sqrt{s(s+1)}\,\hbar$$

$$\text{Orbital angular momentum} = \sqrt{l(l+1)}\,\hbar$$

Total spin of an atom or ion is a multiple of $\frac{1}{2}$. Spin multiplicity is a factor to confirm the electronic configuration of an atom or ion.

$$\text{Spin multiplicity} = (2\Sigma s + 1)$$

Answer the following questions:

1. Total spin of Mn^{2+} $(Z = 25)$ ion will be:
   (a) $\frac{3}{2}$ (b) $\frac{1}{2}$ (c) $\frac{5}{2}$ (d) $\frac{7}{2}$
2. Which of the following electronic configurations have four spin multiplicity?
   (a) ↑ ↑ ↑ (b) ↑ ↑ ↓
   (c) ↑ ↓ ↑ (d) ↓ ↓ ↓
3. Which of the following quantum numbers is not derived from Schrödinger wave equation?
   (a) Principal (b) Azimuthal
   (c) Magnetic (d) Spin
4. In any subshell, the maximum number of electrons having same value of spin quantum number is:
   (a) $\sqrt{l(l+1)}$ (b) $l+2$ (c) $2l+1$ (d) $4l+2$
5. The orbital angular momentum for a $2p$-electron is:
   (a) $\sqrt{3}\,\hbar$ (b) $\sqrt{6}\,\hbar$ (c) zero (d) $\sqrt{2}\,\frac{h}{2\pi}$

• Passage 8

Dual nature of matter was proposed by de Broglie in 1923, it was experimentally verified by Davisson and Germer by diffraction experiment. Wave character of matter has significance only for microscopic particles. de Broglie wavelength or wavelength of matter wave can be calculated using the following relation:

$$\lambda = \frac{h}{mv}$$

where, 'm' and 'v' are the mass and velocity of the particle.

de Broglie hypothesis suggested that electron waves were being diffracted by the target, much as X-rays are diffracted by planes of atoms in the crystals.

Answer the following questions:

1. Planck's constant has same dimension as that of:
   (a) work (b) energy
   (c) power (d) angular momentum
2. Wave nature of electrons is shown by:
   (a) photoelectric effect (b) Compton effect
   (c) diffraction experiment (d) Stark effect
3. de Broglie equation is obtained by combination of which of the following theories?
   (a) Planck's quantum theory
   (b) Einstein's theory of mass-energy equivalence
   (c) Theory of interference
   (d) Theory of diffraction
4. Which among the following is not used to calculate the de Broglie wavelength?
   (a) $\lambda = \frac{c}{\nu}$ (b) $\lambda = \frac{h}{mv}$
   (c) $\lambda = \frac{h}{\sqrt{2Em}}$ (d) $\lambda = \frac{h}{\sqrt{2qVm}}$
5. The wavelength of matter waves associated with a body of mass 1000 g moving with a velocity of 100 m/sec is:
   (a) 6.62×10^{-39} cm (b) 6.62×10^{-36} cm
   (c) 6.626×10^{-36} m (d) 3.31×10^{-32} m
6. An electron microscope is used to probe the atomic arrangements to a resolution of 5 Å. What should be the electric potential to which the electrons need to be accelerated? **(VITEEE 2008)**
   (a) 2.5 V (b) 6 V
   (c) 2.5 kV (d) 5 kV

• Passage 9

Orbital is the region in an atom where the probability of finding the electron is maximum. It represents three-dimensional motion of an electron around the nucleus. Orbitals do not specify a definite path according to the uncertainty principle. An orbital is described with the help of wave function ψ. Whenever an electron is described by a wave function, we say that an electron occupies that orbital. Since, many wave functions are possible for an electron, there are many atomic orbitals in an atom. Orbitals have different shapes; except s-orbitals, all other orbitals have directional character. Number of spherical nodes in an orbital is equal to $(n - l - 1)$. Orbital angular momentum of an electron is $\sqrt{l(l+1)}\,\hbar$.

Answer the following questions:

1. Which of the following orbitals is not cylindrically symmetrical about z-axis?
   (a) $3d_{z^2}$ (b) $4p_z$ (c) $6s$ (d) $3d_{yz}$
2. The nodes present in $5p$-orbital are:
   (a) one planar, five spherical (b) one planar, four spherical
   (c) one planar, three spherical (d) four spherical
3. When an atom is placed in a magnetic field, the possible number of orientations for an orbital of azimuthal quantum number 3 is:
   (a) three (b) one (c) five (d) seven
4. Orbital angular momentum of f-electrons is:
   (a) $\sqrt{2}\,\hbar$ (b) $\sqrt{3}\,\hbar$ (c) $\sqrt{12}\,\hbar$ (d) $2\,\hbar$
5. Which of the following orbitals has/have two nodal planes?
   (a) d_{xy} (b) d_{yz} (c) $d_{x^2-y^2}$ (d) All of these

• Passage 10

The hydrogen-like species Li^{2+} is in a spherically symmetric state S_1 with one radial node. Upon absorbing light the ion undergoes transition to a state S_2. The state S_2 has one radial node and its energy is equal to the ground state energy of the hydrogen atom.

(IIT 2010)

Answer the following questions:

1. The state S_1 is :
   (a) $1s$ (b) $2s$ (c) $2p$ (d) $3s$
   [**Hint:** $2s$ is symmetrical having one radial node.]
2. Energy of the state S_1 in units of the hydrogen atom ground state energy is:
   (a) 0.75 (b) 1.50 (c) 2.25 (d) 4.50
   [**Hint:** $\frac{E_{Li^{2+}}(2s)}{E_H} = \frac{-\frac{9}{4} \times 13.6}{-13.6} = 2.25$]
3. The orbital angular momentum quantum number of the state S_2 is:
   (a) 0 (b) 1 (c) 2 (d) 3
   [**Hint:** Orbital angular momentum quantum number of $3p$ subshell, *i.e.*, $l = 1$
   $$\underset{2s}{S_1} \xrightarrow{\text{Transition}} \underset{3p}{S_2}$$]

Answers

Passage 1.	1. (b)	2. (c)	3. (c)	4. (b)	5. (c)	
Passage 2.	1. (b)	2. (b)	3. (b)	4. (b)	5. (c)	
Passage 3.	1. (a, d)	2. (a, b)	3. (c)	4. (c)	5. (a)	
Passage 4.	1. (b)	2. (a)	3. (c)	4. (a, c)	5. (b, c, d)	
Passage 5.	1. (b)	2. (a, b, c)	3. (c, d)	4. (a)	5. (b)	
Passage 6.	1. (b)	2. (d)	3. (a, c)	4. (a)	5. (d)	
Passage 7.	1. (c)	2. (a)	3. (d)	4. (c)	5. (d)	
Passage 8.	1. (d)	2. (c)	3. (a, b)	4. (a)	5. (c)	6. (b)
Passage 9.	1. (d)	2. (c)	3. (d)	4. (c)	5. (d)	
Passage 10.	1. (b)	2. (c)	3. (b)			

SELF ASSESSMENT

ASSIGNMENT NO. 2

SECTION-I

Straight Objective Type Questions

This section contains 11 multiple choice questions. Each question has 4 choices (a), (b), (c) and (d), out of which only one is correct.

1. Which one of the following leads to third line of Balmer spectrum from red end (For hydrogen atom)?
(a) $2 \longrightarrow 5$ (b) $5 \longrightarrow 2$ (c) $3 \longrightarrow 2$ (d) $4 \longrightarrow 1$

2. The orbital angular momentum and angular momentum (classical analogue) for the electron of 4s-orbital are respectively, equal to:
(a) $\sqrt{12}\,\frac{h}{2\pi}$ and $\frac{h}{2\pi}$ (b) zero and $\frac{2h}{\pi}$
(c) $\sqrt{6}h$ and $\frac{2h}{\pi}$ (d) $\sqrt{2}\,\frac{h}{2\pi}$ and $\frac{3h}{2\pi}$

3. A sample of hydrogen atom is excited to $n = 4$ state. In the spectrum of emitted radiation, the number of lines in the ultraviolet and visible regions are respectively:
(a) 3 : 2 (b) 2 : 3 (c) 1 : 3 (d) 3 : 1

4. Number of de Broglie waves made by a Bohr electron in an orbit of maximum magnetic quantum number + 2 is:
(a) 1 (b) 2 (c) 3 (d) 4

5. First line of Lyman series of hydrogen atom occurs at $\lambda = x$ Å. The corresponding line of He^+ will occur at:
(a) $4x$ (b) $3x$ (c) $x/3$ (d) $x/4$

6. Electronic transition in He^+ ion takes from n_2 to n_1 shell such that;

$$2n_2 + 3n_1 = 18 \quad \text{...(i)}$$
$$2n_2 - 3n_1 = 6 \quad \text{...(ii)}$$

then what will be the total number of photons emitted when electrons transit to n_1 shell?
(a) 21 (b) 15 (c) 20 (d) 10

7. Which of the following sets of quantum numbers is not possible for an electron ? **[PET (Raj.) 2008]**
(a) $n = 1, l = 0, m_l = 0, m_s = -1/2$
(b) $n = 2, l = 1, m_l = 0, m_s = -1/2$
(c) $n = 1, l = 1, m_l = 0, m_s = +1/2$
(d) $n = 2, l = 1, m_l = 0, m_s = +1/2$

8. The average life of an excited state of hydrogen atom is of the order of 10^{-8} sec. The number of revolutions made by an electron when it returns from $n = 2$ to $n = 1$ is:
(a) 2.28×10^6 (b) 22.8×10^6 (c) 8.23×10^6 (d) 2.82×10^6

9. The wave number of a particular spectral line in the atomic spectrum of a hydrogen like species increases 9/4 times when deuterium nucleus is introduced into its nucleus, then which of the following will be the initial hydrogen like species?
(a) Li^{2+} (b) Li^+ (c) He^+ (d) Be^{3+}

10. Energy of electron in the first Bohr orbit of H-atom is -313.6 kcal mol^{-1}; then the energy in second Bohr orbit will be:
(a) $+313.6$ kcal mol^{-1} (b) -78.4 kcal mol^{-1}
(c) -34.84 kcal mol^{-1} (d) -12.5 kcal mol^{-1}

11. Which phenomenon best supports the theory that matter has a wave nature ? **(VITEEE 2008)**
(a) Electron momentum (b) Electron diffraction
(c) Photon momentum (d) Photon diffraction

SECTION-II

Multiple Answers Type Objective Questions

12. Which of the following is/are correct?
(a) An electron in excited state cannot absorb a photon
(b) Energy of electrons depends only on the principal quantum numbers
(c) Energy of electrons depends only on the principal quantum number for hydrogen atom
(d) Difference in potential energy of two shells is equal to the difference in kinetic energy of these shells

13. Which of the following statements is/are correct?
(a) Energy of $4s$, $4p$, $4d$ and $4f$ are same for hydrogen
(b) Angular momentum of electron $= I\omega$
(c) For all values of 'n', the p-orbitals have the same shape
(d) Orbital angular momentum $= nh/2\pi$

14. Which of the following orbitals are associated with angular nodes?
(a) f (b) d (c) p (d) s

15. The correct statement(s) among the following is/are:
(a) All d-orbitals except d_{z^2} have two angular nodes
(b) $d_{x^2-y^2}$, d_{z^2} lie on the axes
(c) The degeneracy of p-orbitals remains unaffected in the presence of external magnetic field
(d) d-orbitals have 3-fold degeneracy

SECTION-III

Assertion-Reason Type Questions

This section contains 5 questions. Each question contains **Statement-1** (Assertion) and **Statement-2** (Reason). Each question has following 4 choices (a), (b), (c) and (d), out of which only one is correct.

(a) Statement-1 is true; statement-2 is true; statement-2 is a correct explanation for statement-1.
(b) Statement-1 is true; statement-2 is true; statement-2 is not a correct explanation for statement-1.
(c) Statement-1 is true; statement-2 is false.
(d) Statement-1 is false; statement-2 is true.

16. **Statement-1:** Kinetic energy of photoelectrons increases with increase in the frequency of incident radiation.

Because

Statement-2: The number of photoelectrons ejected increases with increase in intensity of incident radiation.

17. **Statement-1:** Photoelectric effect is easily pronounced by caesium metal.

Because

Statement-2: Photoelectric effect is easily pronounced by the metals having high ionization energy.

18. **Statement-1:** Electrons in K-shell revolve in circular orbit.

Because

Statement-2: Principal quantum number 'n' is equal to 1 for the electrons in K-shell.

19. **Statement-1:** Orbit and orbital are synonymous.

Because

Statement-2: Orbit is the path around the nucleus in which electron revolves.

20. **Statement-1:** $C_6 = 1s^2, 2s^1, 2p^3$ is the electronic configuration in first excited state.

Because

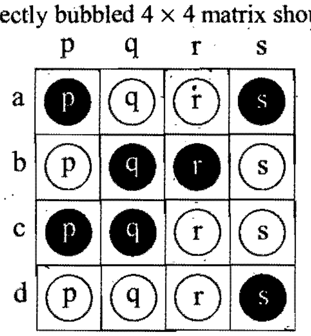
Statement-2: Maximum energy by an electron is possessed in its ground state.

SECTION-IV

Matrix-Matching Type Questions

This section contains 3 questions. Each question contains statements given in two columns which have to be matched. Statements (a, b, c and d) in Column-I have to be matched with statements (p, q, r and s) in Column-II. The answers to these questions have to be appropriately bubbled as illustrated in the following examples:

If the correct matches are (a-p,s); (b-q,r); (c-p,q) and (d-s); then the correctly bubbled 4 × 4 matrix should be as follows:



21. Match the Column-I with Column-II:

	Column-I	Column-II
(a)	$4s$	(p) Circular orbit around nucleus
(b)	$4p$	(q) Non-direction orbitals
(c)	$1s$	(r) Angular momentum = $\frac{2h}{\pi}$
(d)	$3d$	(s) Number of radial node = 0

22. Match the properties of Column-I with the formulae in Column-II:

Column-I	Column-II
(a) Angular momentum of electron	(p) $\sqrt{l(l+1)}\frac{h}{2\pi}$
(b) Orbital angular momentum	(q) $I\omega$
(c) Wavelength of matter wave	(r) $\frac{nh}{2\pi}$
(d) Quantised value(s)	(s) h/p

23. Match the Column-I with Column-II:

	Column-I		Column-II
(a)	Electrons cannot exist in the nucleus	(p)	de Broglie wave
(b)	Microscopic particles in motion are associated with	(q)	Electromagnetic wave
(c)	No medium is required for propagation	(r)	Uncertainty principle
(d)	Concept of orbit was replaced by orbital	(s)	Transverse wave

SECTION-V

Linked Comprehension Type Questions

A chemist was performing an experiment to study the effect of varying voltage on the velocity and de Broglie wavelength of the electrons. In first experiment, the electron was accelerated through a potential difference of 1 kV and in second experiment, it was accelerated through a potential difference of 2 kV.

The wavelength of de Broglie waves associated with electron is given by:

$$\lambda = \frac{h}{\sqrt{2qVm}}$$

where, V is the voltage through which an electron is accelerated.

Putting the values of h, m and q we get:

$$\lambda = \frac{12.3}{\sqrt{V}}\,\text{Å}$$

Answer the following questions:

24. The wavelength of electron will be:
(a) 1.4 times in first case than in second case
(b) 1.4 times in second case than in first case
(c) double in second case than in first case
(d) double in first case than in second case

25. In order to get half velocity of electrons in second case, the applied potential will be:
(a) 0.25 kV (b) 2 kV (c) 0.5 kV (d) 0.75 kV

26. The velocity of electron will be:
(a) same in both cases
(b) 1.4 times in second experiment than in first experiment
(c) double in second experiment than in first experiment
(d) four times in the second case than in first case

Answers

1. (b) **2.** (b) **3.** (a) **4.** (c) **5.** (d) **6.** (d) **7.** (c) **8.** (c)
9. (d) **10.** (b) **11.** (b) **12.** (a, c, d) **13.** (a, b, c) **14.** (a, b, c) **15.** (a, b, c) **16.** (b)
17. (c) **18.** (b) **19.** (d) **20.** (c) **21.** (a-p, q,r) (b-r) (c-p,q) (d-s)
22. (a-q,r) (b-p) (c-s) (d-q,r) **23.** (a-r) (b-p) (c-q,s) (d-r) **24.** (a) **25.** (a) **26.** (b)

CHAPTER 3

RADIOACTIVITY AND NUCLEAR TRANSFORMATION

3.1 RADIOACTIVITY

Radioactivity is a process in which nuclei of certain elements undergo spontaneous disintegration without excitation by any external means.

All heavy elements from bismuth through uranium and a few of lighter elements have naturally occurring isotopes which possess the property of radioactivity. These isotopes have unstable nuclei and attain stability through the phenomenon of radioactivity. The activity results in the emission of a complex type of powerful radiations known as alpha, beta and gamma rays. **All those substances which have the tendency to emit these radiations are termed radioactive materials.** The property of disintegration of a radioactive material is independent of temperature, pressure and other external conditions. **Radioactivity is a nuclear phenomenon,** *i.e.*, the kind of intensity of the radiation emitted by any radioactive substance is absolutely the same whether the element is present as such or in any one of its compounds. $^{226}_{88}Ra$ isotope is radioactive. When this isotope is dissolved in sulphuric acid, it is converted into radium sulphate ($RaSO_4$). The property of radioactivity in radium sulphate and free radium isotope is the same, no doubt that the radium ion in radium sulphate has different number of electrons than free neutral radium isotope.

	Radium atom (Ra)	Radium ion (Ra^{2+})
No. of protons	88	88
No. of electrons	88	86
Atomic mass	226	226

This example clearly shows that the phenomenon of radioactivity does not depend on the orbital electrons but depends only on the composition of nucleus.

In the universe, there are only 81 stable elements having one or more non-radioactive isotopes. No stable isotope exists for the elements above $^{209}_{83}Bi$. Thus, **bismuth** is the heaviest stable nuclide. Two earlier elements **technetium** and **promethium** exist only as radioactive isotopes (see table at the bottom).

3.2 CHARACTERISTICS OF RADIOACTIVE RADIATIONS

The following are the main characteristics of radiations emitted by radioactive materials:

(i) Photographic effect: Radiations affect the photographic plate in a similar manner to that of light. The effect is even observed in dark. The portions of the photographic plate where radiations fall, become blackened after treatment with a developer.

This property is used for the detection of radioactivity.

(ii) Scintillations: When radiations fall on the zinc sulphide (ZnS) screen, flashes of light are produced. This is known as scintillations. The number of particles emitted in unit time can be counted by noting the scintillations produced in the apparatus having a zinc sulphide screen.

The apparatus is called spinthariscope.

(iii) Emission of heat: Radioactive materials continuously emit energy in the form of kinetic energy. Heat energy is

								H									He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Lr	Rf	Db	Sg	Bh	Hs	Mt	Uun	Uuu	Uub		Uuq		Uuh		

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No

Shaded elements are radioactive

produced when the radiation particles collide with matter. Radium produces 134.7 cals of heat per gram per hour.

(iv) Physiological effect: Radioactive radiations have serious physiological effects, which may be cumulative over a period of time. Even a short exposure to intense source of radiations is sufficient to cause painful inflammation. **Gamma rays are most effective.** In general, abnormal cells are more affected than the healthy cells. On account of this property radiations are used to destroy cancerous tumours.

(v) Ionisation of gases: This is the most important effect observed in the case of radioactive radiations. Radiations produce ionisation in the gases through which they are passed. **This effect is used for quantitative measurement of radioactivity.** The radiations cause a number of molecules of the gas to lose electrons and pass into positive ions. The electrons immediately become attached to the neutral molecules, thus making them negative ions. **The total ions of one sign are equal to the total ions of the other type. The rate of production of these ions is proportional to the intensity of radiation.** The extent, to which a definite quantity of a gas is rendered a conductor by a radioactive substance, is a measure of the radioactive power of a radioactive substance. The apparatus used for this purpose is called **electroscope** (Fig. 3.1).

Geiger-Muller counter is based on this effect. The ionisation chamber consists of 90% argon and 10% ethyl alcohol vapour at 10 mm pressure. Due to ionisation, a flow of current occurs, which is measured after amplification.

3.3 HISTORY OF THE DISCOVERY OF RADIOACTIVITY

In 1895, **Henri Becquerel** was studying the effect of sunlight on various phosphorescent minerals, among them a uranium ore. During a period of several cloudy days, he left the uranium sample in a drawer along with some photographic paper wrapped in black paper. Much to his surprise, he discovered that the photographic paper had been fogged by exposure to some invisible radiation from uranium. He called this mysterious property of the ore 'radioactivity'. **Radioactivity means ray-emitting activity.** He further observed that the radioactive mineral emitted these mysterious radiations day after day and month after month and the emission seemed to be endless. The emission was completely unaffected by physical and chemical conditions. A year later, in 1896, **Marie Curie** found that besides uranium and its compounds, thorium was another element which possessed the property of radioactivity. In 1898, **Marie Curie** and her husband **P. Curie** observed that the uranium ore '*pitchblende*', contained more activity than was expected from the uranium which it contained. It must be obviously due to the presence of some other radioactive elements which were far more radioactive than uranium. Finally, they isolated two new radioactive elements **polonium** and **radium**.

Almost in the same period, **G. C. Schmidt** reported that thorium compounds possessed radioactivity. In 1901, **A. Debierne and F.S. Giesel** discovered another new radioactive element actinium in uranium minerals. Further systematic researches led to the discovery of many more radioactive elements. At present over forty such materials are known to exist in nature.

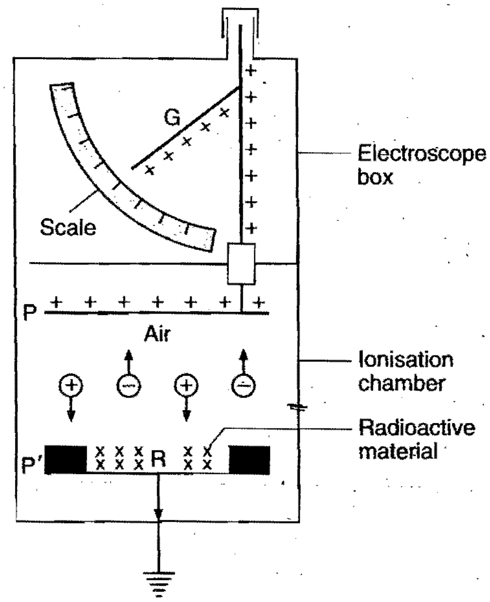


Fig. 3.1 Electroscope for measurement of radioactivity

3.4 ANALYSIS OF RADIOACTIVE RADIATIONS

In 1904, **Rutherford and his co-workers** observed that when radioactive radiations were subjected to a magnetic field or a strong electric field, these were split into three types, as shown in Fig. 3.2. The rays which are attracted towards the negative plate are positively charged and are called **alpha (α) rays.** The rays which are deflected towards the positive plate are negatively charged and are called **beta (β) rays.** The third type of rays which are not deflected on any side but move straight are known as **gamma (γ) rays.** The important properties of these radiations are tabulated on next page:

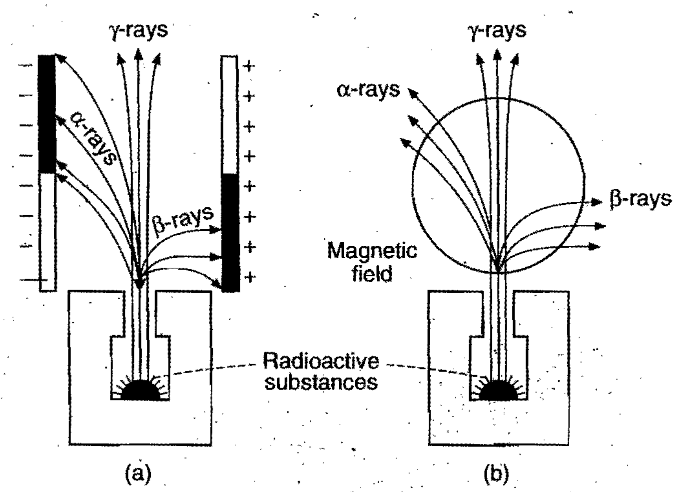


Fig. 3.2 (a) Deflection of radioactive rays in electric field and (b) Emission of radioactive rays and their deflection in a magnetic field.

Property	α-rays	β-rays	γ-rays
1. Nature	These consist of small positively charged particles which are merely nuclei of helium atoms, each consisting of 2 protons and 2 neutrons. These are represented as ${}_2^4He$.	These consist of negatively charged particles which have the same e/m value as the cathode rays. β-rays are merely electrons. The β-rays are represented as ${}_{-1}^{0}\beta$ or ${}_{-1}^{0}e$.	γ-rays are similar to X-rays. These are neutral in nature. They have very small wavelengths of the order of 10^{-10} to 10^{-13} m.
2. Velocity	The α-rays are ejected with high velocities ranging from 1.4×10^9 to 1.7×10^9 cm/sec. The velocity of α-rays depends upon the kind of nucleus from which they are emitted.	The β-rays are much faster than α-rays. They have generally different velocities sometimes approaching the velocity of light.	They travel with the velocity of light.
3. Penetrating power	α-particles have small penetrating power due to relatively larger size. They are stopped by a piece of aluminium foil of 0.1 mm thickness.	β-rays are more penetrating than α-particles. This is due to small size and high velocity. These are stopped by a 1 cm thick sheet of aluminium.	Due to high velocity and non-material character, γ-rays are 10^{10} times more penetrating than α-rays.

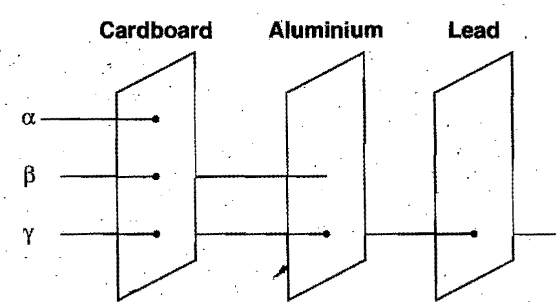


Fig. 3.2 (c) Comparison of penetrating power of α, β and γ-rays

Property	α-rays	β-rays	γ-rays
4. Ionising power	α-particles produce intense ionisation in gases. Ionising power is 100 times greater than β-rays and 10,000 times greater than γ-rays. This is due to high kinetic energy.	Due to low value of kinetic energy ionising power is less than α-particles but 100 times greater than γ-rays.	γ-rays produce minimum ionisation or no ionisation.
5. Photographic effect	α-particles affect the photographic plate.	β-rays effect on a photographic plate is greater than α-particles. Like cathode rays, β-rays produce X-rays.	γ-rays have little effect on photographic plate.
6. Effect on zinc sulphide screen	α-particles produce luminosity on ZnS screen due to high kinetic energy.	β-particles have little effect on ZnS screen due to low kinetic energy.	γ-rays have very little effect on ZnS screen.

Note :

(i) The quantum energy of γ-rays emitted by a radioactive substance can have unique and discrete values.

(ii) The energy of α-particles emitted by a radioactive substance (α-emitter) has unique value.

(iii) The energy of β-particles emitted by a radioactive substance, (β-emitter) can have any value between zero and end point energy.

3.5 CAUSE OF RADIOACTIVITY

Except in the case of ordinary hydrogen, all other nuclei contain both neutrons and protons. A look at the stable nuclei shows that the ratio n/p (neutrons/protons) in them is either equal to 1 or more than 1. The ratio is ≈ 1 in all the light-stable nuclei up to calcium (${}_{20}^{40}Ca$) and thereafter the ratio is greater than 1 and increases up to 1.6 for heavy stable nuclei as shown in the following table:

Neutron-proton ratio in some stable nuclei

Isotope	${}_6^{12}C$	${}_7^{14}N$	${}_8^{16}O$	${}_{10}^{20}Ne$	${}_{20}^{40}Ca$	${}_{30}^{64}Zn$	${}_{40}^{90}Zr$	${}_{50}^{120}Sn$	${}_{60}^{150}Nd$	${}_{80}^{202}Hg$
n	6	7	8	10	20	34	50	70	90	122
p	6	7	8	10	20	30	40	50	60	80
n/p	1	1	1	1	1	1.13	1.25	1.40	1.50	1.53

The variation of n *versus* p for some nuclei is shown in Fig. 3.3.

The stable nuclei lie within the shaded area which is called the **region** or **zone of stability.** All the nuclei falling outside this zone are invariably radioactive and unstable in nature. **Nuclei that fall above the stability zone have an excess of neutrons while those lying below have more protons.** Both of these cause instability. These nuclei attain stability by making adjustment in the n/p ratio.

Two cases thus arise:

(i) n/p ratio is higher than required for stability. Such nuclei have a tendency to emit β-rays, *i.e.*, transforming a neutron into proton.

$$ {}_0^1n \longrightarrow {}_1^1p + {}_{-1}^{0}e \text{ (β-particle)} $$

Thus, in β-emission n/p ratio decreases. For example, in the change of ${}_6^{14}C$ to ${}_7^{14}N$, n/p ratio decreases from 1.33 to 1.

$$^{14}_{6}C \longrightarrow ^{14}_{7}N + ^{\ 0}_{-1}e$$

Similarly, in the following examples, the n/p ratio decreases during β-emission:

	$^{32}_{15}P \longrightarrow$	$^{32}_{16}S + ^{\ 0}_{-1}e$	
n/p ratio	17/15	16/16	
	$^{87}_{36}Kr \longrightarrow$	$^{87}_{37}Rb + ^{\ 0}_{-1}e$	
n/p ratio	51/36	50/37	

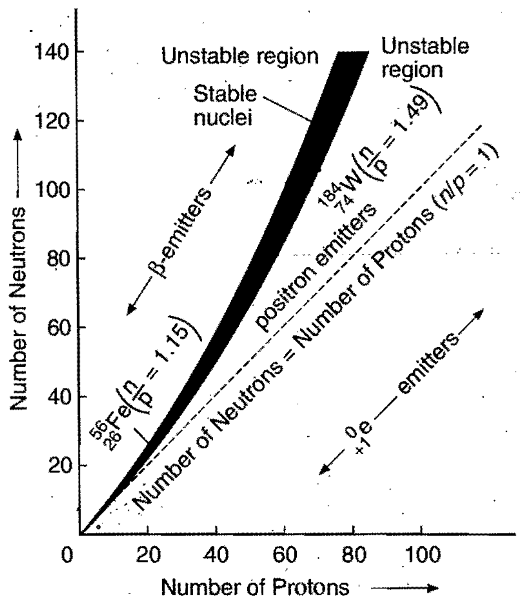


Fig. 3.3 Variation of number of neutrons with number of protons in stable non-radioactive nuclei

(ii) n/p ratio is lower than required for stability. Such nuclei can increase n/p ratio by adopting any one of the following three ways:

(a) By emission of an alpha particle (Natural radioactivity):

	$^{238}_{92}U \longrightarrow$	$^{234}_{90}Th + ^{4}_{2}He$ (α-particle)	
n/p ratio	146/92 = 1.58	144/90 = 1.60	

(b) By emission of a positron (Artificial emission):

	$^{13}_{7}N \longrightarrow$	$^{13}_{6}C + ^{\ 0}_{+1}e$	
n/p ratio	6/7	7/6	

(c) By K-electron capture:

	$^{194}_{79}Au + ^{\ 0}_{-1}e \longrightarrow$	$^{194}_{78}Pt$	
n/p ratio	115/79	116/78	

Alpha emission is usually observed in natural radioactive isotopes while emission of positron or K-electron capture is observed in artificial radioactive isotopes. The unstable nuclei continue to emit alpha or beta particles until a stable nucleus comes into existence.

Conclusion: (i) For the elements (mass number $A \le 40$), nature prefers the number of protons and neutrons in the nucleus to be same or perhaps one more neutron than protons. *e.g.*, $^{7}_{3}Li$, $^{12}_{6}C$, $^{16}_{8}O$, $^{32}_{16}S$, etc.

(ii) For the elements (mass number $A \ge 40$), there is preference for the number of neutrons to be greater than the number of protons ($n > Z$), *e.g.*, $^{11}_{5}B$ is stable but $^{11}_{6}C$ is not. There are two stable elements $^{1}_{1}H$ and $^{3}_{2}He$, in which number of neutrons is less than that of protons.

(iii) Beyond Bi ($Z = 83$), all isotopes are unstable and radioactive. These elements do not have a strong nuclear **"superglue"** to hold nucleons together.

Illustration :

Nuclide	$\frac{n}{p}$ Ratio	Nature of Emission
$^{35}_{16}S$	$\frac{19}{16} = 1.2$	β-emission $^{35}_{16}S \longrightarrow ^{35}_{17}Cl + ^{\ 0}_{-1}e$
$^{17}_{9}F$	$\frac{8}{9}\left(\frac{n}{p} < 1\right)$	Positron emission $^{17}_{9}F \longrightarrow ^{17}_{8}O + ^{\ 0}_{+1}e$
$^{105}_{47}Ag$	$\frac{n}{p} = \frac{58}{47} = 1.23$	Lies below stability belt, it has a heavy nucleus and it decays by K-electron capture. $^{105}_{47}Ag + ^{\ 0}_{-1}e \longrightarrow ^{105}_{46}Pd + h\nu$
$^{238}_{92}U$	$\frac{n}{p} = \frac{146}{92} = 1.59$	It is a neutron rich species. It undergoes decay by α-emission. $^{238}_{92}U \longrightarrow ^{234}_{90}Th + ^{4}_{2}He$

Some other examples : $^{60}_{29}Cu$ (positron emission), $^{140}_{54}Xe$ (α or β-decay), $^{240}_{93}N_p$ (α-decay), $^{33}_{15}P$ (β-decay), $^{125}_{53}I$ (K-electron capture).

3.6 THEORY OF RADIOACTIVE DISINTEGRATION

Rutherford and **Soddy**, in 1903, postulated that radioactivity is a nuclear phenomenon and all the radioactive changes are taking place in the nucleus of the atom. They presented an interpretation of the radioactive processes and the origin of radiations in the form of a theory known as **theory of radioactive disintegration.** The main points of the theory are:

(i) The atomic nuclei of the radioactive elements are unstable and liable to disintegrate any moment.

(ii) The disintegration is spontaneous, *i.e.*, constantly breaking. The rate of breaking is not affected by external factors like temperature, pressure, chemical combination, etc.

(iii) During disintegration, atoms of new elements called daughter elements having different physical and chemical properties than the parent element come into existence.

(iv) During disintegration, either alpha or beta particles are emitted from the nucleus.

The disintegration process may proceed in one of the following two ways:

(a) α-particle emission: When an α-particle [$^{4}_{2}He$] is emitted from the nucleus of an atom of the parent element, the nucleus of the new element, called daughter element, possesses atomic mass or atomic mass number less by four units and nuclear charge or atomic number less by 2 units because α-particle has mass of 4 units and nuclear charge of two units.

The daughter element after α-emission is called an isodiaphere of parent element.

	Parent element $\xrightarrow{-\alpha}$	**Daughter element**
Atomic mass	W	$W-4$
Atomic number	Z	$Z-2$

For example, in the following transformations, each α-particle emission is accompanied by decrease of atomic mass by 4 and of atomic number by 2.

$$\underset{\text{(Radium)}}{{}^{226}_{88}\text{Ra}} \longrightarrow \underset{\text{(Radon)}}{{}^{222}_{86}\text{Rn}} + {}^{4}_{2}\text{He}$$

$$\underset{\text{(Uranium)}}{{}^{238}_{92}\text{U}} \longrightarrow \underset{\text{(Thorium)}}{{}^{234}_{90}\text{Th}} + {}^{4}_{2}\text{He}$$

$$\underset{\text{(Bismuth)}}{{}^{213}_{83}\text{Bi}} \longrightarrow \underset{\text{(Thallium)}}{{}^{209}_{81}\text{Tl}} + {}^{4}_{2}\text{He}$$

$$\underset{\text{(Polonium)}}{{}^{215}_{84}\text{Po}} \longrightarrow \underset{\text{(Lead)}}{{}^{211}_{82}\text{Pb}} + {}^{4}_{2}\text{He}$$

(b) β-particle emission: β-particle is merely an electron which has negligible mass. Whenever a beta particle is emitted from the nucleus of a radioactive atom, the nucleus of the new element formed possesses the same atomic mass but nuclear charge or atomic number is increased by 1 unit over the parent element. Beta particle emission is due to the result of decay of neutron into proton and electron.

$${}^{1}_{0}n \longrightarrow {}^{1}_{1}p + {}^{0}_{-1}e$$

The electron produced escapes as a beta particle leaving proton in the nucleus.

	Parent element $\xrightarrow{-\beta}$	**Daughter element**
Atomic mass	W	W
Atomic number	Z	$Z+1$

For example, in the following transformations, beta particle emission results in increase of atomic number by one without any change in atomic mass, *i.e.*, **daughter element is an isobar of parent element.** (See table 3.1)

Table 3.1 Isotopes, Isobars, Isotones, Isomers, Isoters and Isodiapheres

	Characteristics Z = at. no., A = mass no., N = neutrons, P = protons	**Examples**
Isotopes	Z = same, A = different	${}^{1}_{1}\text{H}$, ${}^{2}_{1}\text{H}$, ${}^{3}_{1}\text{H}$, ${}^{235}_{92}\text{U}$, ${}^{238}_{92}\text{U}$
Isobars	Z = different, A = same	${}^{228}_{88}\text{Ra}$, ${}^{228}_{89}\text{Ac}$, ${}^{228}_{90}\text{Th}$
Isotones	N = same, nucleons = different, Z = different	${}^{39}_{18}\text{Ar}$, ${}^{40}_{19}\text{K}$
Isomers	N = same, P = same, Z = same, A = same, Nuclear energy levels = different	U - X_2, U - Z
Isoters	No. of atoms = same, No. of electrons = same, physical properties = same.	CO_2, N_2O
Isodiapheres	Isotopic excess mass (N—P) = same.	${}_{92}U^{235}$, ${}_{90}Th^{231}$

$$\underset{\text{(Lead)}}{{}^{214}_{82}\text{Pb}} \longrightarrow \underset{\text{(Bismuth)}}{{}^{214}_{83}\text{Bi}} + {}^{0}_{-1}e$$

$$\underset{\text{(Thorium)}}{{}^{234}_{90}\text{Th}} \longrightarrow \underset{\text{(Protoactinium)}}{{}^{234}_{91}\text{Pa}} + {}^{0}_{-1}e$$

$$\underset{\text{(Bismuth)}}{{}^{213}_{83}\text{Bi}} \longrightarrow \underset{\text{(Polonium)}}{{}^{213}_{84}\text{Po}} + {}^{0}_{-1}e$$

Special case: If in a radioactive transformation, 1 alpha and 2 beta particles are emitted, the resulting nucleus possesses the same atomic number but atomic mass is less by 4 units. A radioactive transformation of this type always produces an isotope of the parent element.

$${}^{W}_{Z}A \xrightarrow{-\alpha} {}^{W-4}_{Z-2}B \xrightarrow{-\beta} {}^{W-4}_{Z-1}C \xrightarrow{-\beta} {}^{W-4}_{Z}D$$

A and D are isotopes.

(v) Gamma rays are emitted due to secondary effects. After the emission of an alpha particle or a beta particle, the nucleus is left behind in excited state due to recoil. The excess of energy is released in the form of gamma rays. **Thus, γ-rays arise from energy rearrangements in the nucleus.** As γ-rays are short wavelength electromagnetic radiations with no charge and no mass, **their emission from a radioactive element does not produce a new element.**

On passing through an absorbing material, the intensity of γ-radiation decreases exponentially with the thickness traversed and is given by :

$$I = I_0 e^{-\mu x}$$

where μ = Absorption coefficient,
x = Thickness
I_0 = Initial intensity
I = Transmitted intensity

All radioactive nuclei have the same probability of disintegration. However, a radioactive nucleus may undergo decay next moment while some other may have to wait for billions of years to decay one cannot predict, when a particular atom will decay.

(vi) Internal conversion: An excited nucleus, in some cases, may return to its ground state by giving up its excitation energy to one of the orbital electrons around it. The emitted electron has a kinetic energy equal to the lost nuclear excitation energy minus the binding energy of the electron in the atom.

Kinetic energy of the ejected electron

= Available excitation energy

− Binding energy of the ejected electron

This process is called *internal conversion* and emitted electron is called conversion electron.

(vii) Brems strahlung (German word meaning 'Breaking Radiation'): Continuous γ-radiations emitted when β-particles are slowed down by interaction with atomic nucleus.

Note: Counting of the number of α and β-particles in a radioactive transformation:

$$\text{Parent element} \longrightarrow \text{Daughter element}$$

$$_{Z_1}A^{M_1} \longrightarrow {}_{Z_2}B^{M_2}$$

$$\text{Number of } \alpha\text{-particles} = \frac{\text{Change in mass number}}{4}$$

$$= \frac{M_1 - M_2}{4}$$

Let 'x' α and 'y' β-particles be emitted.

Atomic number of parent element $- 2x + y$
= Atomic number of daughter element

$$Z_1 - 2x + y = Z_2$$

3.7 GROUP DISPLACEMENT LAW

This law was presented by **Fajan, Soddy** and **Russel** in 1913 to explain the changes which occur when an alpha particle or a beta particle is emitted from a radioactive element. According to this law, **"when an α-particle is emitted, the daughter element has atomic number 2 units less than that of the parent element. It is consequently displaced two places (groups) to the left in the periodic table. When a β-particle is emitted, the daughter element has an atomic number 1 unit higher than that of the parent element. It is consequently displaced one place (group) to the right in the periodic table."**

Examples

(i) Polonium ($^{214}_{84}$Po) belongs to group 16 (VIA) of the periodic table. On losing an alpha particle, it is transformed into lead ($^{210}_{82}$Pb) which belongs to group 14 (IVA), *i.e.*, two places to the left of the parent element, polonium.

$$\underset{\substack{16 \\ (\text{VI A})}}{^{214}_{84}\text{Po}} \longrightarrow \underset{\substack{14 \\ (\text{IV A})}}{^{210}_{82}\text{Pb}}$$

(ii) Bismuth ($^{213}_{83}$Bi) belongs to group 15 (VA) of the periodic table. It emits an alpha particle resulting in the formation of thallium which belongs to group 13 (IIIA), *i.e.*, two places to the left of the parent element, bismuth.

$$\underset{\substack{15 \\ (\text{V A})}}{^{213}_{83}\text{Bi}} \longrightarrow \underset{\substack{13 \\ (\text{III A})}}{^{209}_{81}\text{Tl}}$$

(iii) Carbon ($^{14}_{6}$C) belongs to group 14 (IV A) and emits a β-particle forming nitrogen ($^{14}_{7}$N) which belongs to group 15 (V A), *i.e.*, one place to the right of the parent element.

$$\underset{(\text{IV A})}{^{14}_{6}\text{C}} \longrightarrow \underset{(\text{V A})}{^{14}_{7}\text{N}} + {}^{0}_{-1}e$$

(iv) Phosphorus ($^{32}_{15}$P) belongs to group 15 (VA) and emits a β-particle forming sulphur ($^{32}_{16}$S) which belongs to group 16 (VIA), *i.e.*, one place right to the parent element.

$$\underset{(\text{V A})}{^{32}_{15}\text{P}} \longrightarrow \underset{(\text{VI A})}{^{32}_{16}\text{S}} + {}^{0}_{-1}e$$

The above examples follow group displacement law rigidly in accordance with the statement. However, there are a number of examples where confusion arises regarding the position of the element in the periodic table if the above statement is followed rigidly.

$^{27}_{12}$Mg is β-radioactive. It belongs to group 2 (IIA) of the periodic table. On losing a beta particle, it is transformed to aluminium ($^{27}_{13}$Al) which belongs to group 13 (IIIA), *i.e.*, 11 places right to the parent element.

$$^{27}_{12}\text{Mg} \longrightarrow {}^{27}_{13}\text{Al} + {}^{0}_{-1}e$$

$^{234}_{90}$Th is a member of actinide series. All the fourteen members of the actinide series have been placed along with actinium in the III B group, *i.e.*, group 3. It emits a beta particle and is transformed to protactinium ($^{234}_{91}$Pa) which also belongs to actinide series, *i.e.*, group 3 of the periodic table.

$$^{234}_{90}\text{Th} \longrightarrow {}^{234}_{91}\text{Pa} + {}^{0}_{-1}e$$

Hence, group displacement law should be applied with great care especially in the case of elements of lanthanide series (57 to 71), actinide series (89 to 103), VIII group (26 to 28; 44 to 46; 76 to 78), IA and IIA groups. It is always beneficial to keep in mind the setup and skeleton of the extended form of periodic table.

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII		
1	2	3	4	5	6	7	8	9	10

IB	IIB	IIIA	IVA	VA	VIA	VIIA	Zero
11	12	13	14	15	16	17	18

SOME SOLVED EXAMPLES

Example 1. *Calculate the number of neutrons in the remaining atom after emission of an alpha particle from $^{238}_{92}U$ atom.*

Solution: On account of emission of an alpha particle, the atomic mass is decreased by 4 units and atomic number by 2 units.

So, Atomic mass of daughter element = 234

Atomic number of daughter element = 90

Number of neutrons = atomic mass − atomic number

= 234 − 90 = 144

Example 2. *Radioactive disintegration of $^{226}_{88}Ra$ takes place in the following manner into RaC,*

$$Ra \xrightarrow{-\alpha} Rn \xrightarrow{-\alpha} RaA \xrightarrow{-\alpha} RaB \xrightarrow{-\beta} RaC$$

Determine mass number and atomic number of RaC.

Solution: Parent element is $^{226}_{88}$Ra.

Atomic mass = 226

Atomic number = 88

Ra*C* is formed after the emission of 3 alpha particles. Mass of 3 alpha particles = 3 × 4 = 12

So, Atomic mass of RaC = (226 − 12) = 214

With emission of one α-particle, atomic number is decreased by 2 and with the emission of one β-particle, atomic number is increased by 1.

So, Atomic number of RaC $= 88-(3\times 2)+1=83$

Example 3. *A radioactive element A disintegrates in the following manner,*

$$A \xrightarrow{-\alpha} B \xrightarrow{-\beta} C \xrightarrow{-\beta} D$$

which ones of the elements A, B, C and D are isotopes and which ones are isobars?

Solution: Let the mass number and atomic number of element A be M and Z, respectively. The following changes shall occur during disintegration:

$$^{M}_{Z}A \xrightarrow{-\alpha} {}^{M-4}_{Z-2}B \xrightarrow{-\beta} {}^{M-4}_{Z-1}C \xrightarrow{-\beta} {}^{M-4}_{Z}D$$

A and D are isotopes as both have same value of Z.

B, C and D are isobars as these have same values of atomic mass.

Example 4. *$^{234}_{90}Th$ disintegrates to give $^{206}_{82}Pb$ as the final product. How many alpha and beta particles are emitted in this process?*

Solution: $\underset{\text{Parent}}{^{234}_{90}\text{Th}} \longrightarrow \underset{\text{End product}}{^{206}_{82}\text{Pb}}$

Decrease in mass $= (234-206)=28$

Mass of α-particle $= 4$

So, Number of α-particles emitted $= \dfrac{28}{4} = 7$

No. of β-particles emitted $= 2\times$ No. of α-particles–(At. No. of parent – At. No. of end product)

$= 2\times 7-(90-82)=6$

Example 5. *The atomic mass of thorium is 232 and its atomic number is 90. During the course of its radioactive disintegration 6α and 4β-particles are emitted. What is the atomic mass and atomic number of the final atom?*

Solution: Decrease in mass due to emission of 6α-particles $= 6\times 4 = 24$.

So, Atomic mass of the product atom $= (232-24)=208$

No. of β-particles emitted $= 2\times$ No. of α-particles $-(Z_{\text{Thorium}} - Z_{\text{Final atom}})$

$$4 = 2\times 6-(90-Z_{\text{Final atom}})$$

or $Z_{\text{Final atom}} = 82$

Example 6. *An atom has atomic mass 232 and atomic number 90. During the course of disintegration, it emits 2β-particles and few α-particles. The resultant atom has atomic mass 212 and atomic number 82. How many α-particles are emitted during this process?*

Solution: The decrease in atomic mass

$= (232-212)=20$

Decrease in mass occurs due to emission of α-particles. Let x be the number of alpha particles emitted.

Mass of 'x' α-particles $= 4x$

So, $4x = 20$

or $x = \dfrac{20}{4} = 5$

Alternative method: This can also be determined by the application of the following equation:

No. of β-particles emitted $= 2\times$ No. of α-particles emitted $-(Z_{\text{Parent}} - Z_{\text{End product}})$

$$2 = 2\times x-(90-82)$$

or $x=5$

Example 7. *How many moles of helium are produced when one mole of $^{238}_{92}U$ disintegrates into $^{206}_{82}Pb$?*

Solution: Radioactive change is

$$^{238}_{92}\text{U} \longrightarrow {}^{206}_{82}\text{Pb}$$

Decrease in mass $= (238-206)=32$

Let the number of α-particles emitted be x.

$$4x = 32$$
$$x = 8$$

Thus, 8 moles of helium are produced when one mole of $^{238}_{92}$U disintegrates into $^{206}_{82}$Pb.

Example 8. *How many 'α' and 'β'-particles will be emitted when $_{90}Th^{234}$ changes into $_{84}Po^{218}$?*

Solution: The change is;

$$\underset{\text{Parent}}{_{90}\text{Th}^{234}} \longrightarrow \underset{\text{End product}}{_{84}\text{Po}^{218}}$$

Decrease in mass $= (234-218)=16$ amu

Mass of 1 α-particle $= 4$ amu

Therefore, number of α-particles emitted $= \dfrac{16}{4} = 4.$

Number of β-particles emitted

$= 2\times$ No. of α-particles emitted – (At. No. of parent – At. No. of end product)

$= 2\times 4-(90-84)=(8-6)=2$

Hence, number of α-particles emitted $= 4$

and number of β-particles emitted $= 2$

Example 9. *$^{238}_{92}U$ is a natural an α-emitter. After α-emission, the residual nucleus U_{X_1} in turns emits a β-particle to produce another nucleus U_{X_2}. Find out the atomic number and mass number of U_{X_1} and U_{X_2}. Also if uranium belongs to IIIrd group to which group U_{X_1} and U_{X_2} belong.*

Solution: $^{238}_{92}\text{U} - {}^{4}_{2}\text{He} \longrightarrow {}^{234}_{90}\text{U}_{X_1}$

$$^{234}_{90}\text{U}_{X_1} - ({}^{0}_{-1}e) \longrightarrow {}^{234}_{91}\text{U}_{X_2}$$

Both U_{X_1} and U_{X_2} will belong to IIIrd group because both lie in actinide series.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

1. During the transformation of $^{a}_{c}X$ to $^{b}_{d}Y$, the number of β-particles emitted are: **[PET (Kerala) 2006, 08]**

(a) $\frac{a-b}{4}$ (b) $d+\frac{a-b}{2}+c$

(c) $d+\left(\frac{a-b}{2}\right)-c$ (d) $2c-d+a-b$

[**Ans.** (c)]

[**Hint:** No. of α-particles $=\left(\frac{a-b}{4}\right)$

$$Z_1 - 2\alpha + \beta = Z_2$$
$$\beta = Z_2 - Z_1 + 2\alpha$$
$$= d - c + 2\frac{(a-b)}{4}$$
$$= d + \frac{(a-b)}{2} - c]$$

2. A radioactive nuclide emits γ-rays due to:

(a) K-electron capture
(b) nuclear transition from higher to lower energy state
(c) presence of greater number of neutrons than protons
(d) presence of greater number of protons than neutrons

[**Ans.** (b)]

[**Hint:** After α, β-emission, nucleus goes to excited state; when it returns to normal state, emission of γ-radiations takes place.]

3. In which of the following transformations, the β-particles are emitted?

(a) Proton to neutron (b) Neutron to proton
(c) Proton to proton (d) Neutron to neutron

[**Ans.** (b)]

[**Hint:** ${}_0^1n \longrightarrow {}_1^1\mathrm{H} + {}_{-1}^{\ 0}e$ + antineutrino]

4. In the radioactive decay:

$${}_Z^AX \longrightarrow {}_{Z+1}^{\ \ A}Y \longrightarrow {}_{Z-1}^{A-4}Z \longrightarrow {}_{Z-1}^{A-4}Z^*$$

the sequence of emission is:

(a) α, β, γ (b) β, α, γ (c) γ, α, β (d) β, γ, α

[**Ans.** (b)]

[**Hint:** ${}_Z^AX - {}_{-1}^{\ 0}e \longrightarrow {}_{Z+1}^{\ \ A}Y;\ {}_{Z+1}^{\ \ A}Y - {}_2^4\mathrm{He} \longrightarrow {}_{Z-1}^{A-4}Z;$
${}_{Z-1}^{A-4}Z - \gamma \longrightarrow {}_{Z-1}^{A-4}Z^*$]

5. Which of the following elements is an isodiaphere of ${}_{92}^{235}\mathrm{U}$?

(a) ${}_{83}^{209}\mathrm{Bi}$ (b) ${}_{82}^{212}\mathrm{Pb}$ (c) ${}_{90}^{231}\mathrm{Th}$ (d) ${}_{91}^{231}\mathrm{Pa}$

[**Ans.** (c)]

[**Hint:** Isodiapheres are formed by α-emission.
${}_{92}^{235}\mathrm{U} - {}_2^4\mathrm{He} \longrightarrow {}_{90}^{231}\mathrm{Th}$ (Isodiaphere)]

6. A certain radioactive material ${}_Z^AX$ starts emitting α and β particles successively such that the end product is ${}_{Z-3}^{A-8}Y$. The number of α and β particles emitted are : **(VITEEE 2008)**

(a) 4 and 3 respectively (b) 2 and 1 respectively
(c) 3 and 4 respectively (d) 3 and 8 respectively

[**Ans.** (b)]

[**Hint:** ${}_Z^AX \longrightarrow {}_{Z-3}^{A-8}Y$

$$\text{Number of α-particles} = \frac{\text{Change in mass number}}{4} = \frac{8}{4} = 2$$

$Z - 2 \times$ Number of α-particles + Number of β-particles $= Z - 3$

$Z - 2 \times 2 +$ Number of β-particles $= Z - 3$

Number of β- particles = 1]

3.8 RADIOACTIVE DISINTEGRATION SERIES

Elements beyond bismuth are all radioactive in nature. Most of them have several radioactive isotopes. These radioactive elements disintegrate to give new elements which again disintegrate to form other elements and so on. The process continues till a non-radioactive end product is reached.

"The whole chain of such elements starting from the parent element (radioactive) to the end element (non-radioactive) is called a radioactive series or a family".

All the naturally occurring radioactive elements above atomic number 82 belong to one of the three radioactive series. These are known as:

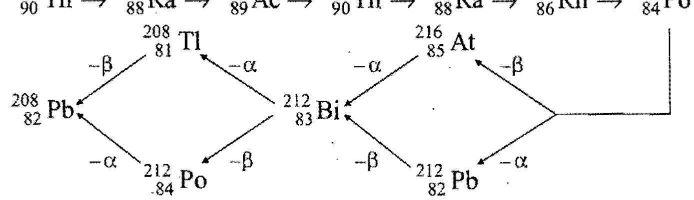
(i) Thorium series (ii) Uranium series
(iii) Actinium series

Uranium and thorium series have been named on the basis of long lived isotopes of ^{238}U and ^{232}Th. The parent element of **actinium** series is ^{235}U but originally it was thought to be an isotope of actinium, ^{227}Ac. The three series are also referred to $4n$ (thorium), $4n+2$ (uranium) and $4n+3$ (actinium) series as when the mass numbers of various members belonging to these series when divided by four, either there is no remainder (as in thorium series) or the remainder is 2 (as in uranium series) or 3 (as in actinium series). The end product in all the three series is an isotope of lead which is stable and non-radioactive in nature. The following table shows the main characteristics of three radioactive series:

Series	First member	Half life of first member in years	Last member	Atomic masses when divided by 4, the remainder	No. of α-particles emitted	No. of β-particles emitted
Thorium ($4n$)	${}_{90}^{232}\mathrm{Th}$	1.4×10^{10}	${}_{82}^{208}\mathrm{Pb}$	0	6	4
Uranium ($4n+2$)	${}_{92}^{238}\mathrm{U}$	4.51×10^9	${}_{82}^{206}\mathrm{Pb}$	2	8	6
Actinium ($4n+3$)	${}_{92}^{235}\mathrm{U}$	7.07×10^8	${}_{82}^{207}\mathrm{Pb}$	3	7	4

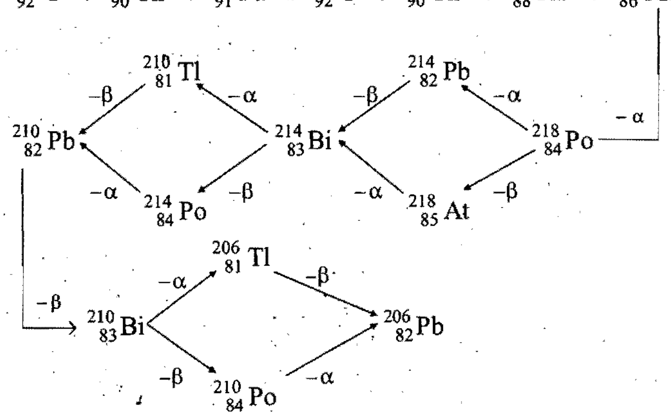
(i) Thorium series ($4n$ series):

$${}_{90}^{232}\mathrm{Th} \xrightarrow{-\alpha} {}_{88}^{228}\mathrm{Ra} \xrightarrow{-\beta} {}_{89}^{228}\mathrm{Ac} \xrightarrow{-\beta} {}_{90}^{228}\mathrm{Th} \xrightarrow{-\alpha} {}_{88}^{224}\mathrm{Ra} \xrightarrow{-\alpha} {}_{86}^{220}\mathrm{Rn} \xrightarrow{-\alpha} {}_{84}^{216}\mathrm{Po}$$

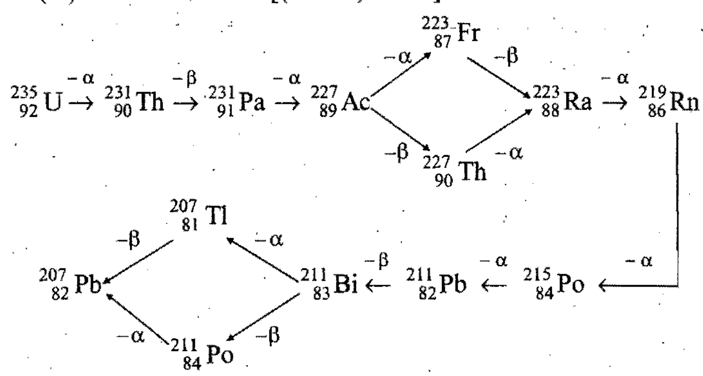


(ii) Uranium series [(4n + 2) series]:

$$^{238}_{92}U \xrightarrow{-\alpha} {}^{234}_{90}Th \xrightarrow{-\beta} {}^{234}_{91}Pa \xrightarrow{-\beta} {}^{234}_{92}U \xrightarrow{-\alpha} {}^{230}_{90}Th \xrightarrow{-\alpha} {}^{226}_{88}Ra \xrightarrow{-\alpha} {}^{222}_{86}Rn$$



(iii) Actinium series [(4n + 3) series]:



Only 18 radioactive isotopes with atomic number 82 or less are found in nature. ^{14}C is the exception because it is continuously synthesized in our atmosphere. All these natural radioactive elements have half-life longer than 10^9 yrs (age of earth). Another 45 radioactive isotopes having atomic number greater than 82 are also found in nature and fall in above three natural decay series.

Similarities between Radioactive Series

(i) In all the series, there is an element of zero group with atomic number 86. This element comes in the gaseous state and is called emanation. Different names are given to three isotopes. These are **radon** in uranium series, **thoron** in thorium series and **actinon** in actinium series.

(ii) In all the series, the last product is an isotope of lead (atomic number 82), ^{206}Pb in uranium series, ^{207}Pb in actinium series while ^{208}Pb in thorium series. Due to this reason, lead is found in nature as a mixture of these three isotopes.

(iii) In all the series, there are certain elements which disintegrate in a branching process by emitting either α or β-particles. The species thus formed are then disintegrated in such a way as to give a common product.

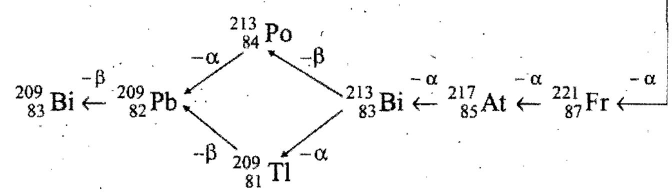
Neptunium series [(4n + 1) series]: For many years scientists speculated upon the failure to find a disintegration series in nature whose isotopic masses carry a numerical relationship of 4n + 1. The most reasonable explanation for the absence of this series in nature was that no member of this series was sufficiently long lived to have survived over the years since the series might have been formed. Except the last member, all other members of this series have been obtained by artificial means. The name of this series is given on the long lived isotope of neptunium (Half life $^{237}_{93}Np = 2.25 \times 10^6$ years). This family differs from the other three naturally occurring series in the following respects:

(a) The last member of this series is an isotope of bismuth ($^{209}_{83}Bi$) and not an isotope of lead.

(b) The only member of this series which is found in nature is the last member.

(c) The series does not contain gaseous emanation.

$$^{237}_{93}Np \xrightarrow{-\alpha} {}^{233}_{91}Pa \xrightarrow{-\beta} {}^{233}_{92}U \xrightarrow{-\alpha} {}^{229}_{90}Th \xrightarrow{-\alpha} {}^{225}_{88}Ra \xrightarrow{-\beta} {}^{225}_{89}Ac$$



In this series, seven alpha and four beta particles are emitted.

3.9 RATE OF DISINTEGRATION AND HALF LIFE PERIOD

The radioactive decay of the different radioactive substances differ widely. The rate of disintegration of a given substance depends upon the **nature of disintegrating substance and its total amount. The law of radioactive disintegration may be defined as "the quantity of radioactive substance which disappears in unit time is directly proportional to the amount* of radioactive substance present or yet not decayed."** The quantity of the radioactive substance which disintegrates or disappears in unit time is called rate of disintegration.

The rate of disintegration decreases with time as the amount of radioactive substance decreases with time. One of the most important characteristics of the radioactive disintegration is that a certain definite fraction of a radioactive sample undergoes disintegration in a definite period of time. This time period does not depend upon the initial amount of the radioactive substance.

For example, whatever be the amount (initial) of ^{131}I taken, it becomes half within 8 days. This has been shown in Fig. 3.4 (a).

Initial amount of ^{131}I	**The amount of ^{131}I after 8 days**
20 grams	10 grams
10 grams	5 grams
5 grams	2.5 grams, etc.

Rutherford introduced a constant known as **half life period.** It is defined **"as the time during which half the amount of a given sample of the radioactive substance disintegrates".**

* Amount of a radioactive substance can be taken in terms of number of atoms or moles or grams, milligrams, etc.

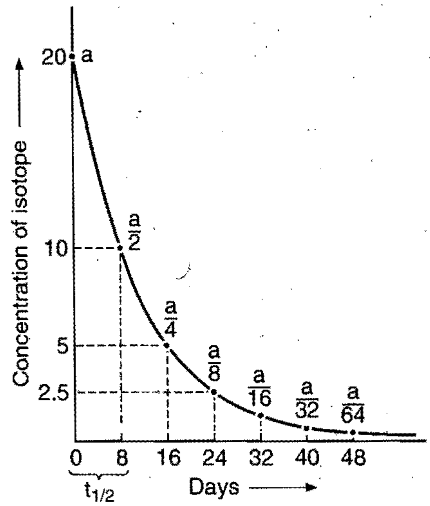


Fig. 3.4 (a)

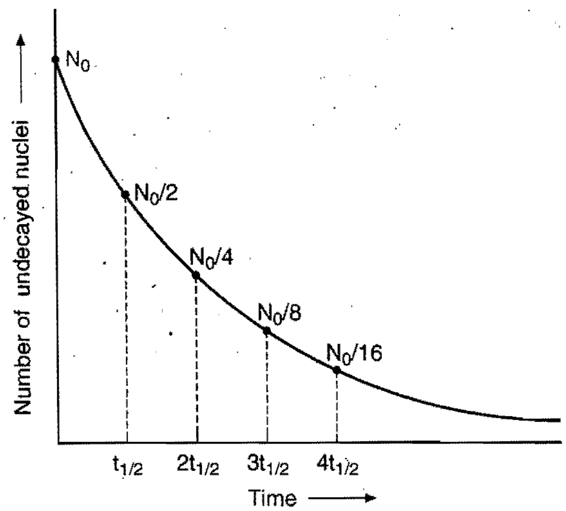


Fig. 3.4 (b)

Every radioactive element is characterised by a definite constant value of half life period. Half life period of an element is also a measure of its radioactivity, since shorter the half life period, the greater is the number of disintegrations and hence greater its radioactivity. Half life periods vary from billions of years for some radioisotopes to a fraction of a second.

Half life period is represented as $t_{1/2}$.

Let the initial amount of a radioactive substance be N_0.

After one half life period ($t_{1/2}$), it becomes $= N_0/2$

After two half life periods ($2t_{1/2}$), it becomes $= N_0/4$

After three half life periods ($3t_{1/2}$), it becomes $= N_0/8$

and After n half life periods ($nt_{1/2}$), it shall become $= \left(\frac{1}{2}\right)^n N_0$

Thus, for the total disintegration of a radioactive substance an infinite time will be required.

Time (T)	Amount of radioactive substance (N)	Amount of radioactive substance decomposed ($N_0 - N$)
0	N_0	0
$t_{1/2}$	$\frac{1}{2}N_0 = \left(\frac{1}{2}\right)^1 N_0$	$\frac{1}{2}N_0 = \left[1-\frac{1}{2}\right]N_0$
$2t_{1/2}$	$\frac{1}{4}N_0 = \left(\frac{1}{2}\right)^2 N_0$	$\frac{3}{4}N_0 = \left[1-\frac{1}{4}\right]N_0$
$3t_{1/2}$	$\frac{1}{8}N_0 = \left(\frac{1}{2}\right)^3 N_0$	$\frac{7}{8}N_0 = \left[1-\frac{1}{8}\right]N_0$
$4t_{1/2}$	$\frac{1}{16}N_0 = \left(\frac{1}{2}\right)^4 N_0$	$\frac{15}{16}N_0 = \left[1-\frac{1}{16}\right]N_0$
$nt_{1/2}$	$\left(\frac{1}{2}\right)^n N_0$	$[1-\left(\frac{1}{2}\right)^n]N_0$

Amount of radioactive substance left after n half life periods

$$N = \left(\frac{1}{2}\right)^n N_0$$

and total time $T = n \times t_{1/2}$

where, n is a whole number.

SOME SOLVED EXAMPLES

Example 10. *The half life period of radium is 1580 years. How do you interpret this statement?*

Solution: Whatever quantity of radium is taken, it shall become half after the expiry of 1580 years. The following table explains the statement:

Quantity of radium at present	Quantity of radium after 1580 years
100 atoms	50 atoms
50 gram	25 gram
5 mole	2.5 mole

Example 11. *The radioactive isotope ^{137}Cs has a half life period of 30 years. Starting with 1 mg of ^{137}Cs, how much would remain after 120 years?*

Solution: At this time, we have 1.0 mg of ^{137}Cs; after 30 years, we shall have one half of the original, or 0.50 mg; after 60 years, we shall have 0.25 mg; after 90 years, we shall have 0.125 mg and, finally, after 120 years, we shall have 0.0625 mg.

After	30 years	—	0.50 mg
	60 years	—	0.25 mg
	90 years	—	0.125 mg
	120 years	—	0.0625 mg

Alternative solution: Total time = 120 years

We know that, total time $= n \times t_{1/2}$

So, $120 = n \times 30$

$n = 4$

Thus, the quantity of the isotope left after

four half life periods $= \left(\frac{1}{2}\right)^4 N_0 = \left(\frac{1}{2}\right)^4 \times 1$

$= \frac{1}{16} = 0.0625$ mg

Example 12. *A radioactive element has half life period of 30 days. How much of it will be left after 90 days?*

Solution: Total time = 90 days

Half life $(t_{1/2}) = 30$ days

We know that, total time $= n \times t_{1/2}$

So, $90 = n \times 30$

$n = 3$

Thus, quantity left after three half life periods

$$= (\frac{1}{2})^3 N_0 \quad [N_0 = \text{original amount}]$$

$$= \frac{1}{8} \times N_0 = \frac{1}{8} N_0$$

Example 13. *The half life period of $^{210}_{84}Po$ is 140 days. In how many days 1 g of this isotope is reduced to 0.25 g?*

Solution: Original quantity of the isotope $(N_0) = 1$ g

Final quantity of the isotope $N = 0.25$ g

We know that, $N = \left(\frac{1}{2}\right)^n N_0$

So, $\frac{1}{4} = \left(\frac{1}{2}\right)^n \times 1$

or $\left(\frac{1}{2}\right)^2 = \left(\frac{1}{2}\right)^n$

or $n = 2$

Time taken $T = n \times t_{1/2} = 2 \times 140 = 280$ days

Example 14. *The half life period of ^{234}U is 2.5×10^5 years. In how much time is the quantity of an isotope reduced to 25% of the original amount?*

Solution: Initial amount of this isotope $N_0 = 100$

Final amount of the isotope $N = 25$

We know that, $N = \left(\frac{1}{2}\right)^n N_0$

So, $25 = \left(\frac{1}{2}\right)^n \times 100$

or $\frac{25}{100} = \left(\frac{1}{2}\right)^n$

or $\frac{1}{4} = \left(\frac{1}{2}\right)^n$

or $\left(\frac{1}{2}\right)^2 = \left(\frac{1}{2}\right)^n$

or $n = 2$

Time taken $T = n \times t_{1/2}$

$= 2 \times 2.5 \times 10^5 = 5 \times 10^5$ years

$$\log \frac{N_0}{N} = \lambda t \text{ or } \log \frac{N}{N_0} = -\lambda t \text{ or } \frac{N}{N_0} = e^{-\lambda t} \text{ or } N = N_0 e^{-\lambda t}$$

Example 15. *A radioisotope has $t_{1/2} = 5$ years. After a given amount decays for 15 years, what fraction of the original isotope remains?*

Solution: Half life $(t_{1/2}) = 5$ years

Time for decay $(T) = 15$ years

We know that, $T = n \times t_{1/2}$

So, $15 = n \times 5$

or $n = 3$

Let the original amount be $= N_0$

Let the amount left after three half life periods be $= N$

fraction $= N / N_0$

We know that, $N = \left(\frac{1}{2}\right)^n N_0$

or $\frac{N}{N_0} = \left(\frac{1}{2}\right)^n = \left(\frac{1}{2}\right)^3 = \frac{1}{8}$

Thus, after 15 years $\frac{1}{8}$th of the original amount remains.

Example 16. *If in 3160 years, a radioactive substance becomes one-fourth of the original amount, find its half life period.*

Solution: $\frac{N}{N_0} = \frac{1}{4}$

So, $\frac{1}{4} = \left(\frac{1}{2}\right)^n$

or $\left(\frac{1}{2}\right)^2 = \left(\frac{1}{2}\right)^n$

or $n = 2$

We know that, total time 'T' $= n \times t_{1/2}$

So, $3160 = 2 \times t_{1/2}$

or $t_{1/2} = \frac{3160}{2} = 1580$ years

The half life period of the radioactive substance is 1580 years.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

7. Half life of a radioactive sample is $2x$ years. What fraction of this sample will remain undecayed after x years?

(a) $\frac{1}{2}$ (b) $\frac{1}{\sqrt{2}}$ (c) $\frac{1}{\sqrt{3}}$ (d) 2

[Ans. (b)]

[Hint: $\lambda = \frac{2.303}{t} \log\left(\frac{N_0}{N}\right)$

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{t} \log_{10}\left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{2x} = \frac{2.303}{x} \log_{10}\left(\frac{N_0}{N}\right)$$

$$\frac{1}{2} \log_{10} 2 = \log\left(\frac{N_0}{N}\right)$$

$$\frac{N}{N_0} = \frac{1}{\sqrt{2}}$$

Fraction undecayed $= \frac{1}{\sqrt{2}}$]

8. Half life of a radioactive element is 10 days. What percentage of the element will remain undecayed after 100 days?
(a) 10% (b) 0.1%
(c) 0% (d) 99%
[Ans. (b)]
[Hint: In ten times of half life 99.9%, the element undergoes decay; then percentage of undecayed radioactive element will be 0.1%.]

9. Which among the following relations is correct?
(a) $t_{1/2} = 2t_{3/4}$ (b) $t_{1/2} = 3t_{3/4}$
(c) $t_{3/4} = 2t_{1/2}$ (d) $t_{3/4} = 3t_{1/2}$
[Ans. (c)]
[Hint: $\frac{0.693}{t_{1/2}} = \frac{2.303}{t} \log\left(\frac{N_0}{N}\right)$

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{t_{3/4}} \log\frac{100}{75}$$

$t_{3/4} = 2t_{1/2}$]

10. Select the correct statement:
(a) Same amount will decay in every half life
(b) Amount decayed in first half life is maximum
(c) Amount decayed in first half life is minimum
(d) Amount decayed in a half life depends on the nature of element
[Ans. (b)]
[Hint: Amount decayed in first half life is maximum. Half of the initial amount is decayed in first half life.]

11. The half life period of a radioactive mineral is 15 min. What percent of radioactivity of that mineral will remain after 45 min? [UGET (Manipal Medical) 2006]
(a) 17.5% (b) 15% (c) 12.5% (d) 10%
[Ans. (c)]
[Hint: $n = \frac{45}{15} = 3$ = No. of half lives

$$N = N_0\left(\frac{1}{2}\right)^n = 100 \times \left(\frac{1}{2}\right)^3 = 12.5\%]$$

12. Half life of a radioactive element is 16 hrs. What time will it take for 75% disintegration? (DCE 2006)
(a) 32 days (b) 32 hrs (c) 48 hrs (d) 16 hrs
[Ans. (b)]
[Hint: 75% decay takes place in $t_{3/4}$ (3 / 4th life)

$$t_{3/4} = 2t_{1/2} = 2 \times 16 = 32 \text{ hrs}]$$

Disintegration constant: A chemical reaction whose rate of reaction varies directly as the concentration of one molecular species only, is termed a first order reaction. Radioactive disintegration is similar to such a chemical reaction as one radioactive species changes into other. This change can be represented by the equation:

$$A \longrightarrow B$$

Suppose the number of atoms of a radioactive substance present at the start of observation, *i.e.*, when $t = 0$, is N_0 and after time t, the number of atoms remaining unchanged is N. At this instant a very small number of atoms dN disintegrate in a small time dt; the rate of change of A into B is given by $-\frac{dN}{dt}$. The negative sign indicates that number of atoms decreases as time increases. Since, rate of disintegration or change is proportional to the total number of atoms present at that time, the relation becomes

$$-\frac{dN}{dt} = \lambda \cdot N \quad \text{... (i)}$$

'λ' is called the **disintegration constant** or **decay constant.**

Evidently $-\frac{dN}{N} = \lambda \cdot dt$ ... (ii)

If $dt = 1$ second, $\lambda = -\frac{dN}{N}$ ... (iii)

Thus, λ may be defined as **the fraction of the total number of atoms which disintegrate per second at any time.** This is constant for a given radioactive isotope.

Integrating eq. (ii), $-\int \frac{dN}{N} = \lambda \int dt$

$$-\log N = \lambda t + C \quad \text{... (iv)}$$

C is the integration constant.

When $t = 0$, $N = N_0$

Putting the values in eq. (iv),

$$-\log N_0 = C \quad \text{... (v)}$$

Putting the value of C in eq. (iv)

$$-\log N = \lambda t - \log N_0 \quad \text{or} \quad \log N_0 - \log N = \lambda t$$

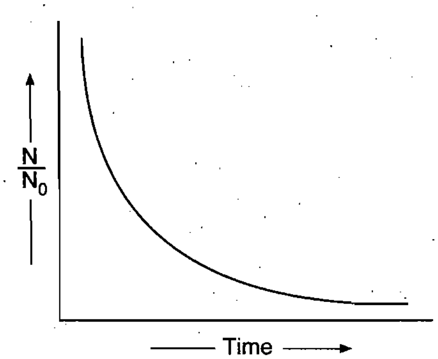


Fig. 3.5 Fraction of radioisotope remaining versus time

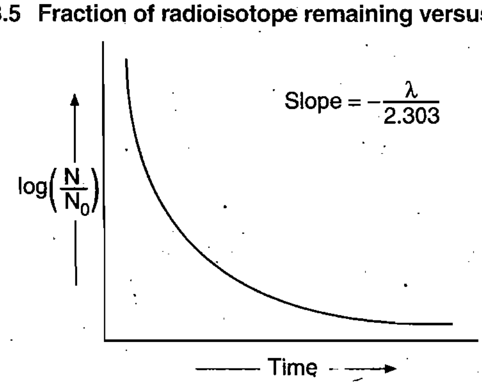


Fig. 3.6 Log of fraction remaining versus time

or $\log \frac{N_0}{N} = \lambda t$ or $2.303 \log_{10} \frac{N_0}{N} = \lambda t$

or
$$\lambda = \frac{2.303}{t} \log_{10} \frac{N_0}{N} \quad \text{... (vi)}$$

This equation is called kinetic equation and is obeyed by first order reactions.

Relationship between half life period and radioactive disintegration constant.

When $t = t_{1/2}$, $N = \frac{N_0}{2}$

Putting the values in eq. (vi)
$$\lambda = \frac{2.303}{t_{1/2}} \log_{10} \frac{N_0}{N_0/2} = \frac{2.303}{t_{1/2}} \log_{10} 2$$

The value of $\log_{10} 2$ is 0.3010.

So, $\lambda = \frac{0.693}{t_{1/2}}$ or $t_{1/2} = \frac{0.693}{\lambda}$

Thus, half life period of a given radioactive substance does not depend on the initial amount of a radioactive substance but depends only on the disintegration constant of the radioactive element.

3.10 AVERAGE LIFE

It is the sum of the periods of existence of all the atoms divided by the total number of atoms of the radioactive substance.

$$\text{Average life} = \frac{\text{Total life time of all the atoms}}{\text{Total number of atoms}}$$
$$= \frac{\int_0^\infty t\,dN}{N_0} = \frac{1}{\lambda}$$

Thus, **average life** of a radioactive element is the inverse of its disintegration or decay constant.
$$\text{Average life} = \frac{1}{\lambda} = \frac{t_{1/2}}{0.693} = 1.44 t_{1/2}$$

The **average life** of a radioactive substance is 1.44 times of its half life period.

Alternatively:

We know that, $\lambda = \frac{2.303}{t} \log_{10} \left(\frac{N_0}{N}\right)$

or $\lambda t = \log_e \left(\frac{N_0}{N}\right)$

$$e^{\lambda t} = \frac{N_0}{N}$$

or $\frac{N}{N_0} = e^{-\lambda t}$

Let $t = \frac{1}{\lambda}$; then $\frac{N}{N_0} = e^{-1} = \frac{1}{e}$

$$\frac{N}{N_0} = \frac{1}{2.718} = 0.3679$$

$$\%\text{ remaining amount} = \frac{N}{N_0} \times 100 = 36.79$$

$$\%\text{ decayed amount} = 100 - 36.79 = 63.21$$

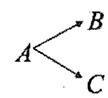
Time during which 63.21% substance undergoes decay is called average life.

Relation between rate of decay and mass of given element

$$\text{Rate}\left(-\frac{dN}{dt}\right) = \lambda \times N$$
$$= \lambda \times \text{No. of atoms of element undergoing decay}$$
$$= \frac{0.693}{t_{1/2}} \times \frac{\text{mass}}{\text{atomic mass}} \times \text{Avogadro's number}$$

Parallel Path Decay

Let a radioactive element 'A' decays to 'B' and 'C' in two parallel paths:



Decay constant of

'A' = Decay constant of 'B' + Decay constant of 'C'

$$\lambda_A = \lambda_B + \lambda_C \quad \text{... (i)}$$

Here, $\lambda_B = [\text{fractional yield of } B] \times \lambda_A$

$\lambda_C = [\text{fractional yield of } C] \times \lambda_A$

Maximum Yield of Daughter Element

Let a radioactive element 'A' decays to daughter element 'B'.
$$A \longrightarrow B$$

λ_A and λ_B are decay constants of 'A' and 'B'. Maximum activity time of daughter element can be calculated as:
$$t_{max} = \frac{2.303}{(\lambda_B - \lambda_A)} \log_{10}\left[\frac{\lambda_B}{\lambda_A}\right]$$

3.11 RADIOACTIVE EQUILIBRIUM

Let us consider that a radioactive element A disintegrates to give B which is also radioactive and disintegrates into C.
$$A \longrightarrow B \longrightarrow C$$

The element B is said to be in radioactive equilibrium with A if its rate of formation from A is equal to its rate of decay into C. If λ_1 and λ_2 are the disintegration constants of A and B, N_1 and N_2 are the number of atoms of each radioactive element present at equilibrium, then we have

Rate of formation of B = Rate of decay of $A = \lambda_1 N_1$

and Rate of decay of $B = \lambda_2 N_2$

At radioactive equilibrium,
$$\lambda_1 N_1 = \lambda_2 N_2$$

or
$$\frac{N_1}{N_2} = \frac{\lambda_2}{\lambda_1} = \frac{1/\lambda_1}{1/\lambda_2} = \frac{\text{Average life of A}}{\text{Average life of B}} = \frac{Z_A}{Z_B}$$

Thus, the number of atoms of A and B are in the ratio of their average life periods.

$$\frac{N_1}{N_2} = \frac{\lambda_2}{\lambda_1} = \frac{0.693/(t_{1/2})_2}{0.693/(t_{1/2})_1}$$
$$\frac{N_1}{N_2} = \frac{(t_{1/2})_1}{(t_{1/2})_2}$$

When λ_A of parent element is less than λ_B of daughter element, but both are not very small then a *transient equilibrium* is reached, when

$$\frac{N_B}{N_A} = \frac{\lambda_A}{\lambda_B - \lambda_A}$$

in fact it is steady state.

3.12 UNITS OF RADIOACTIVITY

In radioactivity, the number of atoms which disintegrate in unit time is of real importance rather than the total amount of the radioactive substance expressed by mass or number of atoms, *i.e.*, the activity of a radioactive substance is the rate of decay or number of disintegrations per second.

The unit of radioactivity called Curie (Ci) **is defined as that quantity of any radioactive substance which has a decay rate of 3.7×10^{10} disintegrations per second.**

This unit is a large one and hence smaller units like **millicurie** (mCi) and **microcurie** (μCi) are used.

$$1 \text{ millicurie} = 3.7 \times 10^7 \text{ disintegrations per sec}$$

$$1 \text{ microcurie} = 3.7 \times 10^4 \text{ disintegrations per sec}$$

There is another unit Rutherford (Rd) which is also used these days. **It is defined as the amount of a radioactive substance which undergoes 10^6 disintegrations per second.** Smaller units like milli-Rutherford and micro-Rutherford are also used.

$$1 \text{ milli-Rutherford} = 10^3 \text{ disintegrations per sec (dps)}$$

$$1 \text{ micro-Rutherford} = 1 \text{ disintegration per sec}$$

The SI unit of radioactivity is proposed as Becquerel which refers to one dps.

$$1 \text{ curie} = 3.7 \times 10^4 \text{ Rutherford} = 3.7 \times 10^{10} \text{ Becquerel}$$

$$1 \text{ curie} = 37 \text{ GBq}$$

Here, G stands for 10^9, *i.e.*, giga

Gray (Gy) = 1 kg tissue receiving 1 J of energy

Sievert (Sv) = gray × quality number of radiation

Quality number of 1α-particle = 20

Quality number of 1β-particle = 1

Specific activity of a radionuclide is its activity per kilogram (or dm^3) of the radioactive material.

(In some cases, specific activity is taken as the activity per gram.)

Radiation counter: There are two main radiation counters in practice.

1. Geiger-Muller counter: It is used to count charged particles, *e.g.*, α and β-particles, emitted by a radioactive nucleus. This counter is simply a metal tube filled with a gas like argon.

In order to count and detect neutrons, boron trifluoride (BF_3) is added along with a gas in the G.M. counter. Neutron strikes ${}^{10}_{5}B$ nuclei to produce α-particle, which is then detected and counted in Geiger counter.

$${}^{10}_{5}B + {}^{1}_{0}n \longrightarrow {}^{7}_{3}Li + {}^{4}_{2}He$$

2. Scintillation counter: γ-radiations are detected by Scintillation counter. A phosphor is used in this counter which produces flash of light when it is struck by electromagnetic radiation like γ-rays, for detection of γ-rays. Sodium iodide (NaI) and thallium iodide (TlI) are used as phosphor. Rutherford first of all used zinc sulphide (ZnS) as phosphor in detection of α-particles.

SOME SOLVED EXAMPLES

Example 17. *The half life period of radium is 1600 years. Calculate the disintegration constant of radium. Mention its unit.*

Solution: Disintegration constant $\lambda = \frac{0.693}{t_{1/2}}$

Since, $t_{1/2} = 1600$ years

So, $\lambda = \frac{0.693}{1600}$

or $\lambda = 4.33 \times 10^{-4}\ \text{year}^{-1}$

Example 18. *The disintegration constant of ${}^{238}U$ is 1.54×10^{-10} year^{-1}. Calculate the half life period of ${}^{238}U$.*

Solution: Half life period, $t_{1/2} = \frac{0.693}{\lambda}$

Since, $\lambda = 1.54 \times 10^{-10}\ \text{year}^{-1}$

So, $t_{1/2} = \frac{0.693}{1.54 \times 10^{-10}} = 4.5 \times 10^9$ years

Example 19. *The half life period of radon is 3.8 days. After how many days will only one-twentieth of radon sample be left over?*

Solution: We know that, $\lambda = \frac{0.693}{t_{1/2}} = \frac{0.693}{3.8} = 0.182\ \text{day}^{-1}$

Let the initial amount of radon be N_0 and the amount left after t days be N which is equal to $\frac{N_0}{20}$.

Applying the equation,

$$t = \frac{2.303}{\lambda} \log_{10} \frac{N_0}{N}$$

$$= \frac{2.303}{0.182} \log_{10} \frac{N_0}{N_0/20} = \frac{2.303}{0.182} \log_{10} 20$$

$$= 16.54 \text{ days}$$

Example 20. *A counter rate meter is used to measure the activity of a radioactive sample. At a certain instant, the count rate was recorded as 475 counters per minute. Five minutes later, the count rate recorded was 270 counts per minute. Calculate the decay constant and half life period of the sample.*

Solution: Let N_0 and N be the number of atoms of the radioactive substance present at the start and after 5 minutes respectively.

Rate of disintegration at the start $= \lambda N_0 = 475$

and rate of disintegration after 5 minutes $= \lambda N = 270$

Dividing both, $\frac{\lambda N_0}{\lambda N} = \frac{475}{270}$

or $\frac{N_0}{N} = 1.76$

We know that, $\lambda = \frac{2.303}{t} \log_{10} \frac{N_0}{N}$

$$\lambda = \frac{2.303}{5} \log_{10} 1.76 = 0.113 \text{ minute}^{-1}$$

$$\text{Half life period} = \frac{0.693}{\lambda} = \frac{0.693}{0.113} = 6.1 \text{ minutes}$$

Example 21. *You have* 0.1 g *atom of a radioactive isotope* $^A_Z X$ *(half life* = 5 *days). How many atoms will decay during the* 11*th day?*

Solution: Amount of radioactive substance = 0.1 g atom

So, $N_0 = 0.1 \times$ Avogadro's number

$$= 0.1 \times 6.02 \times 10^{23}$$

$$= 6.02 \times 10^{22} \text{ atoms}$$

$$\text{Number of atoms after 5 days} = \frac{6.02 \times 10^{22}}{2} = 3.01 \times 10^{22}$$

$$\text{Number of atoms after 10 days} = \frac{3.01 \times 10^{22}}{2} = 1.505 \times 10^{22}$$

Let the number of atoms left after 11 days be N.

We know that,

$$t = \frac{2.303}{\lambda} \log_{10} \frac{N_0}{N}$$

Given, $t = 11$, $\lambda = \frac{0.693}{5}$, $N_0 = 6.02 \times 10^{22}$

So, $$11 = \frac{2.303 \times 5}{0.693} \log_{10} \frac{6.02 \times 10^{22}}{N}$$

or $$\log_{10} \frac{6.02 \times 10^{22}}{N} = \frac{11 \times 0.693}{2.303 \times 5} = 0.6620$$

$$\frac{6.02 \times 10^{22}}{N} = \text{Antilog } 0.6620 = 4.592$$

So, $$N = \frac{6.02}{4.592} \times 10^{22} = 1.3109 \times 10^{22}$$

Atoms decayed during 11th day

$$= [1.5050 \times 10^{22} - 1.3109 \times 10^{22}]$$

$$= 0.1941 \times 10^{22}$$

$$= 1.941 \times 10^{21}$$

Example 22. 10 *g atoms of an* α*-active radioisotope are disintegrating in a sealed container. In one hour, the helium gas collected at STP is* 11.2 cm^3. *Calculate the half life of the radio isotope.*

Solution: Amount of radioactive isotope = 10 g atoms

or $N_0 = 10 \times 6.023 \times 10^{23}$ atoms

$$= 6.023 \times 10^{24} \text{ atoms}$$

22400 cm^3 of helium contains = 6.023×10^{23} atoms

$$11.2 \text{ cm}^3 \text{ of helium will contain} = \frac{6.023 \times 10^{23}}{22400} \times 11.2 \text{ atoms}$$

$$= 3.01 \times 10^{20} \text{ atoms}$$

As one helium atom is obtained by disintegration of one atom of radioisotope, the total number of atoms of the radioactive isotope which have disintegrated in one hour

$$= 3.01 \times 10^{20} \quad \text{or} \quad 0.0003 \times 10^{24}$$

The number of atoms of the radioactive isotope left after one hour,

$$N = (6.023 \times 10^{24} - 0.000301 \times 10^{24})$$

$$= 6.0227 \times 10^{24}$$

Using, $$\lambda = \frac{2.303}{t} \log \frac{N_0}{N}$$

$$\lambda = \frac{2.303}{t} \log \frac{6.023 \times 10^{24}}{6.0227 \times 10^{24}}$$

$$= 2.303 \times 2.1632 \times 10^{-5} = 4.982 \times 10^{-5} \text{ hr}^{-1}$$

$$t_{1/2} = 0.693 / (4.982 \times 10^{-5} \times 24 \times 365) = 1.58 \text{ years}$$

Example 23. *Calculate the average life of a radioactive substance whose half life period is* 1650 *years.*

Solution: Average life = $1.44 \times t_{1/2}$

$$= 1.44 \times 1650 = 2376 \text{ years}$$

Example 24. ^{90}Sr *shows* β*-activity and its half life period is* 28 *years. What is the activity of a sample containing* 1 *g of* ^{90}Sr?

Solution: Activity = No. of atoms disintegrating per second

$$= \lambda \times \text{total number of atoms}$$

$$\lambda = \frac{0.693}{28 \times 365 \times 24 \times 60 \times 60}$$

$$\text{Total number of atoms in 1 g of } ^{90}\text{Sr} = \frac{6.023 \times 10^{23}}{90}$$

$$\text{Activity} = \frac{0.693}{28 \times 365 \times 24 \times 60 \times 60} \times \frac{6.023 \times 10^{23}}{90}$$

$$= 5.25 \times 10^{12} \text{ disintegrations per second}$$

$$= \frac{5.25 \times 10^{12}}{3.7 \times 10^{10}} = 141.89 \text{ curie}$$

Example 25. *A chemist prepares* 1.00 *g of pure* $^{11}_6C$. *This isotope has half life of* 21 *minutes, decaying by the equation:*

$$^{11}_6C \rightarrow {}^{11}_5B + {}^0_1e$$

(*a*) *What is the rate of disintegration per second (dps) at start?*

(*b*) *What are the activity and specific activity of* $^{11}_6C$ *at start?*

(*c*) *How much of this isotope* ($^{11}_6C$) *is left after 24 hours of its preparation?*

Solution: (a) Applying, $-\frac{dN}{dt} = \lambda N_0$

$$= \frac{0.693}{21 \times 60} \times \frac{1 \times 6.02 \times 10^{23}}{11}$$

$$= 3 \times 10^{19} \text{ dps}$$

(b) Activity $= \frac{3\times 10^{19}}{3.7\times 10^{10}}$ (1 curie $= 3.7\times 10^{10}$ dps)

$= 8.108\times 10^{8}$ curie

Sp. activity $= 3\times 10^{19}\times 10^{3} = 3\times 10^{22}$ dis/kg s

$= 8.108\times 10^{11}$ curie

(c) Applying, $N = N_0\left(\frac{1}{2}\right)^n$ $\left[n = \frac{t}{t_{1/2}} = \frac{24\times 60}{21} = 68.57\right]$

$$N = 1\times\left(\frac{1}{2}\right)^{68.57} = 2.29\times 10^{-21}\text{ g}$$

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

13. The time of decay for the nuclear reaction is given by $t = 5t_{1/2}$. The relation between mean life τ and time of decay 't' is given by:

(a) $2\tau \ln 2$ (b) $5\tau \ln 2$ (c) $2\tau^4 \ln 2$ (d) $\frac{1}{\tau^4}\ln 2$

[Ans. (b)]

[Hint: $t = 5t_{1/2}$

$t = 5\times\frac{\ln 2}{\lambda}$

$t = 5\tau \ln 2$]

14. The activity of a sample of radioactive element ^{100}A is 6.02 curie. Its decay constant is $3.7\times 10^4\ s^{-1}$. The initial mass of the sample will be:

(a) 10^{-14} g (b) 10^{-6} g (c) 10^{-15} g (d) 10^{-3} g

[Ans. (c)]

[Hint: Activity $= \lambda\times\frac{w}{\text{At. wt.}}\times 6.023\times 10^{23}$

$6.02\times 3.7\times 10^{10} = 3.7\times 10^4\times\frac{w}{100}\times 6.023\times 10^{23}$

$w = 10^{-15}$ g]

15. A freshly prepared radio medicine has half life 2 hours. Its activity is 64 times the permissible safe value. The minimum time after which it would be possible to treat the patients with the medicine is:

(a) 3 hrs (b) 9 hrs (c) 24 hrs (d) 12 hrs

[Ans. (d)]

[Hint: $N = N_0\left(\frac{1}{2}\right)^n$

$\frac{N}{N_0} = \left(\frac{1}{2}\right)^n$

$\frac{1}{64} = \left(\frac{1}{2}\right)^n$; $n = 6$ half lives

$\therefore$ time $= 2\times 6 = 12$ hrs]

16. One gram of ^{226}Ra has an activity of nearly 1 Ci. The half life of ^{226}Ra is:

(a) 1582 yrs (b) 12.5 hrs
(c) 140 days (d) 4.5×10^9 yrs

[Ans. (a)]

[Hint: Use the following relation for calculation of activity:

$$\text{Activity} = \frac{0.693}{t_{1/2}}\times\frac{w}{\text{At. wt.}}\times 6.023\times 10^{23}$$

$$3.7\times 10^{10} = \frac{0.693}{t_{1/2}}\times\frac{1}{226}\times 6.023\times 10^{23}$$

It will give the half life in seconds.]

17. Assuming that ^{226}Ra ($t_{1/2} = 1.6\times 10^3$ yrs) is in secular equilibrium with ^{238}U ($t_{1/2} = 4.5\times 10^9$ yrs) in a certain mineral, how many grams of radium will be present in for every gram of ^{238}U in this mineral?

(a) 3.7×10^{-7} (b) 3.4×10^{7}
(c) 3.4×10^{-7} (d) 3.7×10^{7}

[Ans. (c)]

[Hint: $\frac{N_1\ {}^{226}\text{Ra}}{N_2\ {}^{238}\text{U}} = \frac{t_{1/2}\ {}^{226}\text{Ra}}{t_{1/2}\ {}^{238}\text{U}}$

$\frac{w/226}{1/238} = \frac{1.6\times 10^3}{4.5\times 10^9}$; $w = 3.4\times 10^{-7}$ g]

18. A certain radioactive isotope decay has α-emission,

$${}^{A_1}_{Z_1}X \longrightarrow {}^{A_1-4}_{Z_1-2}Y$$

half life of X is 10 days. If 1 mol of X is taken initially in a sealed container, then what volume of helium will be collected at STP after 20 days?

(a) 22.4 L (b) 11.2 L (c) 16.8 L (d) 33.6 L

[Ans. (c)]

[Hint: After 20 days 0.75 mol helium will be formed.

$\therefore$ Volume of helium at STP $= 0.75\times 22.4$

$= 16.8$ L]

3.13 ARTIFICIAL TRANSMUTATION

Transmutation is defined as the conversion of one element into another or one type of atom into another. When this conversion is achieved by artificial means, it is termed as artificial transmutation.

The conversion of elements into one another has been the dream of the human race for many centuries. In the middle ages, it was popular under the name of '*Alchemy*'. Alchemists were unsuccessful in this attempt as they were having very little knowledge about the structure of atom. With the background of the clear picture of the structure of the atom, modern scientists have realised that to convert one element into another, the nucleus should be attacked and altered.

The first indication that a stable nucleus could be disrupted was given by Rutherford in 1919. He observed that when nitrogen was bombarded with high speed α-particles from ^{214}Po, protons were emitted. Thus, nitrogen was changed into an isotope of oxygen.

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

Later on, Rutherford and Chadwick showed that many other elements from boron to potassium with the exception of carbon and oxygen could be transmuted by bombardment with α-particles. However with heavier elements, there was only scattering of α-particles as these suffered a force of repulsion. It was, thus, concluded that to bring transmutation in heavier elements, the projectiles must have higher energies than α-particles obtained from natural sources. It was suggested by Gamow in 1928 that a proton (^1_1H) would be a much more effective projectile than an α-particle, but it was not available as a high speed particle.

The charged particles, like alpha particles, protons, deuterons can be made much more effective projectiles if they have high velocity. Out of all the instruments which have been devised for accelerating projectiles, the one which has attracted the widest interest is the cyclotron of E.O. Lawrence. The projectile can be accelerated to the speed of 25,000 miles per second.

The discovery of neutron by Chadwick, in 1932, added another projectile for transmutation. The neutron being electrically neutral can penetrate easily into the atomic nucleus. Although neutrons are the most effective and versatile of projectiles, yet they suffer the objection that they must be produced by transmutation at the time of use. High speed neutrons are obtained when beryllium-9 is bombarded with α-particles,

$$^9_4Be + ^4_2He \longrightarrow ^{12}_6C + ^1_0n$$

and slow neutrons are obtained by bombarding lithium-7 with protons.

$$^7_3Li + ^1_1H \longrightarrow ^7_4Be + ^1_0n$$

In general, for the transmutation of lighter elements, charged particles like alpha particles, protons, deuterons are used while for heavier elements, neutrons are used.

Nuclear Reactions

The reactions in which nuclei of atoms interact with other nuclei or elementary particles such as alpha particle, proton, deuteron, neutron, etc., resulting in the formation of a new nucleus and one or more elementary particles are called nuclear reactions. Nuclear reactions are expressed in the same fashion as chemical reactions, *i.e.*, reactants on left hand side and the products on right hand side of the sign of (=) or ($\rightarrow$). **In all nuclear reactions, the total number of protons and neutrons are conserved as in chemical reactions, the number of atoms of each element are conserved.** The symbols ^1_0n, ^1_1H, ^4_2He, ^2_1H, $^0_{-1}e$, $^0_{+1}e$ and γ are used to represent neutron, proton, α-particle, deuteron, electron, positron, γ-rays respectively. A short hand notation is often used for the representation of nuclear reactions. As for example, the nuclear reaction

$$^{14}_7N + ^4_2He \longrightarrow ^{17}_8O + ^1_1H$$

is represented as $^{14}_7N(\alpha, p)\,^{17}_8O$. Some of the characteristics that differentiate between nuclear reactions and ordinary chemical reactions are summarised below:

Nuclear reactions	Chemical reactions
1. Elements may be converted from one to another.	No new element can be produced.
2. Particles within the nucleus are involved.	Only outermost electrons participate.
3. Often accompanied by release or absorption of tremendous amount of energy.	Accompanied by release or absorption of relatively small amount of energy.
4. Rate of reaction is independent of external factors such as temperature, pressure and catalyst.	Rate of reaction is influenced by external factors.

Example 26. *Calculate the energy in the reaction*

$$2\,^1_1H + 2\,^1_0n \rightarrow ^4_2He$$

Given, H = 1.00813 amu, n = 1.00897 amu and He = 4.00388 amu

Solution: Loss of mass in the given nuclear reaction

$$= 2(1.00813 + 1.00897) - 4.00388$$

$$= 0.03032 \text{ amu}$$

Energy released = $0.03032 \times 931 = 28.3$ MeV

Types of Nuclear Reactions

(a) Projectile capture reactions: The bombarding particle is absorbed with or without the emission of γ-radiations.

$$^{238}_{92}U + ^1_0n \longrightarrow ^{239}_{92}U + \gamma$$

$$^{27}_{13}Al + ^1_0n \longrightarrow ^{28}_{13}Al + \gamma$$

(b) Particle-particle reactions: Majority of nuclear reactions come under this category. In addition to the product nucleus, an elementary particle is also emitted.

$$^{23}_{11}Na + ^1_1H \longrightarrow ^{23}_{12}Mg + ^1_0n$$

$$^{23}_{11}Na + ^2_1H \longrightarrow ^{24}_{11}Na + ^1_1H$$

$$^{23}_{11}Na + ^4_2He \longrightarrow ^{26}_{12}Mg + ^1_1H$$

$$^{14}_7N + ^1_0n \longrightarrow ^{14}_6C + ^1_1H$$

(c) Spallation reactions: High speed projectiles with energies approximately 40 MeV may chip fragments from a heavy nucleus, leaving a smaller nucleus.

$$^{63}_{29}Cu + ^4_2He + 400\text{ MeV} \longrightarrow ^{37}_{17}Cl + 14\,^1_1H + 16\,^1_0n$$

$$^{75}_{33}As + ^2_1H \longrightarrow ^{56}_{25}Mn + 9\,^1_1H + 12\,^1_0n$$

$$^{238}_{92}U + ^4_2He \longrightarrow ^{187}_{74}W + 20\,^1_1H + 35\,^1_0n$$

(d) Fission reactions: A reaction in which a heavy nucleus is broken down into two or more medium heavy fragments. The process is usually accompanied with emission of neutrons and large amount of energy.

$$^{235}_{92}U + ^1_0n \longrightarrow ^{141}_{56}Ba + ^{92}_{36}Kr + 3\,^1_0n + 200\text{ MeV}$$

(e) Fusion reactions: Light nuclei fuse together to reproduce comparatively heavier nuclei.

$$^2_1H + ^3_1H \longrightarrow ^4_2He + ^1_0n + 17.6\text{ MeV}$$

A fusion reaction is the source of tremendous amount of energy.

Pair production : Pair production is the most striking example of mass-energy equivalence.

We can write pair production symbolically as :

$$\text{Photon} + \text{Photon} \longrightarrow \text{Particle} + \text{Antiparticle}$$
$$\text{Particle} + \text{Antiparticle} \longrightarrow \text{Photon} + \text{Photon}$$

A particle and antiparticle can collide and annihilate each other, producing two high-energy gamma ray photons. Pair production must obey the law of conservation of energy and momentum.

The following are the important contributions of artificial transmutation:

(i) Discovery of neutron
(ii) Artificial radioactivity
(iii) Nuclear fission
(iv) Nuclear fusion

3.14 ARTIFICIAL RADIOACTIVITY

In 1934, Irene Curie and F. Joliot observed that when boron and aluminium were bombarded by α-particles, neutrons, protons and positrons were emitted. When bombardment was stopped, the emission of protons and neutrons ceased but that of positrons did not. The emission of positrons continued with time but decreased exponentially in a manner similar to natural radioactivity. Curie and Joliot explained this observation by saying that during bombardment, a metastable isotope is formed which behaves as a radioactive element. This process was termed as **artificial radioactivity.**

"The process in which a stable isotope is converted into a radioactive element by artificial transmutation is called artificial radioactivity."

When $^{27}_{13}\text{Al}$ is bombarded by α-particles, radioactive isotope $^{30}_{15}\text{P}$ is formed.

$$^{27}_{13}\text{Al} + {}^{4}_{2}\text{He} \longrightarrow {}^{30}_{14}\text{Si} + {}^{1}_{1}\text{H} \text{ (95\% of total conversion)}$$
$$\longrightarrow {}^{30}_{15}\text{P}^{*} + {}^{1}_{0}n \text{ (5\% of total conversion)}$$
$$\longrightarrow {}^{30}_{14}\text{Si} + {}^{0}_{+1}e$$

Positron

In a similar manner, the artificial radioactivity was observed when $^{10}_{5}\text{B}$ was bombarded by α-particles.

$$^{10}_{5}\text{B} + {}^{4}_{2}\text{He} \longrightarrow {}^{13}_{6}\text{C} + {}^{1}_{1}\text{H}$$
$$\longrightarrow {}^{13}_{7}\text{N}^{*} + {}^{1}_{0}n$$
$$\longrightarrow {}^{13}_{6}\text{C} + {}^{0}_{+1}e$$

The following are some of the nuclear reactions in which radioactive isotopes are formed.

$$^{23}_{11}\text{Na} + {}^{2}_{1}\text{H} \longrightarrow {}^{24}_{11}\text{Na}^{*} + {}^{1}_{1}\text{H} \quad [{}^{24}_{11}\text{Na} — \beta \text{ radioactive}]$$
$$^{238}_{92}\text{U} + {}^{1}_{0}n \longrightarrow {}^{239}_{92}\text{U}^{*} + \gamma \quad [{}^{239}_{92}\text{U} — \beta \text{ radioactive}]$$
$$^{12}_{6}\text{C} + {}^{1}_{1}\text{H} \longrightarrow {}^{13}_{7}\text{N}^{*} + \gamma \quad [{}^{13}_{7}\text{N} — \text{positron radioactive}]$$
$$^{25}_{12}\text{Mg} + {}^{4}_{2}\text{He} \longrightarrow {}^{28}_{13}\text{Al}^{*} + {}^{1}_{1}\text{H} \quad [{}^{28}_{13}\text{Al} — \beta \text{ radioactive}]$$

* Half life period of $^{30}_{15}\text{P}$ is 3.2 minutes.

3.15 NUCLEAR FISSION

"The process of artificial transmutation in which heavy nucleus is broken down into two lighter nuclei of nearly comparable masses with release of large amount of energy is termed nuclear fission." The word fission is derived from its resemblance to the biological process called fission in which a living cell breaks up into two cells of roughly same size.

After the discovery of neutron, **Fermi,** in 1934, made an attempt to synthesise transuranic elements from uranium by bombarding with neutrons. This experiment was repeated in Germany by **Hahn** and **Strassmann.** In one of the chemical tests, they found that one of the products was an isotope of barium along with the formation of an isotope of the element with atomic number 93 (neptunium). In 1939, they proposed that uranium after capturing neutron undergoes two types of reactions—one with ^{238}U isotope and the other with ^{235}U isotope.

(a) $^{238}_{92}\text{U}$ is converted into $^{239}_{93}\text{Np}$ and $^{239}_{94}\text{Pu}$.

$$^{238}_{92}\text{U} + {}^{1}_{0}n \longrightarrow {}^{239}_{92}\text{U} \xrightarrow{-\beta} {}^{239}_{93}\text{Np} \xrightarrow{-\beta} \underset{\text{(Plutonium)}}{{}^{239}_{94}\text{Pu}}$$

(b) $^{235}_{92}\text{U}$ captures slow neutron and splits up into fragments.

$$^{235}_{92}\text{U} + {}^{1}_{0}n \longrightarrow {}^{236}_{92}\text{U} \longrightarrow {}^{144}_{56}\text{Ba} + {}^{90}_{36}\text{Kr} + 2\,{}^{1}_{0}n$$

It has been observed that during fission of $^{235}_{92}\text{U}$ not only isotopes of Ba and Kr are formed but isotopes of various other elements come into existence. These isotopes fall under two groups. First type—isotopes having atomic masses from 80 to 110 and atomic numbers from 35 to 43 and second type—isotopes having atomic masses from 120 to 150 and atomic numbers 51 to 57. It is believed that only two isotopes are first formed as primary fission products which then give rise to secondary products by successive disintegration.

$$\left.\begin{aligned} {}^{235}_{92}\text{U} + {}^{1}_{0}n &\longrightarrow {}^{141}_{56}\text{Ba} + {}^{92}_{36}\text{Kr} + 3\,{}^{1}_{0}n \\ &\longrightarrow {}^{139}_{54}\text{Xe} + {}^{95}_{38}\text{Sr} + 2\,{}^{1}_{0}n \\ &\longrightarrow {}^{140}_{54}\text{Xe} + {}^{95}_{38}\text{Sr} + {}^{1}_{0}n \end{aligned}\right\} \text{Primary products}$$

$$\underbrace{{}^{140}_{54}\text{Xe} \xrightarrow{-\beta} {}^{140}_{55}\text{Cs} \xrightarrow{-\beta} {}^{140}_{56}\text{Ba} \xrightarrow{-\beta} {}^{140}_{57}\text{La} \xrightarrow{-\beta} {}^{140}_{58}\text{Ce}}_{\text{Secondary products}}$$

During fission, there is always loss of mass which is converted into energy according to Einstein equation $E = mc^2$. There is a loss of about 0.215 amu mass during one fission. Thus, energy released in one fission is equal to 0.215×931, *i.e.*, 200 MeV.

Chain reaction: Whatever are the primary products of fission of uranium, it is certain that neutrons are always set free. If the conditions are so arranged that each of these neutrons can, in turn, bring about the fission, the number of neutrons will increase at a continuously accelerating rate until whole of the material is exhausted. Such type of reaction is called **chain reaction.** It takes very small time and is uncontrolled. It ends in a terrible explosion due to release of enormous amount of energy. The chain reaction is shown in Fig. 3.7.

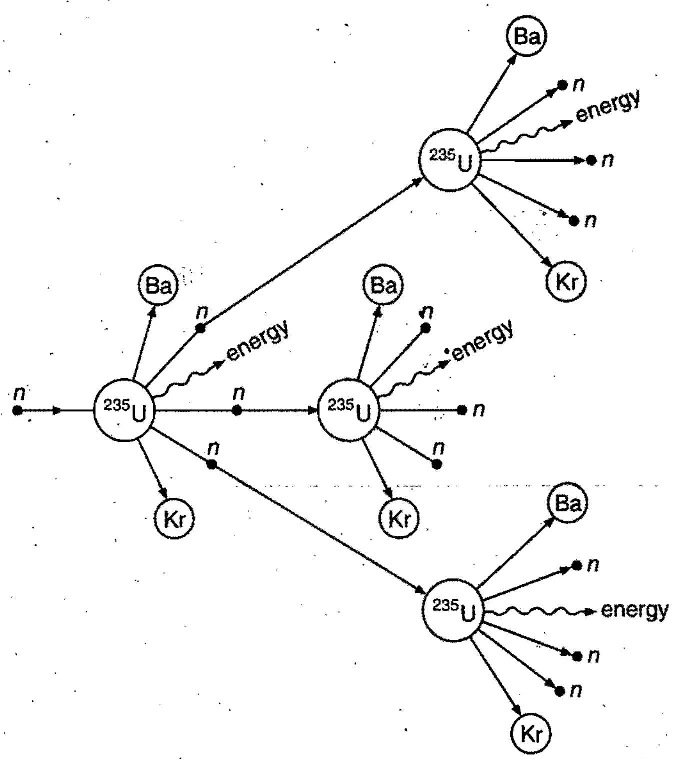


Fig. 3.7 Chain reaction in uranium-235

The chain reaction is self-propagating if the value of multiplication factor is more than 1.

Multiplication factor,

$$K = \frac{\text{No. of neutrons produced in one step}}{\text{No. of neutrons produced in preceding step}}$$

The value of K is 2.5 for ^{235}U and 0.5 for ^{238}U. This shows that if all other factors are ignored, natural uranium which is a mixture of three isotopes ($^{238}U = 99.29\%$, $^{235}U = 0.7\%$, $^{234}U = 0.0006\%$) is not suitable for a chain reaction. The following two factors hinder the self-propagation of a chain reaction:

(i) Leakage of neutrons from the system.

(ii) Presence of non-fissionable material.

If the system is such that loss of neutrons is more than their production, it is a subcritical stage. When the loss of neutrons is equal to their production, it is said to be the critical stage and when loss of neutrons is less than their production, it is the over-critical stage. Over-critical stage is required for self-propagation of chain reaction. The leakage of neutrons from the system can be reduced by suitable choice of size and shape of the fissionable material. The second source of loss of neutrons is due to absorption of neutrons by non-fissionable material. It may be reduced by careful purification of natural uranium, *i. e.*, natural uranium is submitted to the process of enrichment by which the percentage of ^{235}U in the sample is increased. The chain reaction can be carried out under two conditions: (a) uncontrolled (atom bomb) and (b) controlled (nuclear reactors).

Nuclear fuels: Nuclear fuels are of two types:

(i) Fissile materials: These, on bombardment with slow neutrons, directly produce a chain reaction leading to release of energy. Three fissile materials are in use at present. These are ^{235}U, ^{239}Pu and ^{233}U. ^{235}U is obtained from natural sources while ^{239}Pu and ^{233}U are obtained by artificial transmutation.

(ii) Fertile materials: A fertile material is one which by itself is non-fissile in nature, can be converted into a fissile material by reaction with neutrons. ^{238}U and ^{232}Th are fertile materials. ^{238}U is converted into ^{239}Pu by the following nuclear reaction:

$$^{238}_{92}U + ^{1}_{0}n \longrightarrow ^{239}_{92}U^* \xrightarrow[(23\text{ min})]{-\beta} ^{239}_{93}Np^* \xrightarrow[(2.3\text{ days})]{-\beta} ^{239}_{94}Pu$$

Similarly ^{232}Th is converted into ^{233}U.

$$^{232}_{90}Th + ^{1}_{0}n \longrightarrow ^{233}_{90}Th^* \xrightarrow[(23\text{ min})]{-\beta} ^{233}_{91}Pa^* \xrightarrow[(27\text{ days})]{-\beta} ^{233}_{92}U$$

Applications of nuclear fission: Three practical applications of nuclear fission are:

(a) Atomic bomb, (b) Nuclear reactor and (c) Power plant.

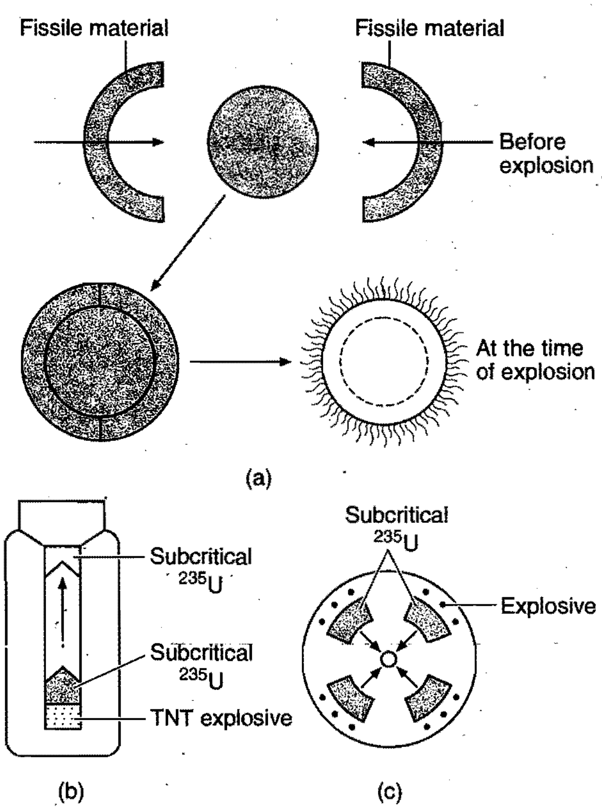


Fig. 3.8 (a), (b) and (c) various designs used in the assembly of atom bomb. (The atom bomb is made in two or more pieces of the fissile material each smaller than the critical size. The moment these pieces are forced together, the bomb explodes with terrific violence).

(a) Atomic bomb: It is based on uncontrolled chain reaction. The shape and size of the fissionable material is so adjusted at the time of explosion that it reaches the over-critical stage. In the atom bomb, a few pounds of fissionable material (^{235}U or ^{239}Pu) is taken in the form of a number of separate pieces; each piece is in subcritical stage (surface area is very large, *i.e.*, loss of neutrons is high.) At the time of explosion, these pieces are driven together rapidly by using explosives like TNT (trinitro toluene) lying behind each of ^{235}U pieces as to make one large piece of fissionable material. At this instant, the over-critical stage is achieved and a fast chain reaction is set- up. This results in a violent explosion with the release of tremendous amount of energy. Fig. 3.8 shows some of the designs of atomic bomb.

On account of explosion, the fragments fly apart with tremendous speeds. These collide with each other and kinetic energy is changed to heat energy. The amount of energy liberated in an atomic explosion is of the order of the detonation of about 20,000 or 30,000 tons of TNT raising the temperature to about 10^7°C. Air expands suddenly and a shock wave of great destructive impulse travels across. The explosion also produces a violent and intense blast of highly penetrating γ-rays which are exceedingly dangerous. The radioactive dust (fallout) scatters over wide areas causing contamination.

The first atomic bomb dropped over Hiroshima city during the second World War in 1945 utilised ^{235}U and the second atomic bomb dropped on Nagasaki made use of ^{239}Pu. India exploded their first atomic bomb at **Pokhran** in **Rajasthan** in May 1974, and used ^{239}Pu as the fissionable material.

Nuclear Power and India (Recent Developments)

Indian scientists recently repeated the history of 11th May 1974. Our great scientists successfully conducted five underground nuclear tests at **Pokhran** range in **Rajasthan,** 24 years after the nation had conducted the first such test. Three tests were conducted at 3.45 p.m. on 11th May 1998 and the two tests were made later on 13th May. These tests were up to the mark and as per our expectations.

(b) Nuclear reactor or atomic reactor or atomic pile: The reactor is the furnace of the atomic age, the place where fissionable material is burnt for useful purposes. **It is essentially an instrument designed to allow a nuclear chain to develop, under control.** All the neutrons produced are not allowed to carry out the chain reaction. A fission reactor has five main components: (i) fuel, (ii) moderator, (iii) control rods, (iv) cooling system and (v) shielding.

(i) Fuel: Either enriched uranium or natural uranium is usually used as fuel. Heterogeneous reactors employ the fuel in the form of rods, plates or hollow cylinders. Homogeneous reactors employ solution of the fuel prepared in the moderator.

(ii) Moderator: The most efficient fission reactions occur with slow neutrons. Thus, the fast neutrons ejected during fission must be slowed down by collisions with atoms of comparable mass that do not absorb them. Such materials are called **moderators.** The most commonly used moderators are ordinary water and graphite. The most efficient moderator is helium. The next most efficient one is heavy water (D_2O) but this is so expensive that it has been used only in research reactors.

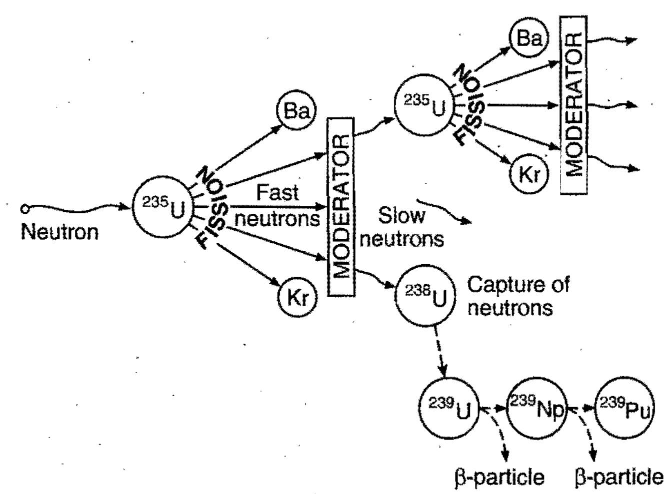


Fig. 3.9 Nuclear fission in a nuclear reactor using enriched uranium

(iii) Control rods: Boron or cadmium steel rods are used as control rods. These rods absorb neutrons and thereby control the rate of fission, *e.g.*,

$$_5B^{10} + {}_0n^1 \longrightarrow {}_3Li^7 + {}_2He^4$$

(iv) Cooling system: Liquid alloy of sodium and potassium is used as coolant; it takes away the heat to the exchanger. Heavy water, polyphenyls and carbon dioxide have also been used as coolants.

(v) Shielding: The reactor is enclosed in a steel containment vessel, which is housed in a thick-walled concrete building. Operating people are protected by a shield of compressed wood fibres.

Nuclear reactors are used:

1. To produce ^{239}Pu and ^{233}U: It is predicted that our limited supply of ^{235}U will last only another 50 years. However, non-fissionable ^{238}U and ^{232}Th are plentiful and can be converted into ^{239}Pu and ^{233}U. This conversion can be done in special type of reactors called **breeder reactors.** These reactors not only produce large quantities of heat from fission but also generate more fuel than they use because neutrons are absorbed in a thorium or uranium blanket to form ^{233}U and ^{239}Pu. This type of reactor requires the use of fast neutrons; no moderator is needed, but control is more difficult. Heat must be transferred very efficiently because ^{239}Pu melts at a relatively low temperature of 640°C. The process in which non fissile Nuclei $^{238}_{92}$U and $^{232}_{90}$Th are converted to fissile nuclei in breeder reactors is given below:

$$^{238}_{92}U + {}^1_0n \longrightarrow {}^{239}_{92}U \xrightarrow{-\beta^-} {}^{239}_{93}Np \xrightarrow{-\beta^-} {}^{239}_{94}Pu$$

$$^{232}_{90}Th + {}^1_0n \longrightarrow {}^{232}_{90}Th \xrightarrow{-\beta^-} {}^{233}_{91}Pa \xrightarrow{-\beta^-} {}^{233}_{92}U$$

2. To produce a strong beam of neutrons: These neutrons are used for making various isotopes which do not occur in nature. For example, ${}^{32}_{15}P$ and ${}^{60}_{27}Co$ are produced from the following nuclear reactions:

$$ {}^{130}_{52}Te + {}^{1}_{0}n \longrightarrow {}^{131}_{53}I + {}^{0}_{-1}e $$

$$ {}^{31}_{15}P + {}^{1}_{0}n \longrightarrow {}^{32}_{15}P $$

$$ {}^{32}_{16}S + {}^{1}_{0}n \longrightarrow {}^{32}_{15}P + {}^{1}_{1}H $$

$$ {}^{59}_{27}Co + {}^{1}_{0}n \longrightarrow {}^{60}_{27}Co $$

The non-radioactive isotope is taken in aluminium capsule which is placed inside the aluminium ball. The ball is rolled into the reactor where it is bombarded by neutrons slowed down by paraffin wax. The bombardment is continued for required period, which varies from element to element.

(c) Power plant (to generate electricity): The heat produced is utilised in generating steam which runs the steam turbines. The electric generator is connected to the turbine. The electric power is obtained from the generator. The atomic reactor when used for production of electricity is termed **power plant.**

The first nuclear reactor was assembled by **Fermi** and his co-workers at the University of Chicago in the United States of America, in 1942. In India, the first nuclear reactor was put into operation at Trombay (Mumbai), in 1956.

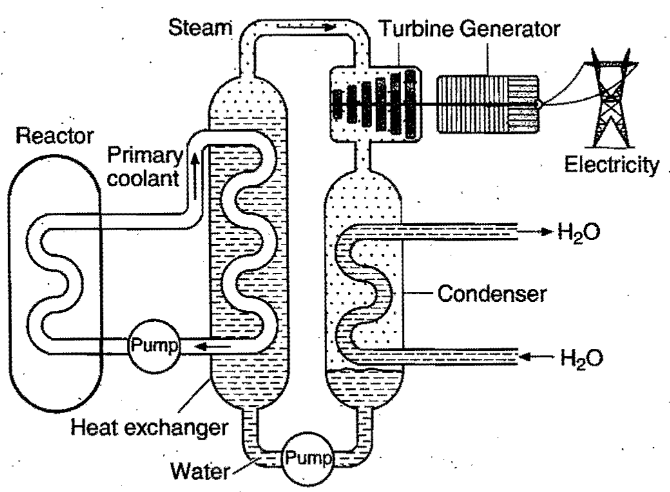


Fig. 3.10 Power plant : Application of nuclear fission for the production of electricity

3.16 NUCLEAR FUSION

A nuclear reaction in which two lighter nuclei are fused together to form a heavier nuclei is called nuclear fusion. In such a process, more stable nuclei come into existence as binding energy per nucleon increases (see sec. 2.23). A fusion reaction is difficult to occur because positively charged nuclei repel each other. At very high temperatures of the order of 10^6 to 10^7 K, the nuclei may have sufficient energy to overcome the repulsive forces and fuse. It is for this reason, fusion reactions are also called **thermonuclear reactions.** Fusion reactions are highly exothermic in nature because loss of mass occurs when heavier nuclei is formed from the two lighter nuclei. To initiate a fusion reaction is difficult, but once it is started, its continuity is maintained due to huge release of energy. Some examples of the fusion reactions are given below:

$$ {}^{2}_{1}H + {}^{2}_{1}H \longrightarrow {}^{4}_{2}He + 24.9\text{ MeV} $$

$$ {}^{3}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + 2\,{}^{1}_{0}n + 11.0\text{ MeV} $$

$$ {}^{1}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + 20.0\text{ MeV} $$

$$ {}^{2}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + {}^{1}_{0}n + 17.8\text{ MeV} $$

$$ {}^{7}_{3}Li + {}^{1}_{1}H \longrightarrow 2\,{}^{4}_{2}He + 17.7\text{ MeV} $$

Hydrogen bomb is based on fusion reactions. Energy released is so enormous that it is about 1000 times that of an atomic bomb. In hydrogen bomb, a mixture of deuterium oxide (D_2O) and tritium oxide (T_2O) is enclosed in a space surrounding an ordinary atomic bomb. The temperature produced by the explosion of the atomic bomb initiates the fusion reaction between ${}^{3}_{1}H$ and ${}^{2}_{1}H$ releasing huge amount of energy. The first hydrogen bomb was exploded in 1952. So far, it has not been possible to bring about fusion under controlled conditions.

It is believed that the high temperature of stars including the sun is due to fusion reactions. **Bethe** and **Weizsaeker,** in 1939, proposed that a carbon-nitrogen cycle is responsible for the production of solar energy in which hydrogen is converted into helium. The cycle is:

$$ {}^{12}_{6}C + {}^{1}_{1}H \longrightarrow {}^{13}_{7}N + \gamma $$

$$ {}^{13}_{7}N \longrightarrow {}^{13}_{6}C + {}^{0}_{+1}e $$

$$ {}^{13}_{6}C + {}^{1}_{1}H \longrightarrow {}^{14}_{7}N + \gamma $$

$$ {}^{14}_{7}N + {}^{1}_{1}H \longrightarrow {}^{15}_{8}O + \gamma $$

$$ {}^{15}_{8}O \longrightarrow {}^{15}_{7}N + {}^{0}_{+1}e $$

$$ {}^{15}_{7}N + {}^{1}_{1}H \longrightarrow {}^{12}_{6}C + {}^{4}_{2}He + \gamma $$

or

$$ 4\,{}^{1}_{1}H \longrightarrow {}^{4}_{2}He + 2\,{}^{0}_{+1}e + 24.7\text{ MeV} $$

${}^{12}_{6}C$ acts as a kind of nuclear catalyst.

E. Saltpeter, in 1953, proposed a proton-proton chain reaction:

$$ {}^{1}_{1}H + {}^{1}_{1}H \longrightarrow {}^{2}_{1}H + {}^{0}_{+1}e + \gamma $$

$$ {}^{2}_{1}H + {}^{1}_{1}H \longrightarrow {}^{3}_{2}He + \gamma $$

$$ {}^{3}_{2}He + {}^{1}_{1}H \longrightarrow {}^{4}_{2}He + {}^{0}_{+1}e + \gamma $$

$$ 4\,{}^{1}_{1}H \longrightarrow {}^{4}_{2}He + 2\,{}^{0}_{+1}e + 24.7\text{ MeV} $$

As a potential source of commercial electrical power, the fusion process has several advantages over the fission reaction. (i) The quantity of energy liberated in the fusion is much greater than in fission. (ii) The products of fusion are non-radioactive. Fission produces many unstable radioactive products. Fission reactors, therefore, pose a waste-disposal problem.

Difference between Nuclear Fission and Nuclear Fusion

Nuclear Fission	Nuclear Fusion
(i) This process occurs in heavy nuclei.	This process occurs in lighter nuclei.
(ii) The heavy nucleus splits into lighter nuclei of comparable masses.	The lighter nuclei fuse together to form a heavy nucleus.
(iii) The binding energy per nucleon increases.	The binding energy per nucleon increases.
(iv) This reaction occurs at ordinary temperature.	This occurs at a very high temperature.
(v) The energy liberated in one fission is about 200 MeV.	The energy liberated in one fusion is about 24 MeV.
(vi) This can be controlled.	This cannot be controlled.
(vii) Products of fission are usually unstable radioactive in nature.	Products of fusion are usually stable and non-radioactive in nature.
(viii) Percentage efficiency is less. % efficiency $=\frac{200}{236\times 931}\times 100 = 0.09$	Percentage efficiency is high. % efficiency $=\frac{17.8}{5\times 931}\times 100 = 0.38$ $[{}_1^2H + {}_1^3H \rightarrow {}_2^4He + {}_0^1n + 17.8\text{ MeV}]$
(ix) The links of fission reactions are neutrons.	The links of fusion reactions are protons.

3.17 SYNTHETIC ELEMENTS INCLUDING TRANSACTINIDES

Elements 43 (technetium), 61 (promethium), 85 (astatine) and all elements with $Z > 92$ do not exist naturally on the earth, because no isotopes of these elements are stable. The elements coming after uranium ($Z = 92$) are named **transuranic** or **transuranium** elements. The actinide series which starts with the element thorium ($Z = 90$) is complete at the element lawrencium ($Z = 103$). The elements with $Z = 104-112$ have been reported recently and are transition (*d*-block-fourth series) elements. These are called **transactinides** or **super heavy elements.** After the discovery of nuclear reactions early in the twentieth century, scientists between 1937 and 1945, set out to make the missing elements, *i.e.*, technetium, promethium and astatine and three members of the actinide series, neptunium ($Z = 93$), plutonium ($Z = 94$) and americium ($Z = 95$). The missing elements and all the elements above atomic number 92 are called **synthetic elements** as these have been synthesised by artificial transmutation, *i.e.*, by nuclear reactions. The credit for the discovery of most of the transuranic elements goes to **Seaborg.**

Much less is known about synthetic elements as these are radioactive and short-lived. This is also due to their limited availability. The production of synthetic elements requires binuclear reactions between two positive nuclei that must be fused together against the force of electrical repulsion. Nuclear accelerators were used for this purpose. High energy deuterons were used to increase the atomic number of target nuclei by one unit.

$$^{98}_{42}Mo + {}^2_1H \longrightarrow {}^{99}_{43}Tc + {}^1_0n$$

$$^{238}_{92}U + {}^2_1H \longrightarrow {}^{238}_{93}Np + 2{}^1_0n$$

Elements 93 and 94 were produced using neutrons (obtained during fission) instead of accelerated positive nuclei. Neutron capture by ^{238}U followed by β-emission gives isotopes with mass number 239.

$$^{238}_{92}U + {}^1_0n \longrightarrow {}^{239}_{92}U \xrightarrow{-\beta} {}^{239}_{93}Np \xrightarrow{-\beta} {}^{239}_{94}Pu$$

$^{239}_{94}Pu$ is an α-emitter with half life of 2.4×10^4 years.

Americium is formed in a similar way.

$$^{239}_{94}Pu \xrightarrow{(n,\gamma)} {}^{240}_{94}Pu \xrightarrow{(n,\gamma)} {}^{241}_{94}Pu \xrightarrow{-\beta} {}^{241}_{95}Am$$

As Z increases, the efficiency of nuclear reactions with neutron bombardment falls sharply. Instead, nuclides in the $Z = 95$ to 99 range are bombarded with beams of helium nuclei accelerated in the cyclotron to form nuclides with atomic numbers 96 to 101.

$$^{239}_{94}Pu + {}^4_2He \longrightarrow {}^{242}_{96}Cm + {}^1_0n$$

$$^{241}_{95}Am + {}^4_2He \longrightarrow {}^{243}_{97}Bk + 2{}^1_0n$$

$$^{242}_{96}Cm + {}^4_2He \longrightarrow {}^{245}_{98}Cf + {}^1_0n$$

$$^{249}_{98}Cf + {}^4_2He \longrightarrow {}^{251}_{100}Fm + 2{}^1_0n$$

$$^{253}_{99}Es + {}^4_2He \longrightarrow {}^{256}_{101}Md + {}^1_0n$$

Beyond element with $Z = 101$, increasingly heavier nuclei are used as projectiles. These projectiles are accelerated by linear rather than circular accelerators. Examples of nuclear reactions of this type are the following:

$$^{246}_{96}Cm + {}^{12}_6C \longrightarrow {}^{254}_{102}No + 4{}^1_0n$$

$$^{252}_{98}Cf + {}^{11}_5B \longrightarrow {}^{257}_{103}Lr + 6{}^1_0n$$

$$^{238}_{92}U + {}^{14}_7N \longrightarrow {}^{249}_{99}Es + 3{}^1_0n$$

$$^{238}_{92}U + {}^{16}_8O \longrightarrow {}^{250}_{100}Fm + 4{}^1_0n$$

The superheavy elements have been discovered by bombardment with medium weight nuclei. For example, the elements with $Z = 107$ and $Z = 109$ have been obtained by bombardment of $^{209}_{83}Bi$ with accelerated $^{54}_{24}Cr$ and $^{58}_{26}Fe$ respectively.

$$^{209}_{83}Bi + {}^{54}_{24}Cr \longrightarrow {}^{261}_{107}Uns + 2{}^1_0n$$

$$^{209}_{83}Bi + {}^{58}_{26}Fe \longrightarrow {}^{266}_{109}Une + {}^1_0n$$

The elements with $Z = 104$, 105, 106 and 108 have also been reported by the applications of the following reactions:

$$^{249}_{98}Cf + {}^{12}_6C \longrightarrow {}^{257}_{104}Unq + 4{}^1_0n$$

$$^{249}_{98}Cf + {}^{15}_7N \longrightarrow {}^{260}_{105}Unp + 4{}^1_0n$$

$$^{249}_{98}\text{Cf} + ^{18}_{8}\text{O} \longrightarrow ^{263}_{106}\text{Unh} + 4^{1}_{0}n$$

$$^{208}_{82}\text{Pb} + ^{58}_{26}\text{Fe} \longrightarrow ^{265}_{108}\text{Uno} + ^{1}_{0}n$$

The elements up to 100 (fermium) undergo radioactive decay mainly by emitting α-particles or β-particles. The elements become increasingly unstable as the atomic number increases and nobelium has a half life of only three seconds. With these heavy elements, spontaneous nuclear fission becomes the most important method of decay. ^{252}Cf could become a valuable neutron source.

The IUPAC names for the elements $Z > 100$ have been given below:

101 Unnilunium	Unu	107 Unnilseptium	Uns
102 Unnilbium	Unb	108 Unniloctium	Uno
103 Unniltrium	Unt	109 Unnilennium	Une
104 Unnilquadium	Unq	110 Ununnilium	Uun
105 Unnilpentium	Unp	111 Unununium	Uuu
106 Unnilhexium	Unh	112 Ununbium	Uub

Elements with an even number of protons in the nucleus are usually more stable than their neighbours with odd atomic numbers, *i.e.*, they are less likely to decay. Also nuclei with both an even number of protons and an even number of neutrons are more likely to be stable. A nucleus is more stable than average if the numbers of neutrons or protons are 2, 8, 20, 28, 50, 82 or 126. These are called 'magic numbers' and can be explained by the shell structure of the nucleus. This theory also requires the inclusion of numbers 114, 164 and 184 in the series of magic numbers. The stability is particularly high if number of protons and the number of neutrons are magic numbers. Thus, $^{208}_{82}$Pb is very stable with 82 protons and (208 − 82) 126 neutrons. This suggests that nuclides as Uuq ($Z = 114$, $A = 278$), Uuq ($Z = 114$, $A = 298$) and Ubh ($Z = 126$, $A = 310$) might be stable enough to exist. Considerable efforts are being made to produce elements 114 and 126 but the present techniques have so far only succeeded in producing unstable isotopes. The elements up to $Z = 112$ have been reported so far.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

19. The radioactive isotope $^{60}_{27}$Co which is used in the treatment of cancer can be made by (n, p) reaction. For this reaction, the target nucleus is:
(a) $^{59}_{28}$Ni (b) $^{59}_{27}$Co (c) $^{60}_{28}$Ni (d) $^{60}_{27}$Co
[Manipal (Med.) 2007]

[**Ans.** (c)]

[**Hint:** $^{60}_{28}\text{Ni} + ^{1}_{0}n \longrightarrow ^{60}_{27}\text{Co} + ^{1}_{1}\text{H}$]

20. $^{14}_{7}$N is attacked by doubly charged helium ion, it emits a proton and:
(a) $^{18}_{9}$F (b) $^{17}_{8}$O (c) $^{18}_{8}$O (d) $^{19}_{9}$F
[JEE (Orissa) 2007]

[**Ans.** (b)]

[**Hint:** $^{14}_{7}\text{N} + ^{4}_{2}\text{He} \longrightarrow ^{17}_{8}\text{O} + ^{1}_{1}\text{H}$]

21. A nuclear reaction of $^{235}_{92}$U with a neutron produces $^{90}_{36}$Kr and two neutrons. Other element produced in this reaction is:
(a) $^{137}_{52}$Te (b) $^{144}_{55}$Cs (c) $^{137}_{56}$Ba (d) $^{144}_{56}$Ba
(VITEEE 2007)

[**Ans.** (d)]

[**Hint:** $^{235}_{92}\text{U} + ^{1}_{0}n \longrightarrow ^{90}_{36}\text{Kr} + ^{144}_{56}\text{Ba} + 2\,^{1}_{0}n$]

22. The product P of the nuclear reaction $^{235}_{92}\text{U} + ^{1}_{0}n \longrightarrow P + ^{92}_{36}\text{Kr} + 3(^{1}_{0}n)$ is :
(a) $^{141}_{56}$Sr (b) $^{141}_{56}$La (c) $^{141}_{56}$Ba (d) $^{141}_{56}$Cs
[JEE (WB) 2008]

[**Ans.** (c)]

[**Hint:** Let symbol of element is $^{M}_{Z}P$.

$92 = Z + 36$

$\therefore \quad Z = 56$

$235 + 1 = M + 92 + 3$

$M = 141$

Thus, the element P will be $^{141}_{56}$Ba]

3.18 APPLICATIONS OF RADIOACTIVITY

(a) Use of γ-rays: γ-rays are used for disinfecting food grains and for preserving foodstuffs. Onions, potatoes, fruits and fish, etc., when irradiated with γ-rays, can be preserved for long periods. High yielding disease resistant varieties of wheat, rice, groundnut, jute, etc., can be developed by the application of nuclear radiations. The γ-radiations are used in the treatment of cancer. The γ-radiations emitted by cobalt-60 can burn cancerous cells. γ-radiations are used to sterilize medical instruments like syringes, blood transfusion sets, etc. These radiations make the rubber and plastics objects heat resistant.

(b) The age of the earth: The age of the earth has been estimated by uranium dating technique. The uranium ore (rock) which is found in nature is associated with non-radioactive lead which is believed to be the end product of radioactive disintegration of uranium. A sample of uranium rock is analysed for ^{238}U and ^{206}Pb contents. From this analysis, let the quantities in mole be $N = {}^{238}\text{U}$ mole, $N_0 = {}^{238}\text{U mole} + {}^{206}\text{Pb mole}$.

Applying disintegration equation,

$$\lambda t = 2.303 \log_{10} \frac{N_0}{N}$$

$$= 2.303 \log_{10} \frac{^{238}\text{U} + ^{206}\text{Pb}}{^{238}\text{U}}$$

$$= 2.303 \log_{10} \left[1 + \frac{^{206}\text{Pb}}{^{238}\text{U}}\right]$$

The value of 't' can be calculated by putting the value of λ which is equal to $\frac{0.693}{t_{1/2}}$.

So, $$t = \frac{2.303 \times t_{1/2}}{0.693} \log_{10} \left[1 + \frac{^{206}\text{Pb}}{^{238}\text{U}}\right]$$

Here 't' corresponds to the age of earth which has been found to be 4.5 billion years.

Example 27. *A sample of uranium mineral was found to contain ^{206}Pb and ^{238}U in the ratio of 0.008 : 1. Estimate the age of the mineral. (Half life of ^{238}U is 4.51×10^9 years)*

Solution: We know that, $t = \frac{2.303 t_{1/2}}{0.693} \log \left[1 + \frac{^{206}\text{Pb}}{^{238}\text{U}}\right]$

Given, $t_{1/2} = 4.51 \times 10^9$ years

Ratio by mass of $^{206}\text{Pb} : {}^{238}\text{U} = 0.008 : 1$

Ratio by moles of $^{206}\text{Pb} : {}^{238}\text{U} = \frac{0.008}{206} : \frac{1}{238} = 0.0092$

So, $t = \frac{2.303 \times 4.51 \times 10^9}{0.693} \log [1 + 0.0092]$

$= \frac{2.303 \times 4.51 \times 10^9}{0.693} \times 0.00397$

$= \frac{0.0412}{0.693} \times 10^9 = 0.05945 \times 10^9$ years

Hence, age of the mineral is 5.945×10^7 years.

(c) Radio carbon dating: By using the half life period of ^{14}C, it is possible to determine the age of various objects. In living material the ratio of ^{14}C to ^{12}C remains relatively constant. When a tissue in an animal or plant dies, ^{14}C decreases because the intake and utilization of ^{14}C do not occur. Therefore, in the dead tissue the ratio of ^{14}C to ^{12}C would decrease, depending on the age of the tissue. The age of the dead tissue is determined in the following way. A sample of dead tissue is burnt to carbon dioxide and the carbon dioxide is analysed for the ratio of ^{14}C to ^{12}C. From this data, the age of the dead tissue can be determined. Thus:

$$\lambda = \frac{2.303}{t} \log_{10} \left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{t_{1/2} \text{ of } C^{14}} = \frac{2.303}{t} \log_{10} \left(\frac{N_0}{N}\right)$$

N_0 = Ratio of C^{14} / C^{12} in green plant or atmosphere

N = Ratio of C^{14} / C^{12} in wood

or

N_0 = Activity of green plant per unit mass

N = Activity of wood per unit mass

Although, the method is suitable to a variety of organic materials, accuracy depends on the half life to be used, variations in levels of atmospheric carbon-14 and contamination. (The half life radio carbon was redefined from 5570 ± 30 years to 5730 ± 40 years by IUPAC). The rapid disintegration of carbon-14 generally limits the dating period to approximately 50,000 years.

Example 28. *The amount of $^{14}_{6}C$ isotope in a piece of wood is found to be one-fifth of that present in a fresh piece of wood. Calculate the age of wood. (Half life of ^{14}C = 5577 years)*

Solution: We know that, $t = \frac{2.303 \times t_{1/2}}{0.693} \log \left(\frac{N_0}{N}\right)$

Given, $N = \frac{N_0}{5}$

So, $t = \frac{2.303 \times 5577}{0.693} \log 5$

or $t = \frac{2.303 \times 5577}{0.693} \times 0.6989 = 12953$ years

Example 29. *A piece of wood was found to have $^{14}C/^{12}C$ ratio 0.6 times that in a living plant. Calculate the period when the plant died. (Half life of ^{14}C = 5760 years)*

Solution: We know that, $t = \frac{2.303 \times t_{1/2}}{0.693} \log \left(\frac{N_0}{N}\right)$

So, $t = \frac{2.303 \times 5760}{0.693} \log \left(\frac{1}{0.6}\right)$

$= \frac{2.303 \times 5760}{0.693} \times 0.2201$

$= 4213$ years

(d) Potassium-Argon method: The decay of radioactive potassium isotope to argon is widely used for dating rocks. The geologists are able to date entire rock samples in this way, because potassium-40 is abundant in micas, feldspars and hornblendes. Leakage of Argon is however problem if the rock has been exposed to temperature above 125° C.

(e) Rubidium-Strontium method: This method of dating is used to date ancient igneous and metamorphic terrestrial rocks as well as lunar samples. It is based on disintegration by beta decay of ^{87}Rb to ^{87}Sr. This method is frequently used to check potassium-argon dates, because the strontium daughter element is not diffused by mild heating like argon.

ILLUSTRATIONS OF OBJECTIVE QUESTIONS

23. A wooden artifact sample gave activity of 32 β-particles per second while the freshly cut wood gave activity of 64 β-particles per second in G. M. counter. Calculate the age of the wooden artifact ($t_{1/2}$ of ^{14}C = 5760 yrs):

(a) 11520 yrs (b) 5760 yrs
(c) 2880 yrs (d) 1440 yrs

[**Ans.** (b)]

[**Hint:** $\frac{0.693}{t_{1/2}} = \frac{2.303}{t_{age}} \log_{10} \left(\frac{N_0}{N}\right)$

$\frac{0.693}{5760} = \frac{2.303}{t_{age}} \log_{10} \left(\frac{64}{32}\right)$

$t_{age} = 5760$ yrs]

24. The analysis of a rock shows that the relative number of ^{206}Pb and ^{238}U atoms is Pb/U = 0.25. If $t_{1/2}$ of ^{238}U is 4.5×10^9 yrs, then the age of the rock will be:

(a) $\frac{2.303}{0.693} (4.5 \times 10^9) \log \left(\frac{5}{4}\right)$

(b) $\frac{2.303}{0.693}(4.5\times 10^9)\log\left(\frac{1}{4}\right)$

(c) $\frac{2.303}{0.693}(4.5\times 10^9)\log(4)$

(d) $\frac{2.303}{0.693}(4.5\times 10^9)\log\left(\frac{4}{5}\right)$

[Ans. (a)]

[Hint: $\frac{\text{Pb}}{\text{U}} = 0.25$; $\therefore\ 1+\frac{\text{Pb}}{\text{U}} = 1.25$

$$\frac{\text{U}+\text{Pb}}{\text{U}} = 1.25$$

$$\frac{N_0}{N} = 1.25 = \frac{5}{4}$$

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{t}\log\left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{4.5\times 10^9} = \frac{2.303}{t_{age}}\log\left(\frac{5}{4}\right)$$

$$t_{age} = \frac{2.303}{0.693}(4.5\times 10^9)\log\left(\frac{5}{4}\right)]$$

25. Assuming that about 200 MeV of energy is released per fission of ${}^{235}_{92}$U nuclei, then the mass of ^{235}U consumed per day in a fission reactor of power 1 megawatt will be approximately:

(a) 10^{-2} g (b) 1 g (c) 100 g (d) 1000 g

[Ans. (b)]

[Hint: 1 MW = $10^6 \times 24\times 60\times 60$ J

$$\text{Number of fissions} = \frac{24\times 6\times 6\times 10^8}{200\times 10^6\times 1.6\times 10^{-19}} = 2.7\times 10^{21}$$

$$\text{Mass of uranium} = 2.7\times 10^{21}\times 235\times 1.66\times 10^{-24} = 1.05\text{ g}]$$

26. What is the binding energy of the hydrogen nucleus?

(a) Zero (b) 13.6 eV

(c) More than 13.6 eV (d) Infinite

[Ans. (a)]

[Hint: Nucleus of hydrogen has only one proton; hence its binding energy will be zero.]

27. Which of the following is not the inverse square law force?

(a) Electric force (b) Gravitational force

(c) Nuclear force

(d) Magnetic force between two poles

[Ans. (c)]

[Hint: Nuclear forces are short range forces which do not obey inverse square law.]

28. Lead is the final product formed by a series of changes in which the rate determining stage is the radioactive decay of uranium-238. This radioactive decay is first order with half life of 4.5×10^9 years. What would be the age of a rock sample originally leadfree, in which the molar proportion of uranium to lead is now 1 : 3? [PET (Kerala) 2006]

(a) 1.5×10^9 years (b) 2.25×10^9 years

(c) 4.5×10^9 years (d) 9×10^9 years

(e) 13.5×10^9 years

[Ans. (d)]

[Hint: $\frac{0.693}{t_{1/2}\text{U}^{238}} = \frac{2.303}{t_{age}}\log\left(\frac{N_0}{N}\right)$

$$\frac{0.693}{4.5\times 10^9} = \frac{2.303}{t_{age}}\log_{10}\left(\frac{4}{1}\right)$$

$$t_{age} = 9\times 10^9\text{ years}]$$

(f) Use of radioisotopes (tracers): Tracers have been used in the following fields:

(i) In medicine: Radioisotopes are used to diagnose many diseases. For example, arsenic-74 tracer is used to detect the presence of tumours; sodium-24 tracer is used to detect the presence of blood clots and iodine-131 tracer is used to study the activity of the thyroid gland. It should be noted that the radioactive isotopes used in medicine have very short half life periods.

^{90}Y: This isotope is used in the treatment of joint effusion and arthritis.

^{59}Fe: Used in the detection of anaemia.

^{32}P: This isotope is used in the treatment of polycythaemia, thrombocythaemia, skeletal metastasis, prostate SR and breast SR.

Nuclear Medicine Scan: It is an advanced nuclear technology used in diagnosis of diseases. **Magnetic Resonance Imaging** (MRI), a diagnostic medical imaging technique utilizes the principle of nuclear magnetic resonance. The first images using magnetic resonance were published in early 1970s, and medical applications have accelerated in the world during the decade of 1983 to 1993. MRI is now a most versatile, powerful and sensitive diagnostic imaging modality available. Its medical importance can be summarised briefly as having the ability to non-invasively generate thin section, functional images of any part of the body at any angle and direction in a relatively short period of time. MRI also visualizes the heart with exquisite anatomical detail at any angle and direction.

The principle of MRI is applicable in human body because we are all filled with small biological magnets, the most abundant and responsive of which is the nucleus of hydrogen atom, the proton.

Computerized Axial Tomography: Computerized Axial Tomography (CT or CAT), non-invasive diagnostic technique uses a type of X-ray device that provides a clear view of soft internal organ tissues in the body. CT is used to diagnose various conditions, in particular cancer. A CT scan is the computer analysis of a sharply limited, thin X-ray beam passed circumferentially through an area of the body, producing a cross-sectional image, or slice.

The modern CT scanner comprises five major parts. A high-speed X-ray tube cooled by oil, air and water forms the X-ray source. Its X-ray detector, normally a bank of about 1,000 solid state-crystal microprocessors coated with caesium iodide, receives the attenuated X-ray signal as it passes through the various tissues and bones of the patient being examined. The signal is electronically converted to binary data, which is read by the computer—the heart of the CT imaging system. The CT has a gantry, a framework that is mounted in such a way that it surrounds the patient in a vertical plane, and contains a rotating sub-frame onto which the X-ray source and detectors are mounted. A patient table (or couch) is positioned perpendicular and axial to the gantry so that it is able to travel along that axis.

Topographic images are produced by using an X-ray source and a detector moving in a coupled way relative to the patient. In CT a thin fan beam of radiation rotates in a circular or spiral motion around the patient. Thousands of projected X-ray signals are reconstructed by computer algorithms to produce digital CT images, displayed by a high-resolution monitor. In this way the whole body can be imaged from head to toe.

Radiation Dosage in the Radiotherapy of Cancer

Radiations and the particles emitted by radioactive nuclei are harmful for living organisms. These radiations cause genetic disorders by affecting DNA.

Effect of biological radiations can be measured in terms of the unit called RAD.

RAD = Radiation absorbed dose

1 RAD = The radiation which deposits 1×10^{-2} J of energy per kilogram of tissue.

In order to measure biological destruction by radiation, an other unit REM was introduced.

$$\text{REM} = \text{RAD} \times \text{RBE}$$

RBE = Relative biological effectiveness

RBE for α-particle = 10 unit

RBE for β and γ radiation = 1 unit

RBE for neutron = 5 unit

(ii) In agriculture: The use of radioactive phosphorus ^{32}P in fertilizers has revealed how phosphorus is absorbed by plants. This study has led to an improvement in the preparation of fertilizers. ^{14}C is used to study the kinetics of photosynthesis.

(iii) In industry: Radioisotopes are used in industry to detect the leakage in underground oil pipelines, gas pipelines and water pipes. Radioactive isotopes are used to measure the thickness of materials, to test the wear and tear inside a car engine and the effectiveness of various lubricants. Radioactive carbon has been used as a tracer in studying mechanisms involved in many reactions of industrial importance such as alkylation, polymerisation, catalytic synthesis, etc.

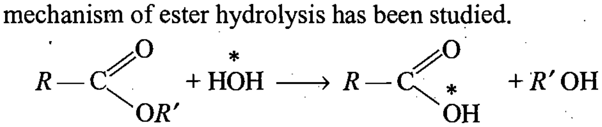
(iv) Analytical studies: Several analytical procedures can be used employing radioisotopes as tracers.

1. Adsorption and occlusion studies: A small amount of radioactive isotope is mixed with an inactive substance and the activity is studied before and after adsorption. Fall in activity gives the amount of substance adsorbed.

2. Solubility of sparingly soluble salts: The solubility of lead sulphate in water may be estimated by mixing a known amount of radioactive lead with ordinary lead. This is dissolved in nitric acid and precipitated as lead sulphate by adding sulphuric acid. Insoluble lead sulphate is filtered and the activity of the water is measured. From this, the amount of $PbSO_4$ still present in water can be estimated.

3. Ion-exchange technique: Ion exchange process of separation is readily followed by measuring activity of successive fractions eluted from the column.

4. Reaction mechanism: By labelling oxygen of the water, mechanism of ester hydrolysis has been studied.



5. Study of efficiency of analytical separations: The efficiency of analytical procedures may be measured by adding a known amount of radioisotope to the sample before analysis begins. After the completion, the activity is again determined. The comparison of activity tells about the efficiency of separation.

MISCELLANEOUS NUMERICAL EXAMPLES

Example 1. *One mole of A present in a closed vessel undergoes decay as:*

$$^{m}_{Z}A \longrightarrow {}^{m-8}_{Z-4}B + 2({}^{4}_{2}He)$$

What will be the volume of helium gas collected at STP after 20 *days* ($t_{1/2}$ *of* $A = 10$ *days*)?

Solution: We know that,

$$N = N_0 \left(\frac{1}{2}\right)^n \text{ where, } N = \text{remaining mole of } A$$

$$N = 1\left(\frac{1}{2}\right)^2 = \frac{1}{4}$$

Number of decayed moles $= 1 - \frac{1}{4} = \frac{3}{4}$

Number of moles of helium formed

$= 2 \times$ number of decayed moles of $A = 2 \times \frac{3}{4} = \frac{3}{2}$

Volume of helium at STP $= \frac{3}{2} \times 22.4 = 33.6$ litre

Example 2. ^{131}I *has half life period* 13.3 *hour. After* 79.8 *hour, what fraction of* ^{131}I *will remain?* [CBSE (PMT) 2005]

Solution: $N = N_0 \left(\frac{1}{2}\right)^n$

$$\frac{N}{N_0} = \left(\frac{1}{2}\right)^6 = \frac{1}{64}$$

Example 3. *A sample of* $^{14}CO_2$ *was to be mixed with ordinary* CO_2 *for a biological tracer experiment. In order that* 10^3 cm^3 *of the diluted gas at NTP should have* 10^4 *dis/min, how many* μCi *of radiocarbon*-14 *are needed to prepare* 60 *L of the diluted gas?*

Solution: 10 cm^3 of the diluted gas at NTP

$$= 10^4 \text{ dis/min} = \frac{10^4}{60} \text{ dps}$$

$\therefore$ 60 L (60,000 cm^3) of the dilute gas at NTP

$$= \frac{10^4 \times 60{,}000}{60 \times 10} \text{ dps}$$

Thus, no. of μCi of $^{14}CO_2$ needed

$$= \frac{10^4 \times 60{,}000}{60 \times 10 \times 3.7 \times 10^4} \quad (1\,\mu\text{Ci} = 3.7 \times 10^4 \text{ dps})$$

$$= 27.03\,\mu\text{Ci}$$

Example 4. *A radioactive nuclide is produced at a constant rate of 'α' per second. Its decay constant is λ. If N_0 be the number of nuclei at time $t = 0$, then what will be the maximum number of possible nuclei?*

(a) $\frac{\alpha}{\lambda}$ (b) $N_0 + \frac{\alpha}{\lambda}$ (c) N_0 (d) $\frac{\lambda}{\alpha} + N_0$

Solution: Maximum number of nuclei will be present when

Rate of decay = Rate of formation

$$\lambda N = \alpha$$

or

$$N = \frac{\alpha}{\lambda}$$

Example 5. *The half life of ^{212}Pb is 10.6 hour. It undergoes decay to its daughter (unstable) element ^{212}Bi of half life 60.5 minute. Calculate the time at which the daughter element will have maximum activity.*

Solution: $\lambda_{Pb} = \frac{0.693}{10.6 \times 60} = 1.0896 \times 10^{-3} \text{ min}^{-1}$

$$\lambda_{Bi} = \frac{0.693}{60.5} = 11.45 \times 10^{-3} \text{ min}^{-1}$$

$$t_{max} = \frac{2.303}{\lambda_{Bi} - \lambda_{Pb}} - \log \frac{\lambda_{Bi}}{\lambda_{Pb}}$$

$$= \frac{2.303}{(11.45 \times 10^{-3} - 1.0896 \times 10^{-3})} \times \log \frac{11.45 \times 10^{-3}}{1.0896 \times 10^{-3}}$$

$$= 227.1 \text{ min}$$

Example 6. *A radioactive isotope is being produced at a constant rate x. Half life of the radioactive substance is 'y'. After sometime, the number of radioactive nuclei becomes constant, the value of this constant is ......*

Solution: At the stage of radioactive equilibrium,

Rate of formation of nuclide = Rate of decay of nuclide

$$x = \lambda N$$

$$N = \frac{x}{\lambda} = \frac{x}{(\ln 2)/y} = \frac{xy}{\ln 2}$$

Example 7. *$^{238}_{92}U$ by successive radioactive decay changes to $^{206}_{82}Pb$. A sample of uranium ore was analysed and found to contain 1.0 g of ^{238}U and 0.1 g of ^{206}Pb. Assuming that all ^{206}Pb has accumulated due to decay of ^{238}U, find the age of the ore (half life of $^{238}U = 4.5 \times 10^9$ yrs).*

Solution: Number of moles of $^{238}U = \frac{1}{238}$

Number of moles of $^{206}Pb = \frac{0.1}{206}$

Applying the relationship,

$$t = \frac{2.303}{\lambda} \log \left[1 + \frac{^{206}Pb}{^{238}U}\right]$$

$$= \frac{2.303}{0.693} \times 4.5 \times 10^9 \log \left[1 + \frac{\frac{0.1}{206}}{\frac{1}{238}}\right]$$

$$= 7.098 \times 10^8 \text{ years}$$

Example 8. *Calculate the mass of C^{14} (half life = 5720 years) atoms which give 3.7×10^7 disintegrations per second.*

Solution: Let the mass of ^{14}C atoms be m g.

Number of atoms in m g of $^{14}C = \frac{m}{14} \times 6.02 \times 10^{23}$

$$\lambda = \frac{0.693}{\text{half life}} = \frac{0.693}{5720 \times 365 \times 24 \times 60 \times 60} = 3.84 \times 10^{-12} \text{ sec}^{-1}$$

We know that, $-\frac{dN}{dt} = \lambda \cdot N$

i.e., Rate of disintegration = $\lambda \times$ no. of atoms

$$3.7 \times 10^7 = \frac{0.693}{5720 \times 365 \times 24 \times 60 \times 60} \times \frac{m}{14} \times 6.02 \times 10^{23}$$

$$= \frac{3.84 \times 10^{-12} \times m \times 6.02 \times 10^{23}}{14}$$

So, $m = 2.24 \times 10^{-4}$ g

Example 9. *Prove that time required for 99.9% decay of a radioactive species is almost ten times its half life period.*

Solution: We know that, $t = \frac{2.303}{\lambda} \log \frac{N_0}{N}$

$$N_0 = 100, \quad N = (100 - 99.9) = 0.1$$

So, Time required for 99.9% decay, $t = \frac{2.303}{\lambda} \log \frac{100}{0.1}$

$$= \frac{2.303}{\lambda} \times 3$$

$$\text{Half life period} = \frac{0.693}{\lambda}$$

So, $\frac{\text{Time required for 99.9\% decay}}{\text{Half life period}} = \frac{2.303 \times 3}{\lambda} \times \frac{\lambda}{0.693}$

$$\approx 10$$

Example 10. *Half life of a radioactive substance A is two times the half life of another radioactive substance B. Initially the number of nuclei of A and B are N_A and N_B respectively. After three half lives of 'A', number of nuclei of both become equal. The ratio of $\frac{N_A}{N_B}$ will be:*

(*a*) $\frac{1}{2}$ (*b*) $\frac{1}{8}$ (*c*) $\frac{1}{3}$ (*d*) $\frac{1}{6}$

Solution: We know that, the amount remaining after n half lives can be calculated as:

$$N = N_0 \left(\frac{1}{2}\right)^n$$

$$\text{Remaining amount of } A = N_A \left(\frac{1}{2}\right)^3$$

$$\text{Remaining amount of } B = N_B \left(\frac{1}{2}\right)^6$$

$$N_A \left(\frac{1}{2}\right)^3 = N_B \left(\frac{1}{2}\right)^6$$

$$\frac{N_A}{N_B} = \frac{8}{64} = \frac{1}{8}$$

Example 11. *1.0 g of $^{198}_{79}Au$ ($t_{1/2} = 65$ hours) decays by β-emission to produce mercury.*

(*a*) *Write the nuclear reaction for the process.*

(*b*) *How much mercury will be present after 260 hours?*

Solution: (a) $^{198}_{79}\text{Au} \longrightarrow {}^{198}_{80}\text{Hg} + {}^{0}_{-1}e$

(b) No. of half lives in 260 hours $= \frac{260}{65} = 4$

$$\text{Amount of gold left after 4 half lives} = \left(\frac{1}{2}\right)^4 = \frac{1}{16}\text{g}$$

$$\text{Amount of gold disintegrated} = 1 - \frac{1}{16} = \frac{15}{16}\text{g}$$

$$\text{Amount of mercury formed} = \frac{15}{16} = 0.9375\text{ g}$$

Example 12. *Calculate the probability (P) of survival of a radioactive nucleus for one mean life.*

Solution: Probability for survival $= \frac{N}{N_0} = e^{-\lambda t}$

$$t = \text{mean life} = \frac{1}{\lambda}$$

$$\therefore \quad \text{Probability} = e^{-\lambda \times 1/\lambda} = \frac{1}{e}$$

Example 13. *1 milligram radium has 2.68×10^{18} atoms. Its half life period is 1620 years. How many radium atoms will disintegrate from 1 milligram of pure radium in 3240 years?*

Solution: No. of half lives in 3240 years $= \frac{3240}{1620} = 2$

$$\text{Amount of radium left after two half lives} = 1 \times \left(\frac{1}{2}\right)^2$$

$$= 0.25\text{ mg}$$

Amount of radium disintegrated $= (1 - 0.25) = 0.75$ mg

No. of atoms which have disintegrated $= 0.75 \times 2.68 \times 10^{18}$
$= 2.01 \times 10^{18}$

Example 14. *A certain radioisotope $^{A}_{Z}X$ (Half life = 10 days) decays to $^{A-4}_{Z-2}Y$. If 1 g atom of $^{A}_{Z}X$ is kept in sealed vessel, how much helium will accumulate in 20 days?*

Solution: $^{A}_{Z}X \longrightarrow {}^{A-4}_{Z-2}Y + {}^{4}_{2}\text{He}$

In two half lives, $\frac{3}{4}$ of the isotope $^{A}_{Z}X$ has disintegrated, *i.e.*, $\frac{3}{4}$ g atom of helium has been formed from $\frac{3}{4}$ g atom of $^{A}_{Z}X$.

Volume of 1 g atom of helium = 22400 mL

So, Volume of $\frac{3}{4}$ g atom of helium $= \frac{3}{4} \times 22400$ mL

$= 16800$ mL

Example 15. *Binding energy per nucleon of $^{2}_{1}H$ and $^{4}_{2}He$ are 1.1 MeV and 7 MeV respectively. Calculate the amount of energy released in the following process:*

$$^{2}_{1}H + {}^{2}_{1}H \longrightarrow {}^{4}_{2}He$$

Solution: Amount of energy released

= Σ Binding energy of products

− Σ Binding energy of reactants

$= [4 \times 7] - [4 \times 1.1]$

$= 23.6$ MeV

Example 16. *Calculate the energy associated with the following nuclear reaction:*

$$\underset{26.9815\ amu}{^{27}_{13}Al} + \underset{2.0141\ amu}{^{2}_{1}H} \longrightarrow \underset{24.9858\ amu}{^{25}_{12}Mg} + \underset{4.0026\ amu}{^{4}_{2}He}$$

Solution: Mass defect = (26.9815 + 2.0141)
− (24.9858 + 4.0026)
= 0.0072 amu

Energy of the reaction = 0.0072 × 931 MeV
= 6.70 MeV

Example 17. *A radioactive isotope $_{Z}A^{m}$ ($t_{1/2} = 10$ days) decays to give $_{Z-6}B^{m-12}$ stable atom alongwith α-particles. If m g of 'A' are taken and kept in a sealed tube, how much 'He' will accumulate in 20 days at STP?*

Solution: $_{Z}A^{m} \longrightarrow {}_{Z-6}B^{m-12} + 3[{}_{2}\text{He}^{4}]$

$$\text{Mole of } A = \frac{m}{m} = 1$$

Number of half lives = 20 / 10 = 2

$$N = N_0 \left(\frac{1}{2}\right)^n$$

$$= 1\left(\frac{1}{2}\right)^2 = \frac{1}{4}$$

Decayed moles $= 1 - 1/4 = 3/4$

Moles of 'He' formed $= 3 \times 3/4 = 9/4$

Volume of 'He' at STP $= 22.4 \times \frac{9}{4}$

$= 50.4$ litre

Example 18. *A sample of pitchblende is found to contain 50% uranium and 2.425% of lead. Of this lead only 93% was Pb^{206} isotope. If the disintegration constant is $1.52 \times 10^{-10}\ yr^{-1}$, how old could be the pitchblende deposits?*

Solution: Moles of $U^{238} = \frac{50}{100 \times 238} = 2.1 \times 10^{-3}$

Moles of $Pb^{206} = \frac{2.425}{100} \times \frac{93}{100 \times 206} = 0.109 \times 10^{-3}$

$N_0 = (x + y) = 2.1 \times 10^{-3} + 0.109 \times 10^{-3} = 2.209 \times 10^{-3}$

$N = x = 2.1 \times 10^{-3}$

$$\lambda = \frac{2.303}{t} \log\left(\frac{N_0}{N}\right)$$

$$1.52 \times 10^{-10} = \frac{2.303}{t} \log \frac{2.209 \times 10^{-3}}{2.1 \times 10^{-3}}$$

$$t = 3.3 \times 10^8 \text{ years}$$

Example 19. *On analysis, a sample of uranium ore was found to contain 0.277 g of $_{82}Pb^{206}$ and 1.667 g of $_{92}U^{238}$. The half life period of U^{238} is 4.51×10^9 yrs. If all the lead were assumed to have come from decay of $_{92}U^{238}$, what is the age of the earth?*

Solution: Moles of $U^{238} = \frac{1.667}{238}$

Moles of $Pb^{206} = \frac{0.277}{206}$

$$N_0 = \frac{1.667}{238} + \frac{0.277}{206}$$

and $$N = \frac{1.667}{238}$$

$$t = \frac{2.303}{\lambda} \log \frac{N_0}{N}$$

$$= \frac{2.303 \times 4.51 \times 10^9}{0.693} \log_{10} \frac{\frac{1.667}{238} + \frac{0.277}{206}}{\frac{1.667}{238}}$$

$= 1.143 \times 10^9$ years

Example 20. *$_{19}K^{40}$ consists of 0.012% potassium in nature. The human body contains 0.35% potassium by weight. Calculate the total radioactivity resulting from $_{19}K^{40}$ decay in a 75 kg human body. Half life of $_{19}K^{40}$ is 1.3×10^9 years.*

Solution:

Weight of radioactive potassium $= \frac{75000 \times 0.012}{100} \times \frac{0.35}{100}$

$= 0.0315$ g

$$\text{Activity} = \frac{0.693}{t_{1/2}} \times \frac{\text{Weight}}{\text{Atomic weight}} \times \text{Avogadro's number}$$

$$\text{Activity} = \frac{0.693}{1.3 \times 10^9 \times 365 \times 24 \times 60} \times \frac{0.0315}{40} \times 6.023 \times 10^{23}$$

$= 4.81 \times 10^5$ dpm

Example 21. *The sun radiates energy at the rate of 4×10^{26} J sec^{-1}. If the energy of fusion process is 27 MeV, calculate the amount of hydrogen that would be consumed per day for the given process.*

$$4\ {}_1^1H \longrightarrow {}_2^4He + 2\ {}_1^0e.$$

Solution: $27 \text{ MeV} = 27 \times 10^6 \times 1.6 \times 10^{-19}$

$= 43.2 \times 10^{-13}$ J

Energy radiated by the sun per day

$= 4 \times 10^{26} \times 3600 \times 24 \text{ J day}^{-1}$

$= 34.56 \times 10^{30} \text{ J day}^{-1}$

43.2×10^{-13} J of energy is obtained from

$= 4$ amu of H

$= 4 \times 1.66 \times 10^{-24}$ g of H

34.56×10^{30} J of energy is obtained from

$$= \frac{4 \times 1.66 \times 10^{-24}}{43.2 \times 10^{-13}} \times 34.56 \times 10^{30}$$

$= 5.31 \times 10^{19}$ g

Example 22. *A radioactive isotope X with half life of 1.37×10^9 years decays to Y, which is stable. A sample of rock from moon was found to contain both the elements X and Y in the ratio 1 : 7. What is the age of the rock?*

Solution: We know that,

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{t_{age}} \log\left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{1.37 \times 10^9} = \frac{2.303}{t_{age}} \log_{10}\left(\frac{1+7}{1}\right)$$

$$\therefore \quad t_{age} = 4.11 \times 10^9 \text{ years}$$

Example 23. *A sample of radioactive substance shows an intensity of 2.3 millicurie at a time 't' and an intensity of 1.62 millicurie, 600 seconds later. What is the half life period of the radioactive material?*

Solution: $$\lambda = \frac{2.303}{t} \log\left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{600} \log\left(\frac{2.3}{1.62}\right)$$

$\therefore$ $t_{1/2} = 1187$ seconds

Example 24. *What mass of ^{226}Ra, whose $t_{1/2} = 1620$ yrs will give the activity of 1 millicurie?*

Solution:

$$\text{Activity} = \frac{0.693}{t_{1/2}} \times \frac{w}{\text{At. wt.}} \times 6.023 \times 10^{23}$$

$$3.7 \times 10^7 = \frac{0.693}{1620 \times 365 \times 24 \times 3600} \times \frac{w}{226} \times 6.023 \times 10^{23}$$

so $w = 10^{-3}$ g

SUMMARY AND IMPORTANT POINTS TO REMEMBER

1. Radius of nucleus is calculated as:
$$R = R_0 A^{1/3},$$
where, $R_0 = 1.1 \times 10^{-15}$ m, A = Mass number of nucleus
Area of cross-section of a nucleus is expressed in barns ($1 \text{ barn} = 10^{-24} \text{ cm}^2$).

2. Nucleus density $\rho = \frac{3 \times \text{Mass}}{4\pi R_0^3}$
Density of all nuclei is constant, nuclear density is very large ($\approx 10^{17}$ kg/m^3) compared to atomic density ($\approx 10^3$ kg/m^3).

3. 1 amu $= 1.66 \times 10^{-27}$ kg
In terms of energy, 1 amu ≈ 931.5 MeV

4. Rate of radioactive decay is given as:
$$\text{Rate} = \lambda \times \frac{\text{Mass}}{\text{Atomic mass}} \times 6.023 \times 10^{23}$$
$$= \frac{0.693}{t_{1/2}} \times \frac{w}{\text{At. wt.}} \times 6.023 \times 10^{23}$$
   (i) Radioactivity is the phenomenon of spontaneous emission of certain radiations. It was discovered by Henri Becquerel in 1895.
   (ii) Marie Curie and her husband Piere Curie isolated two radioactive elements **polonium** and **radium.** Radium is 2 million times more reactive than uranium, it is the most radioactive element.
   (iii) Radium and polonium were isolated from pitchblende (U_3O_8).
   (iv) Francium is a liquid radioactive element in natural state.
   (v) Radon is a gaseous radioactive element in natural state.
   (vi) $^{238}_{92}U$ is the heaviest known natural element and it is radioactive.
   (vii) α-particles evolved from radioactive elements possess energy up to about 10 MeV. They can penetrate an aluminium sheet of 0.02 cm thickness.
   (viii) β-rays can penetrate an aluminium sheet up to 0.2 cm thickness.
   (ix) γ-rays are high energy electromagnetic radiations of short wavelength of the order of 10 pm. These are highly penetrating rays; they can penetrate up to 100 cm thick aluminium sheet.
   (x) After γ-decay, the daughter nuclide is the nuclear isomer of parent nuclide which differs in half life.
   (xi) Potassium uranyl sulphate $K(UO_2)(SO_4)_2$ was the first compound found to be radioactive.
   (xii) Tritium ^3_1H is the lightest radioactive element.

5. Units of rate of decay:
$$1 \text{ curie (Ci)} = 3.7 \times 10^{10} \text{ dis sec}^{-1}$$
$$1 \text{ millicurie (mCi)} = 3.7 \times 10^{7} \text{ dis sec}^{-1}$$
$$1 \text{ microcurie } (\mu\text{Ci}) = 3.7 \times 10^{4} \text{ dis sec}^{-1}$$
$$1 \text{ rutherford (Rd)} = 10^{6} \text{ dis sec}^{-1}$$
$$1 \text{ millicurie (mCi)} = 37 \text{ rutherford}$$
$$1 \text{ becquerel (Bq)} = 1 \text{ dis sec}^{-1}$$

6. Kinetic equation of radioactive decay:
$$N = N_0 e^{-\lambda t} \quad \text{(Exponential form)}$$
$$\lambda = \frac{2.303}{t} \log_{10}\left(\frac{N_0}{N}\right) \quad \text{(logarithmic form)}$$
λ = Decay constant
N_0 = Initial amount of radioactive element
N = Amount remaining after time 't'

7. Half life $t_{1/2} = \frac{0.693}{\lambda}$
$$\text{Average life } \tau = \frac{1}{\lambda} = \frac{t_{1/2}}{0.693}$$
$$\tau = 1.44 \times t_{1/2}$$

8. Amount remaining after 'n' half lives can be calculated as:
$$N = N_0 \left(\frac{1}{2}\right)^n$$
$$n = \frac{\text{Total time}}{\text{Half life}}$$
   (i) A radioactive element undergoes 50% decay in one half life.
   (ii) The time in which 63.2% radioactive element undergoes decay is called average life τ.
   (iii) The radioactive element undergoes 99.9% decay in 10 times of half life.
   (iv) An element undergoes 75% decay in twice of the half life.
   (v) Total life span of a radioactive element is infinite.

9. Some radioactive elements undergo α and β-decay in parallel path.

$$^{227}_{89}Ac \xrightarrow{K_1} {}^{223}_{87}Fr + {}^{4}_{2}He$$
$$^{227}_{89}Ac \xrightarrow{K_2} {}^{227}_{90}Th + {}^{0}_{-1}e$$

Overall decay constant $K = K_1 + K_2$

Fractional yield of Fr $= \dfrac{K_1}{K}$

Fractional yield of Th $= \dfrac{K_2}{K}$

10. At equilibrium,

$$A \longrightarrow B \longrightarrow C \ldots$$
$$\frac{\text{Amount of '}A\text{'}}{\text{Amount of '}B\text{'}} = \frac{\lambda_A}{\lambda_B} = \frac{t_{1/2}B}{t_{1/2}A}$$

11. If an element undergoes simultaneous α and β-decay, then

$$\lambda = \lambda_\alpha + \lambda_\beta; \quad \tau = \frac{\tau_\alpha \tau_\beta}{\tau_\alpha + \tau_\beta}$$

12. α-particles and γ-rays have line spectra, but β-particles have a continuous spectrum.

13. **Geiger-Muller counter** is used for detecting α and β-particles, **cloud chamber** is used for detecting radioactive radiations and for determining their **paths, range** and **energy.** In **scintillation counter,** the particles of radiations are detected by the flashes of light produced in the scintillator.

14. In every nuclear reaction representing transformation of one nucleus to other, the conservation of charge number, nucleons, energy and linear momentum is followed.

15. α-emission takes place when n/p ratio is lower than required for nuclear stability.

$$4\,{}^{1}_{1}H \longrightarrow {}^{4}_{2}He + 2[{}^{0}_{+1}e] + \text{energy}$$

α-particle emission shifts the daughter element two positions left in the periodic table.

$$^{M}_{Z}A - {}^{4}_{2}He \longrightarrow {}^{M-4}_{Z-2}B$$

(Here, A and B are **isodiapheres** to each other.)

16. β-emission takes place when n/p ratio is higher than the required value for nuclear stability.

$$^{1}_{0}n \longrightarrow {}^{1}_{1}H + {}^{0}_{-1}e + \text{antineutrino} + \text{energy}$$

Emission of β-particles increases the atomic number by one hence, the daughter element occupies one position right to the parent element.

$$^{M}_{Z}A - ({}^{0}_{-1}e) \longrightarrow {}_{Z+1}{}^{M}B$$

Here, A and B are **isobars;** thus β-emission is **isobaric** transformation.

17. In artificial radioactive elements, **positrons** are evolved when n/p ratio is lower than the required value for nuclear stability.

$$^{1}_{1}H \longrightarrow {}^{1}_{0}n + {}^{0}_{+1}e + \text{neutrino} + \text{energy}$$

Positron emission and K-electron capture are similar because both processes lower the number of proton by one unit.

18. $$^{M_1}_{Z_1}A \xrightarrow{(x\alpha,\, y\beta)} {}^{M_2}_{Z_2}B$$

Number of α-particles 'x' $= \dfrac{M_1 - M_2}{4}$ ... (i)

Number of β-particles can be calculated using the following relation:

$$Z_1 - 2x + y = Z_2 \quad \text{... (ii)}$$

19. There are three natural and one artificial decay series:

$$^{238}_{92}U \xrightarrow[(8\alpha,\, 6\beta)]{\text{Uranium series } (4n+2)} {}^{206}_{82}Pb$$
$$^{232}_{90}Th \xrightarrow[(6\alpha,\, 4\beta)]{\text{Thorium series } (4n)} {}^{208}_{82}Pb$$
$$^{235}_{92}U \xrightarrow[(7\alpha,\, 4\beta)]{\text{Actinium series } (4n+3)} {}^{207}_{82}Pb$$

Artificial series is also called neptunium series:

$$^{237}_{93}Np \xrightarrow[(7\alpha,\, 4\beta)]{(4n+1)} {}^{209}_{83}Bi$$

20. If both parent and daughter elements belong to actinide series (89–103) then they will belong to same group, *i.e.*, third group.

$$\underset{\text{3rd group}}{^{238}_{92}U} - {}^{4}_{2}He \longrightarrow \underset{\text{3rd group}}{^{234}_{90}Th}$$

21. $\xrightarrow{\alpha < \beta < \gamma}$ Increasing penetrating power

$\xrightarrow{\alpha > \beta > \gamma}$ Decreasing ionising power

$\xrightarrow{\alpha > \beta > \gamma}$ Decreasing luminosity on ZnS screen

22. Emission of one 'α' and two 'β' particles form an isotope of the parent element :

$$^{M}_{Z}A - {}^{4}_{2}He - 2[{}^{0}_{-1}e] \longrightarrow {}^{M-4}_{Z}A$$

23. There are only 81 stable elements having one or more non-radioactive isotopes.

24. No stable isotope exists for the elements above $^{209}_{83}Bi$. Thus, bismuth is the heaviest stable element.

25. Two elements earlier than bismuth (Tc and Pm) are radioactive.

26.

Isotope	Use
^{60}Co	Cancerous tumour detection and treatment.
^{131}I	Detection and treatment of thyroid disorders.
^{59}Fe	Anaemia.
^{32}P	Leucaemia and agriculture research.
^{24}Na	Location of blood clots and circulatory disorders.
^{74}As	Detection of presence of tumours.
^{90}Y	Treatment of joint effusion and arthritis.

^{18}O Used in the study of mechanism of photosynthesis.

Radioactive technetium Brain scan.

Note: Radioactive isotopes of carbon, chlorine and nitrogen are also used in the study of various reactions.

27. **Radiocarbon dating:** This method is used to determine age of wood.

$$\lambda = \frac{2.303}{t}\log_{10}\left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{t_{1/2}\ {}^{14}C} = \frac{2.303}{t_{age}}\log_{10}\left(\frac{N_0}{N}\right)$$

$N_0 = {}^{14}C/{}^{12}C$ in freshly cut wood or in the atmosphere or activity of freshly cut wood.

$N = {}^{14}C/{}^{12}C$ in the given sample of wood or activity of given sample of wood.

28. **Uranium dating or rock dating:** It is used to calculate the age of a sample of rock and mineral, *i.e.*, before how many years it was separated from the fire ball of earth.

$$\lambda = \frac{2.303}{t}\log_{10}\left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{t_{1/2}\ {}^{238}U} = \frac{2.303}{t_{age}}\log_{10}\left(\frac{N_0}{N}\right)$$

$$N_0 = \left(\frac{W}{238} + \frac{w}{206}\right)$$

$$N = \frac{W}{238}$$

where, W = amount of uranium in the sample
w = amount of ^{206}Pb in the sample

29. The force which binds the nucleons together in the nucleus is called nuclear force. These forces are short range forces operating over very small distances (1 fermi, 10^{-15} m). Nuclear forces are 10^{21} times stronger than electrostatic forces.

30. Hideki Yukawa of Japan discovered **mesons** in 1935. Protons and neutrons are held together by their fast mutual exchange.

$$^{1}_{1}H + {}^{0}_{-1}\pi \longrightarrow {}^{1}_{0}n$$
$$^{1}_{0}n + {}^{0}_{+1}\pi \longrightarrow {}^{1}_{1}H$$
$$^{1}_{1}H + {}^{0}_{0}\pi \longrightarrow {}^{1}_{1}H$$
$$^{1}_{0}n + {}^{0}_{0}\mu \longrightarrow {}^{1}_{0}n$$

31. **Artificial nuclear transmutation:** Conversion of one element to other by bombardment of a stable element with high speed subatomic particles. The first artificial transmutation was achieved by Rutherford in 1915 when he bombarded $^{14}_{7}N$ with α-particles emitted by $^{214}_{84}Po$.

$$^{14}_{7}N + {}^{4}_{2}He \longrightarrow {}^{17}_{8}O + {}^{1}_{1}H$$
$$^{27}_{13}Al + {}^{1}_{0}n \longrightarrow {}^{27}_{12}Mg + {}^{1}_{1}H$$

The nucleus bombarded is called **target;** the particles used for bombarding are called **projectiles** and the particles emitted are called **subsidiary particles.**

32. **Particle accelerator:** Various particle accelerators are used to give projectiles like protons, deuterons, α-particles and other cationic projectiles having sufficiently high kinetic energy to overcome the electrostatic repulsions of the target nuclei. Commonly used particle accelerators are linear accelerators, cyclotron and synchrotron. Synchrotron is used as proton accelerator.

33. Reactions of nuclear transformation are represented as:

$$^{9}_{4}Be + {}^{4}_{2}He \longrightarrow {}^{12}_{6}C + {}^{1}_{0}n \quad \text{or} \quad {}^{9}_{4}Be\,(\alpha n)\,{}^{12}_{6}C$$
$$^{23}_{11}Na + {}^{1}_{1}H \longrightarrow {}^{23}_{12}Mg + {}^{1}_{0}n \quad \text{or} \quad {}^{23}_{11}Na\,(pn)\,{}^{23}_{12}Mg$$
$$^{27}_{13}Al + {}^{1}_{0}n \longrightarrow {}^{24}_{11}Na + {}^{4}_{2}He \quad \text{or} \quad {}^{27}_{13}Al\,(n\alpha)\,{}^{24}_{11}Na$$

34. **Artificial radioactivity** was first studied by Irene Curie. In this process, a stable nucleus is converted to radioactive isotope on bombardment of suitable particle. Radioactive isotope produced undergoes artificial decay.

$$\underset{\text{Stable}}{^{27}_{13}Al} + {}^{4}_{2}He \longrightarrow \underset{\text{Unstable}}{^{30}_{15}P} + {}^{1}_{0}n$$
$$^{30}_{15}P \longrightarrow {}^{30}_{14}Si + {}^{0}_{+1}e \quad (t_{1/2} = 2.55\text{ min})$$

35. **Nuclear fission** is the process in which a heavy nucleus breaks up into two smaller nuclei on bombardment with neutrons. Energy is released in the process of fission along with freshly prepared neutrons.

$$^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{236}_{92}U \longrightarrow \begin{cases} {}^{140}_{56}Ba + {}^{93}_{36}Kr + 3\,{}^{1}_{0}n \\ {}^{144}_{54}Xe + {}^{90}_{38}Sr + 2\,{}^{1}_{0}n \\ {}^{144}_{55}Cs + {}^{90}_{37}Rb + 2\,{}^{1}_{0}n \end{cases}$$

Mass defect of the reaction is converted to huge amount of energy.

Δm (Mass defect) = Σ Masses of reactants − Σ Masses of products

Energy released = Δmc^2

If mass defect is 1 amu then 931.5 MeV energy is released.

36. **Critical mass:** It is the minimum mass of fissionable material required that will lead to a self-sustaining chain fission reaction. For $^{235}_{92}U$, the critical mass is between 1 to 100 kg.

37. The material which directly undergoes fission is termed as fissile material such as ^{235}U, ^{239}Pu and ^{233}U. The material which can be converted to fissile material is termed fertile material such as $^{238}_{92}U$ and $^{232}_{90}Th$.

38. Breeder reactors not only involve the fission of $^{235}_{92}U$ but also converts fertile material into fissile material, *e.g.*, ^{238}U is converted to ^{239}Pu:

$$^{238}_{92}U + {}^{1}_{0}n \longrightarrow {}^{239}_{92}U \xrightarrow{-{}^{0}_{-1}e} {}^{239}_{93}Np \xrightarrow{-\beta} {}^{239}_{94}Pu$$

39. Nuclear fission is a chain reaction. If it is uncontrolled, explosion occurs as in the atom bomb. Two or more pieces of fissile material (^{235}U or ^{239}Pu) having subcritical mass are brought together rapidly by means of conventional explosion. The subcritical masses combine to be supercritical and then chain fission starts, releasing large amount of energy.
40. The controlled chain fission reaction takes place in nuclear reactors. In these reactors the energy is used for peaceful purposes. The heat energy produced in the nuclear reactors can be used to generate electricity. A reactor consists of:
    (i) enriched fuel $^{235}_{92}U$(2–3%).
    (ii) heavy water (D_2O) or graphite moderator. It slows down the speed of fast moving neutrons.
    (iii) control rods made of boron and cadmium. These rods absorb some neutrons and thereby control the rate of nuclear fission.
    (iv) liquid alloy of sodium and potassium is used as a coolant.
41. Nuclear fusion is the process in which two nuclei of light atoms fuse to form heavy nuclei with the liberation large amount of energy.

$$^2_1H + ^2_1H \longrightarrow ^4_2He + 23 \times 10^8 \text{ kJ/mol}$$

$$^2_1H + ^3_1H \longrightarrow ^4_2He + ^1_0n + 17.2 \times 10^8 \text{ kJ/mol}$$

42. Fusion reactions are thermonuclear reactions which require very high temperature (10^6 K or more).
43. Hydrogen bomb involves nuclear fusion.
44. Energy of a star (sun) is due to nuclear fusion; this energy is called stellar energy.
45. Hydrogen bomb is much more powerful than atom bomb and there is no restriction of critical mass in this bomb.
46. Neutron activation analysis is a technique of finding the trace amount of an element present with the other. The trace element is activated by bombarding with neutrons. It is a non-destructive method, *e.g.*, traces of silver present in lead paintings can be detected by neutron activation analysis.
47. **Spallation reactions:** It is similar to fission but differ in the fact that they are brought about by high energy bombarding particles or photons. A number of smaller particles are released along with the product, *e.g.*,

$$^{238}_{92}U + ^4_2He \longrightarrow 6\, ^1_1H + 13\, ^1_0n + ^{223}_{88}Ra$$

48. The isotope ^1_1H has $n/p = 0$ and ^3_1H has $n/p = 2$ which is maximum.
49. Only $^{12}_6C$ has zero packing fraction. Packing fraction is maximum for hydrogen and minimum for iron.

$$\text{Packing fraction} = \frac{\text{Isotopic mass} - \text{Mass number}}{\text{Mass number}} \times 10^4$$

Elements with negative packing fraction are stable because some of their mass is converted to binding energy.
50. $^{14}_6C$ is produced in upper atmosphere due to bombardment of cosmic ray neutrons on atmospheric nitrogen.

Questions

1. Match the List-I and List-II and pick the correct answer from the codes given below: **[PMT (Kerala) 2006]**

List-I (Atomic/Molecular species)	**List-II (Corresponding pairs)**
(A) Isotopes	1. $^{228}_{88}Ra$ and $^{228}_{89}Ac$
(B) Isobars	2. $^{39}_{18}Ar$ and $^{40}_{19}K$
(C) Isotones	3. $^{2}_{1}H$ and $^{3}_{1}H$
(D) Isosters	4. $^{235}_{92}U$ and $^{231}_{90}Th$
(E) Isodiapheres	5. CO_2 and N_2O

(a) A—2, B—1, C—4, D—5, E—3
(b) A—2, B—5, C—1, D—4, E—3
(c) A—3, B—1, C—2, D—5, E—4
(d) A—5, B—4, C—1, D—2, E—3
(e) A—5, B—3, C—1, D—2, E—4

2. **Matrix Matching Problems (For IIT aspirants):**

[A] Match the Column-I with Column-II:

Column-I	**Column-II**
(a) Stability of nucleus	(p) Depends on mass number
(b) Density of nucleus	(q) Packing fraction
(c) Spin angular momentum of proton	(r) Binding energy per nucleon
(d) Dimensionless quantity	(s) Independent of mass number

[B] Match the Column-I with Column-II:

Column-I	**Column-II**
(a) 2/3rd life	(p) 63.2% decay
(b) Average life	(q) 75% decay
(c) $1/\lambda$	(r) $2 \times t_{1/2}$
(d) Ten times of half life	(s) 99.9% decay

[C] Match the nuclear transformations of Column-I with the particles emitted of Column-II:

Column-I	**Column-II**
(a) $^{209}_{83}Bi + ^{4}_{2}He \longrightarrow ^{211}_{85}At + ...$	(p) $^{1}_{1}H$
(b) $^{9}_{4}Be + ^{4}_{2}He \longrightarrow ^{12}_{6}C + ...$	(q) $^{4}_{2}He$
(c) $^{24}_{12}Mg(^{1}_{0}n ...)^{24}_{11}Na$	(r) $^{2}_{1}H$
(d) $^{23}_{11}Na(^{2}_{1}H ...)^{21}_{10}Ne$	(s) $^{1}_{0}n$

[D] Match the Column-I with Column-II:

Column-I	**Column-II**
(a) Binding energy per nucleon increases	(p) β-decay
(b) Mass number is conserved	(q) α-decay
(c) Charge number is conserved	(r) Nuclear fusion
(d) Mass of products formed is less than the mass of reactants	(s) Nuclear fission

[E] Match the Column-I with Column-II:

Column-I	**Column-II**
(a) α-rays	(p) Radiations, undeviated in electric field
(b) β-rays	(q) Produced when electrons strike metal surface
(c) γ-rays	(r) Highest deflection in electromagnetic field
(d) X-rays	(s) Nucleus of helium

[F] Match the Column-I with Column-II:

Column-I	**Column-II**
(a) α-emission	(p) Mass number changes
(b) β-emission	(q) Atomic number and mass number are affected
(c) γ-emission	(r) Atomic number decreases
(d) β^+-emission	(s) Atomic number increases

3. Write the complete nuclear reactions:

(a) $^{9}_{4}Be + ^{4}_{2}He \longrightarrow ^{12}_{6}C + ...$
(b) $^{3}_{1}H \longrightarrow ^{3}_{2}He + ...$
(c) $^{14}_{7}N + ^{4}_{2}He \longrightarrow ^{17}_{8}O + ...$
(d) $^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{92}_{38}Sr + Xe + 3\ ^{1}_{0}n$
(e) $^{7}_{3}Li + ^{1}_{0}n \longrightarrow 2\ ^{4}_{2}He + ...$
(f) $^{238}_{92}U + ... \longrightarrow ^{239}_{92}U \longrightarrow ^{239}_{93}Np + ...$
(g) $^{14}_{7}N + ^{1}_{0}n \longrightarrow ^{3}_{1}H + ...$
(h) $^{7}_{3}Li + ... \longrightarrow ^{8}_{4}Be + \gamma$ - radiations
(i) $^{2}_{1}H + ... \longrightarrow ^{4}_{2}He + ^{1}_{0}n$
(j) $^{27}_{13}Al + ^{1}_{0}n \longrightarrow ^{24}_{11}Na + ...$
(k) $^{27}_{13}Al + ^{4}_{2}He \longrightarrow .... + ^{1}_{1}H$
(l) $^{235}_{92}U + ^{1}_{0}n \longrightarrow .... + ^{137}_{52}Te + ^{97}_{40}Zr$ **(IIT 2005)**
(m) $^{86}_{34}Se \longrightarrow .... + 2\ ^{0}_{-1}e$ **(IIT 2005)**

4. Write the particles emitted from each nuclide in the following reactions:

(a) $^{231}_{90}Th \xrightarrow[\text{(i)}]{} ^{231}_{91}Pa \xrightarrow[\text{(ii)}]{} ^{227}_{89}Ac$
(b) $^{217}_{85}At \xrightarrow[\text{(i)}]{} ^{213}_{83}Bi \xrightarrow[\text{(ii)}]{} ^{209}_{81}Tl$
(c) $^{239}_{92}U \xrightarrow[\text{(i)}]{} ^{239}_{93}Np \xrightarrow[\text{(ii)}]{} ^{239}_{94}Pu$
(d) $^{30}_{15}P \xrightarrow[\text{(i)}]{} ^{30}_{14}Si$

5. Find the atomic number and mass number of the last member in the following series:

(a) $^{226}_{188}Ra \xrightarrow{-\alpha} Rn \xrightarrow{-\alpha} RaA \xrightarrow{-\alpha} RaB \xrightarrow{-\beta} RaC$

(b) $^{M}_{Z}A \xrightarrow{-\alpha} B \xrightarrow{-\beta} C \xrightarrow{-\beta} D \xrightarrow{-\alpha} E$

6. Complete the following:

(a) $^{238}X \xrightarrow{-\alpha} Y \xrightarrow{-\beta} {}_{91}Z$

(b) $^{214}A \xrightarrow{-\beta} B \xrightarrow{-\beta} {}_{84}C$

7. Write the equations for the following transformations:

(a) $^{39}_{19}K\,(p, d)$ (b) $^{14}_{7}N\,(n, p)$ (c) $^{23}_{11}Na\,(\alpha, p)$

(d) $^{9}_{4}Be\,(\alpha, n)$.

8. To which radioactive families do the following nuclides belong?

^{222}Rn, ^{228}Ra, ^{207}Pb, ^{209}Bi, ^{233}Pa.

9. To which group of the periodic table does the last member of the following series belong?

(a) $^{239}_{92}U \xrightarrow{-\beta} {}^{239}_{93}Np \xrightarrow{-\beta} {}^{239}_{94}Pu$

III Group

(b) $^{140}_{55}Cs \xrightarrow{-\beta} {}^{140}_{56}Ba \xrightarrow{-\beta} {}^{140}_{57}La \xrightarrow{-\beta} {}^{140}_{58}Ce$

I Group

(c) $^{226}_{88}Ra \xrightarrow{-\alpha} {}^{222}_{86}Rn \xrightarrow{-\alpha} {}^{218}_{84}Po$

II Group

10. Name the process represented below:

(a) $^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{141}_{56}Ba + {}^{92}_{36}Kr + 3\,{}^{1}_{0}n + 200$ MeV

(b) $^{2}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + {}^{1}_{0}n + 17.6$ MeV

(c) $^{63}_{29}Cu + {}^{4}_{2}He + 400$ MeV $\longrightarrow {}^{37}_{17}Cl + 14\,{}^{1}_{1}H + 16\,{}^{1}_{0}n$

(d) $^{10}_{5}B + {}^{4}_{2}He \longrightarrow {}^{13}_{7}N + {}^{1}_{0}n$

$\quad\quad\quad\quad\quad\quad\quad\quad \longrightarrow {}^{13}_{6}C + {}^{0}_{+1}e$

Answers

1. (c) A–3, B–1, C–2, D–5, E–4

2. [A] (a — p, q, r); (b — s); (c — s); (d — q)
[B] (a — q, r); (b — p); (c — p); (d — s)
[C] (a — s); (b — s); (c — p); (d — q)
[D] (a–p, q, r, s); (b–p, q, r, s); (c–p, q, r, s); (d–p, q, r, s)
[E] (a — s); (b — r); (c — p); (d — p, q)
[F] (a — p, r) (b — s) (c — q) (d — q)

3. (a) $^{9}_{4}Be + {}^{4}_{2}He \longrightarrow {}^{12}_{6}C + {}^{1}_{0}n$;

(b) $^{3}_{1}H \longrightarrow {}^{3}_{2}He + {}^{0}_{-1}e$;

(c) $^{14}_{7}N + {}^{4}_{2}He \longrightarrow {}^{17}_{8}O + {}^{1}_{1}H$;

(d) $^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{92}_{38}Sr + {}^{141}_{54}Xe + 3\,{}^{1}_{0}n$;

(e) $^{7}_{3}Li + {}^{1}_{0}n \longrightarrow 2\,{}^{4}_{2}He + {}^{0}_{-1}e$;

(f) $^{238}_{92}U + {}^{1}_{0}n \longrightarrow {}^{239}_{92}U \longrightarrow {}^{239}_{93}Np + {}^{0}_{-1}e$;

(g) $^{14}_{7}N + {}^{1}_{0}n \longrightarrow {}^{12}_{6}C + {}^{3}_{1}H$;

(h) $^{7}_{3}Li + {}^{1}_{1}H \longrightarrow {}^{8}_{4}Be + \gamma$-radiations;

(i) $^{2}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + {}^{1}_{0}n$;

(j) $^{27}_{13}Al + {}^{1}_{0}n \longrightarrow {}^{24}_{11}Na + {}^{4}_{2}He$;

(k) $^{27}_{13}Al + {}^{4}_{2}He \longrightarrow {}^{30}_{14}Si + {}^{1}_{1}H$;

(l) $^{235}_{92}U + {}^{1}_{0}n \longrightarrow 2\,{}^{1}_{0}n + {}^{137}_{52}Te + {}^{97}_{40}Zr$;

(m) $^{86}_{34}Se \longrightarrow {}^{86}_{36}Kr + 2\,{}^{0}_{-1}e$

4. (a) (i) β (ii) α (b) (i) α (ii) α (c) (i) β (ii) β (d) (i) ${}^{0}_{+1}e$

5. (a) Atomic mass = 214, Atomic number = 83
(b) Atomic mass = $M - 8$, Atomic number = $Z - 2$

6. (a) $^{238}_{92}X \xrightarrow{-\alpha} {}^{234}_{90}Y \xrightarrow{-\beta} {}^{234}_{91}Z$

(b) $^{214}_{82}A \xrightarrow{-\beta} {}^{214}_{83}B \xrightarrow{-\beta} {}^{214}_{84}C$

7. (a) $^{39}_{19}K + {}^{1}_{1}H \longrightarrow {}^{38}_{19}K + {}^{2}_{1}H$ (b) $^{14}_{7}N + {}^{1}_{0}n \longrightarrow {}^{14}_{6}C + {}^{1}_{1}H$

(c) $^{23}_{11}Na + {}^{4}_{2}He \rightarrow {}^{26}_{12}Mg + {}^{1}_{1}H$ (d) $^{9}_{4}Be + {}^{4}_{2}He \rightarrow {}^{12}_{6}C + {}^{1}_{0}n$

8. ^{222}Rn belongs to $(4n + 2)$ family, *i.e.*, uranium family.
^{228}Ra belongs to $(4n)$ family, *i.e.*, thorium family.
^{207}Pb belongs to $(4n + 3)$ family, *i.e.*, actinium family.
^{209}Bi belongs to $(4n + 1)$ family, *i.e.*, neptunium series.
^{233}Pa belongs to $(4n + 1)$ family, *i.e.*, neptunium series.

9. (a) $^{239}_{94}Pu$ belongs to actinide series, hence it is present in III group.

(b) $^{140}_{58}Ce$ belongs to lanthanide series, hence it is present in III group.

(c) $^{218}_{84}Po$ belongs to VI group.

10. (a) Nuclear fission (b) nuclear fusion (c) spallation reaction (d) artificial radioactivity.

• PRACTICE PROBLEMS •

1. Half life of ^{24}Na is 14.8 hours. In what period of time will a sample of this element lose 90% of its activity?
[**Ans.** 49.17 hour]

2. A β-particle emitter has a half life of 60.6 min. At any instant of time, a sample of this element registers 2408 counts per second. Calculate the counting rate after 1.5 hours.
[**Ans.** 860 counts per sec]

3. A radio-isotope $^{32}_{15}P$ has half life of 15 days. Calculate the time in which the radioactivity of 1 mg quantity will fall to 10% of the initial value.
[**Ans.** 49.85 days]

4. Consider an α-particle just in contact with $^{238}_{92}U$ nucleus. Calculate the coulombic repulsion energy assuming that the distance between them is equal to the sum of their radii.
[**Ans.** 24.2 MeV]

5. The activity of a certain sample of radioactive element 'A' decreases to $1/\sqrt{2}$ of its value in 4 days. What is its half life? Assuming that,

$$^{M}_{Z}A - {}^{4}_{2}He \longrightarrow {}^{M-4}_{Z-2}B$$

what mass of the sample will be left over after 24 days if we start with one gram of 'A'? Calculate this in terms of M.

[**Ans.** $t_{1/2} = 8$ days; mass of sample left over $= \left(1 - \frac{3-5}{M}\right)$ g]

[**Hint:** $\lambda = \frac{2.303}{t}\log_{10}\left(\frac{N_0}{N}\right)$

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{4}\log_{10}(\sqrt{2}) \quad \because \quad N = \frac{N_0}{\sqrt{2}}$$

$$t_{1/2} = 8 \text{ days}$$

Use $N = N_0\left(\frac{1}{2}\right)^n$ for next part.]

6. The half life of $^{238}_{92}U$ is 4.5×10^9 years. Uranium emits an α-particle to give thorium. Calculate the time required to get the product which contains equal masses of thorium and uranium.
[**Ans.** 4.55×10^9 yrs]

[**Hint:** $N_0 = \frac{1}{238} + \frac{1}{234}$; $N = \frac{1}{238}$

Use, $\frac{0.693}{t_{1/2}} = \frac{2.303}{t}\log_{10}\left(\frac{N_0}{N}\right)$]

7. 32 mg of pure $^{238}_{94}PuO_2$ has an activity of $6.4 \times 10^7\ sec^{-1}$.
(i) What will be the half life of $^{238}_{94}Pu$ in years?
(ii) What amount of PuO_2 will remain if 100 mg PuO_2 is kept for 5000 years?
[**Ans.** (i) 2.45×10^4 years (ii) 86.7 mg]

[**Hint:** (i) Mass of $^{238}Pu = \frac{238}{270} \times 32 = 28.207$ mg

$$\text{Rate} = \frac{0.693}{t_{1/2}} \times \frac{w}{\text{At. wt.}} \times N$$

$$6.4 \times 10^7 = \frac{0.693}{t_{1/2}} \times \frac{28.207 \times 10^{-3}}{238} \times 6.023 \times 10^{23}$$

$$t_{1/2} = 7.729 \times 10^{11} \text{ sec} = 2.45 \times 10^4 \text{ years}$$

(ii) Use, $\frac{0.693}{t_{1/2}} = \frac{2.303}{t}\log_{10}\left(\frac{N_0}{N}\right)$

$$\frac{0.693}{2.45 \times 10^4} = \frac{2.303}{5000}\log_{10}\left(\frac{100}{N}\right)$$

$N = 86.7$ mg]

8. A radioactive isotope decays as:

$$^{M}_{Z}A \longrightarrow {}^{M-4}_{Z-2}B \longrightarrow {}^{M-4}_{Z-1}C$$

The half lives of A and B are 6 and 10 months respectively. Assuming that initially only A was present, will it be possible to achieve the radioactive equilibrium for B? If so, what would be the ratio of A and B at equilibrium? What would happen if the half lives of A and B were 10 and 6 months respectively?
[**Hint:** At equilibrium, ratio of amounts of A and B will be

$$\frac{N_A}{N_B} = \frac{t_{1/2}A}{t_{1/2}B} = \frac{6}{10} = 0.6$$

If the half lives of A and B are 10 and 6 months respectively, then B will decay faster than 'A', hence equilibrium will not be achieved.]

9. Lowest level of ^{14}C activity for experimental detection is 0.03 dis per min per gram. What is the maximum age of an object that can be determined by ^{14}C method? The activity of ^{14}C in the atmosphere is 15 dis per min per gram of ^{14}C ($t_{1/2}$ for $^{14}C = 5730$ yrs).
[**Ans.** 51379.28 yrs]

10. An analysis of a rock shows that relative number of ^{87}Sr and ^{87}Rb atoms is 0.052, *i.e.*, ($^{87}Sr/^{87}Rb = 0.052$). Determine the age of the rock. Given that half life period for β-decay of Rb to Sr is 4.7×10^{10} years.
[**Ans.** 3.43×10^9 years]

[**Hint:**

$$\frac{^{87}Rb}{^{87}Sr} = \frac{x}{y} = \frac{1}{0.052}$$

$$\frac{x}{x+y} = \frac{1}{1.052}$$

$$\frac{N}{N_0} = \frac{1}{1.052}$$

$$\lambda = \frac{2.303}{t}\log_{10}\frac{N_0}{N}$$

$$\frac{0.693}{4.7 \times 10^{10}} = \frac{2.303}{t}\log_{10} 1.052$$

$t = 3.43 \times 10^9$ years]

11. Hydrolysis of ester was studied by isotopic labelling method. Write down the structures of products A and B in the given reaction: **(IIT 2000)**

$$CH_3-\overset{O}{\overset{||}{C}}-\overset{18}{O}-H + HOH \longrightarrow A + B$$

[**Ans.** $(A)\ CH_3-\overset{O}{\overset{||}{C}}-OH;\ (B)\ C_2H_5-\overset{18}{O}-H$]

12. Arrange the following species in decreasing order of chemical reactivity and radioactivity:

$${}^1_1H,\ {}^2_1H,\ {}^3_1H$$

[**Ans.** Reactivity ${}^1_1H > {}^2_1H > {}^3_1H$

Radioactivity ${}^3_1H > {}^2_1H > {}^1_1H$]

13. The half life of ${}^{212}Pb$ is 10.6 hours and that of its daughter element ${}^{212}Bi$ is 60.5 minutes. After how much time will the daughter element have maximum activity?

[**Ans.** 3.78 hours]

[**Hint:** $\lambda_p = \dfrac{0.693}{10.6 \times 60}$; $\lambda_d = \dfrac{0.693}{60.5} = 0.01145\ \text{min}^{-1}$

$= 0.001089\ \text{min}^{-1}$

$$t_{max} = \frac{2.303}{(\lambda_d - \lambda_p)} \log_{10} \frac{\lambda_d}{\lambda_p}$$

$$= \frac{2.303}{0.01145 - 0.001089} \log_{10} \left[\frac{0.01145}{0.001089}\right]$$

$$= 222.2758 \log_{10} \frac{0.01145}{0.001089}$$

$= 227.1$ min

$= 3.785$ hours]

14. Radioactive element is spread over a room, its half life is 30 days. Its activity is 50 times the permissible value. After how many days will it be safe?

[**Ans.** 169.30 days]

[**Hint:** $N_0 = 50N$

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{t} \log_{10} \left(\frac{N_0}{N}\right)$$

$$\frac{0.693}{30} = \frac{2.303}{t} \log_{10} \left(\frac{50N}{N}\right)$$

$t = 169.3$ days]

15. Calculate the energy released in joules and MeV in the following nuclear reaction:

$${}^2_1H + {}^2_1H \longrightarrow {}^3_2He + {}^1_0n$$

Assume that the masses of ${}^2_1H$, ${}^3_2He$ and neutron respectively are 2.0141, 3.0160 and 1.0087 amu.

[**Ans.** 5.223×10^{-13} J; 3.260 MeV]

16. A radioactive element due to an accident in research laboratory gets embedded in its floor and walls. The initial rate of decay is 64 times the safe limit. The half life of the element is 32 days. Calculate the time after which the laboratory will be safe for use.

[**Ans.** 192 days]

[**Hint:** $N_0 = 64N$

$$\frac{0.693}{t_{1/2}} = \frac{2.303}{t} \log_{10} \frac{N_0}{N}$$

$$\frac{0.693}{32} = \frac{2.303}{t} \log_{10} \left[\frac{64N}{N}\right]$$

$t = 192$ days]

17. Radium has a half life 1600 years and its daughter element radon has a half life 3.82 days. In an enclosure, the volume of radon was found constant for a week. Explain and calculate the ratio of the number of radium and radon nuclei. Will the ratio be constant after 400 years?

[**Ans.** 1.528×10^5]

[**Hint:** $\dfrac{N_1(\text{Ra})}{N_2(\text{Rn})} = \dfrac{t_{1/2}(\text{Rn})}{t_{1/2}(\text{Ra})}$

$$\frac{N_1}{N_2} = \frac{1600}{3.82} \times 365$$

$= 1.528 \times 10^5$]

18. Calculate the radius and density of ${}^{235}_{92}U$.

[**Ans.** $R = 6.8 \times 10^{-13}$ cm; $d = 2.979 \times 10^{14}$ g / cc]

[**Hint:** $R = R_0 A^{1/3} = 1.1 \times 10^{-15} (235)^{1/3} = 6.788 \times 10^{-15}$ m

$= 6.788 \times 10^{-13}$ cm

$$d = W/V = \frac{235 \times 1.66 \times 10^{-24}}{\frac{4}{3} \times \pi \times (6.788 \times 10^{-13})^3} \text{ g / cc}$$

$= 2.979 \times 10^{14}$ g / cc]

19. ${}^{235}_{92}U$ decays with emission of α and β-particles to form ultimately ${}^{207}_{82}Pb$. How many α and β-particles are emitted per atom of Pb produced?

[**Ans.** 7 α and 4 β]

20. The half life of radium is 1600 years. After how much time, $\frac{1}{16}$th part of radium will remain undisintegrated in a sample?

[**Ans.** 6400 years]

21. The half life of polonium is 140 days. In what time will 15 g of polonium be disintegrated out of its initial mass of 16 g?

[**Ans.** 560 days]

[**Hint:** Polonium left is $\frac{1}{16}$th of the initial, *i.e.*, 4 half lives.]

22. The activity of a radioactive isotope falls to 12.5% in 90 days. Calculate the half life and decay constant of the radioactive isotope.

[**Ans.** 30 days, 0.0231 day^{-1}]

23. The radioactivity of an element was found to be one millicurie. What will be its radioactivity after 42 days if it has half life of 14 days?

[**Ans.** 0.125 millicurie]

24. There are 10^6 radioactive nuclei in a given radioactive element. Its half life is 20 seconds. How many nuclei will remain after 10 seconds? (Given, $\sqrt{2} = 1.41$)

[**Ans.** 7×10^5 (approximately)]

[**Hint:** $N = N_0 \left(\frac{1}{2}\right)^n = 10^6 \times \left(\frac{1}{2}\right)^{1/2}$ as $n = \frac{10}{20} = \frac{1}{2} = \frac{10^6}{1.41}$]

25. A radioactive element decays at such a rate that after 68 minutes only one-fourth of its original amount remains. Calculate its decay constant and half life period.

[**Ans.** $\lambda = 0.0204 \text{ min}^{-1}$, $t_{1/2} = 34$ min.]

26. One gram of a radioactive element decays by β-emission to 0.125 in 200 hours. How much more time will elapse until only 0.10 g of it is left?

[**Ans.** 21.46 hours]

27. A wooden article found in a cave has only 40% as much ^{14}C activity as a fresh piece of wood. How old is the article? ($t_{1/2}$ for $^{14}C = 5760$ years)

[**Ans.** 7617 years]

28. A sample of carbon derived from one of dead sea scrolls is found to be decaying at the rate of 12.0 disintegrations per minute per gram of carbon. Estimate the age of dead sea scrolls when carbon from living plants disintegrates at the rate of 15.3 disintegrations per minute per gram. ($t_{1/2}$ for $^{14}C = 5760$ years)

[**Ans.** 2020 years]

29. One μg of a radioactive iodine contained in thyroxine is injected into the blood of a patient. How long will it take for radioactivity to fall to 50%, 25% and 10% of the initial value? ($t_{1/2}$ for $^{131}_{53}I = 8.05$ days)

[**Ans.** 8.05 days, 16.1 days, 26.75 days]

30. 1 g radium is reduced by 2.1 mg in 5 years by alpha decay, calculate the half life period.

[**Ans.** Half life = 1672 years]

[**Hint:** Mass of radium left after 5 years = (1.0 − 0.0021) g

$$= 0.9979 \text{ g}$$

Apply $\lambda = \frac{2.303}{t} \log_{10} \frac{N_0}{N} = \frac{2.303}{5} \log_{10} \frac{1}{0.9979}$]

31. The activity of a radioactive substance falls to 87.5% of the initial value in 5 years. What is the half life of the element? Calculate the time in which the activity will fall by 87.5%.

[**Ans.** Half life = 9.52 years, t = 28.58 years]

32. Starting with 1.0 g of a radioactive sample, 0.25 g of it is left after 5 days. Calculate the amount which was left after one day.

[**Ans.** 0.758 g]

33. A sample of wooden artifact is found to undergo 9 disintegrations per minute per gram of carbon. What is the approximate age of the artifact? The half life of $^{14}_{6}C$ is 5730 years and radioactivity of wood recently cut is 15 disintegrations per minute per gram of carbon.

34. Xenon-127 has a half life of 36.4 days. How much of a sample of xenon that originally weighed 1.0 g remains after 20 days? **(Dhanbad 1992)**

[**Ans.** 0.6835 g]

35. Calculate the ratio of $\frac{N}{N_0}$ after an hour has passed for a radioactive material of half life 47.2 seconds.

[**Ans.** $\frac{N}{N_0} = 1.12 \times 10^{-23}$]

[**Hint:** $\lambda t = 2.303 \log_{10} \frac{N_0}{N}$

$$\frac{0.693}{47.2} \times \frac{60 \times 60}{2.303} = \log_{10} \frac{N_0}{N}$$]

36. The activity of the hair of an Egyptian mummy is 7 disintegrations minute^{-1} of ^{14}C. Find the age of the mummy. Given, $t_{1/2}$ of ^{14}C is 5770 years and disintegration rate of fresh sample of ^{14}C is 14 disintegrations minute^{-1}.

[**Ans.** 5770 years]

37. On analysis a sample of ^{238}U ore was found to contain 20.6 g of $^{206}_{82}Pb$ and 23.8 g of $^{238}_{92}U$. The half life period of ^{238}U is 4.50×10^9 years. If all the lead were assumed to have come from decay of $^{238}_{92}U$, what is the age of the ore? **(IIT 1996)**

[**Ans.** 4.49×10^9 years]

38. It is known that 1 g of ^{226}Ra emits 11.6×10^{17} atoms of α per year. Given, the half life of ^{226}Ra to be 1600 years, compute the value of Avogadro's number.

[**Ans.** 6.052×10^{23}]

[**Hint:** Rate = λ × number of atoms in one gram

$$= \lambda \times \frac{\text{Avogadro's number}}{226}$$]

39. A uranium mineral contains ^{238}U and ^{206}Pb in the ratio of 4 : 1 by weight. Calculate the age of the mineral $t_{1/2}$ ^{238}U $= 4.5 \times 10^9$ years. Assume that all the lead present in the mineral is formed from disintegration of ^{238}U.

[**Ans.** 1.648×10^9 years]

40. In a sample of pitchblende, the atomic ratio of $^{206}Pb : {}^{238}U$ is 0.23 : 1. Calculate the age of the mineral if half life of uranium is 4.5×10^9 years. Assume that all lead has originated from uranium.

[**Ans.** 1.34×10^9 years]

41. The ratio of the atoms of two elements A and B at radioactive equilibrium is 5.0×10^5 : 1 respectively. Calculate half life of B if half life of A is 245 days.

[**Ans.** 4.9×10^{-4} days]

42. Calculate the energy released in MeV during the reaction $^7_3Li + {}^1_1H \rightarrow 2[^4_2He]$ if the masses of ^7_3Li, ^1_1H and ^4_2He are 7.018, 1.008 and 4.004 amu respectively.

[**Ans.** 16.76 MeV]

43. 1.0 g of $^{226}_{88}Ra$ is placed in a sealed vessel. How much helium will be collected in the vessel in 100 days? ($t_{1/2}$ of radium = 1600 years)

[**Ans.** 2.12×10^{-6} g]

44. The half life period of $^{141}_{58}Ce$ is 13.11 days. It is a β-particle emitter and the average energy of the β-particle emitted is 0.442 MeV. What is the total energy emitted per second in watts by 10 mg of $^{141}_{58}Ce$?

[**Ans.** 1.84 watt]

[**Hint:** Rate of disintegrations per sec = $\lambda \times$ No. of atoms

$$= \frac{0.693}{13.11 \times 24 \times 60 \times 60} \times \frac{6.023 \times 10^{23}}{141} \times 0.01$$

Total β-particles emitted $= 2.61 \times 10^{13}$

Total energy emitted $= 2.61 \times 10^{13} \times 0.442 = 1.1536 \times 10^{13}$ MeV

$$\text{Energy in erg} = (1.1536 \times 10^{13})(1.6 \times 10^{-6})$$

$$\text{Energy in watt} = \frac{1.1536 \times 10^{13} \times 1.6 \times 10^{-6}}{10^7} = 1.84 \text{ watt}]$$

45. A sample of $^{90}_{38}Sr$ has an activity of 0.5 mCi. What is its specific activity? ($t_{1/2}$ of $^{90}_{38}Sr$ = 19.9 years)

[**Ans.** 7.4×10^{12} dis g^{-1} s^{-1}]

[**Hint:** Rate of disintegrations = $\lambda \times$ No. of atoms

So, No. of atoms

$$= \frac{0.5 \times 3.7 \times 10^7}{0.693} \times 19.9 \times 365 \times 24 \times 60 \times 60$$

$$= 1.675 \times 10^{16}$$

$$\text{Mass} = \frac{90 \times 1.675 \times 10^{16}}{6.023 \times 10^{23}} = 2.50 \times 10^{-6} \text{ g}$$

$$\text{Specific activity} = \frac{0.5 \times 3.7 \times 10^7}{2.5 \times 10^{-6}} = 7.4 \times 10^{12} \text{ dis g}^{-1} \text{ s}^{-1}]$$

46. Calculate the Q-value of the reaction;

$$^6_3Li + ^1_0n \longrightarrow ^4_2He + ^3_1H$$

Given, $^6_3Li = 6.015126$ amu, $^4_2He = 4.002604$ amu

$^3_1H = 3.016049$ amu, $^1_0n = 1.008665$ amu

[**Ans.** +4.7835 MeV]

47. The disintegration rate of a certain radioactive sample at any instant is 4750 dpm. Five minutes later, the rate becomes 2700 dpm. Calculate half life of sample.

[**Ans.** $t_{1/2} = 6.13$ minute]

48. One of the hazards of nuclear explosion is the generation of ^{90}Sr and its subsequent incorporation in bones. This nuclide has a half life of 28.1 years. Suppose one microgram was absorbed by a new-born child, how much ^{90}Sr will remain in his bones after 20 years? (IIT 1995)

[**Ans.** 0.61 μg]

[**Hint:** $t = \frac{2.303}{\lambda} \log \frac{\text{initial}}{\text{remaining}}$

$$\lambda = \frac{2.303}{28.1} y^{-1} t = 20 y, \text{ initial} = 1 \mu g, \text{remaining} = x \mu g$$

$$20 = \frac{2.303}{0.693} \times 28.1 \log \frac{1}{x}$$

$$x = 0.61 \mu g]$$

49. It has been estimated that the carbon-14 in the atmosphere is responsible for producing 60 atoms of nitrogen-14 and 60 electrons every hour for each gram of carbon. We can quote this disintegration rate as 60 counts $hour^{-1}$ g^{-1}. A sample of sea shell found near a sea shore was found to have a count of 4 counts $hour^{-1}$ g^{-1}. Estimate the age of the shell. ($t_{1/2}$ for ^{14}C = 5730 years).

[**Ans.** 21000 years (approximately)]

50. Upon irradiating californium with neutrons, a scientist discovered a new nuclide having mass number of 250 and half life of 0.5 hours. Three hours after the irradiation, the observed radioactivity due to the nuclide was 10 dis/min. How many atoms of the nuclide were prepared initially?

[**Ans.** 2.8×10^4]

OBJECTIVE QUESTIONS

Set-1: Questions with single correct answer

1. Natural radioactivity was discovered by:
(a) Rutherford (b) Becquerel
(c) Curie (d) Schmidt
2. Radioactivity is due to:
(a) stable electronic configuration
(b) unstable electronic configuration
(c) stable nucleus
(d) unstable nucleus
3. Radioactivity is essentially:
(a) a chemical activity (b) a physical property
(c) a nuclear property (d) a property of non-metals
4. Radioactivity is generally found in:
(a) light nuclei (b) stable nuclei
(c) heavy nuclei (d) nuclei of intermediate mass
5. The activity of radioisotope changes with:
(a) temperature (b) pressure
(c) chemical environment (d) none of these
6. The rays are given off by a radioactive element from:
(a) nucleus (b) valence electrons
(c) all the orbits (d) outer orbit
7. The alpha particles are:
(a) high energy electrons
(b) positively charged hydrogen ions
(c) high energy X-ray radiations
(d) double positively charged helium nuclei
8. The emission of beta particles is from: (CBSE 1999)
(a) the valence shell of an atom
(b) the inner shell of an atom
(c) the nucleus due to the nuclear conversion proton $\rightarrow$ neutron + electron
(d) the nucleus due to the nuclear conversion neutron $\rightarrow$ proton + electron
9. Identify the nuclear reaction that differs from the rest:
(a) Positron emission (b) K-capture
(c) β-decay (d) α-decay
(e) γ-decay [PET(Kerala) 2008]
[**Hint :** Only γ-emission does not change the n/p (Neutron/Proton, ratio) of the parent element.]
10. Gamma rays are: (MLNR 1990)
(a) high energy electrons
(b) low energy electrons
(c) high energy electromagnetic waves
(d) high energy positrons
11. Radium is a radioactive substance. It dissolves in dilute H_2SO_4 and forms a compound radium sulphate. The compound is:
(a) no longer radioactive
(b) half as radioactive as the radium content
(c) as radioactive as the radium content
(d) twice as radioactive as the radium content
12. The velocity of α-rays is approximately:
(a) equal to that of the velocity of light
(b) $\frac{1}{10}$th of the velocity of light
(c) 10 times more than the velocity of light
(d) uncomparable to the velocity of light
13. α-rays have ionisation power because they possess:
(a) lesser kinetic energy
(b) higher kinetic energy
(c) lesser penetration power
(d) higher penetration power
14. The radiations from a naturally occurring radioactive substance as seen after deflection by a magnetic field in one direction are:
(a) definitely α-rays (b) definitely β-rays
(c) both α and β-rays (d) either α or β-rays
15. Which of the following statements about radioactivity is wrong?
(a) It involves outer electrons activity
(b) It is not affected by temperature or pressure
(c) It is an exothermic process
(d) The radioactivity of an element is not affected by any other element compounded by it
16. The radioactivity of uranium minerals is usually more in comparison to pure uranium. This is due to presence of . . . in the mineral.
(a) actinium (b) thorium
(c) radium (d) plutonium
17. Radioactive disintegration differs from a chemical change in being: (MLNR 1991)
(a) an exothermic change
(b) a spontaneous process
(c) a nuclear process
(d) an unimolecular first order reaction
18. The ionising power of α, β and γ-rays is in the decreasing order:
(a) $\alpha > \beta > \gamma$ (b) $\beta > \alpha > \gamma$
(c) $\gamma > \alpha > \beta$ (d) $\beta > \gamma > \alpha$
19. Which of the following radiations have least effect on both the photographic plate and zinc sulphide screen?
(a) α-rays (b) β-rays
(c) γ-rays (d) All have equal effect
20. γ-rays are emitted from a nucleus due to:
(a) high n/p ratio
(b) excess energy possessed by nucleus after emission of α or β-particles
(c) fission reaction
(d) fusion reaction
21. If a radioactive substance is placed in vacuum at $100^\circ C$, its rate of disintegration in comparison to one atmospheric pressure:

(a) is not affected
(b) increases
(c) decreases
(d) increases when the product is gas

22. In α-decay, n/p ratio:
(a) may increase or decrease
(b) remains constant
(c) decreases
(d) increases

23. In β-decay, n/p ratio:
(a) remains unchanged (b) may increase or decrease
(c) increases (d) decreases

24. A device used for the measurement of radioactivity is:
(a) mass spectrometer (b) cyclotron
(c) nuclear reactor (d) G.M. counter

25. Which of the following does not contain material particles? [CET (Pb.) 1991]
(a) α-rays (b) β-rays (c) γ-rays (d) Anode rays

26. If by mistake some radioactive substance gets into human body, then from the point of view of radiation damage, the most harmful will be one that emits:
(a) γ-rays (b) neutrons (c) β-rays (d) α-rays

27. Radioactive decay is a reaction of:
(a) zero order (b) first order
(c) second order (d) third order

28. If n/p ratio is high, the nucleus tends to stabilise by:
(a) the emission of a β-particle
(b) neutron capture
(c) losing a positron
(d) any one of the above

29. Emission of β-particles by an atom of an element results in the formations of:
(a) isobar (b) isomer (c) isotope (d) isotone

30. Which of the following process will cause the emission of X-ray?
(a) α-emission (b) β-emission
(c) K-electron capture (d) γ-emission

31. When a β-particle is emitted by the atom of a radioactive element, the new species formed possesses: [PET (MP) 1990]
(a) same atomic mass and atomic number less by one unit
(b) same atomic mass and atomic number less by two units
(c) same atomic mass and atomic number higher by one unit
(d) same atomic mass and atomic number higher by two units

32. Successive emission of an α-particle and two β-particles by an atom of a radioactive element results in the formation of its: [IIT (Screening) 1993]
(a) isobar (b) isomer (c) isotone (d) isotope

33. The isotope $^{235}_{92}U$ decays in a number of steps to an isotope of $^{207}_{82}Pb$. The groups of particles emitted in this process will be:
(a) $4\alpha, 7\beta$ (b) $6\alpha, 4\beta$ (c) $7\alpha, 4\beta$ (d) $10\alpha, 8\beta$

34. The number of α and β-particles emitted in the nuclear reaction $^{228}_{90}Th \longrightarrow {}^{212}_{83}Bi$ are: [MLNR 1992; JEE (Orissa) 2010]
(a) $8\alpha, 1\beta$ (b) $4\alpha, 7\beta$ (c) $3\alpha, 7\beta$ (d) $4\alpha, 1\beta$

35. $^{210}_{84}Po \longrightarrow {}^{206}_{82}Pb + {}^{4}_{2}He$
In above reaction, predict the position of Po in the periodic table when lead belongs to IVB group:
(a) IIA (b) VIB (c) IVB (d) VB

36. When $^{226}_{88}Ra$ emits an α-particle, the new element formed belongs to:
(a) third group (b) zero group
(c) fourth group (d) second group

37. The radius of nucleus is: (VITEEE 2007)
(a) proportional to its mass number
(b) inversely proportional to its mass number
(c) proportional to the cube root of its mass number
(d) not related to its mass number

38. The last product of $4n$ series is:
(a) $^{208}_{82}Pb$ (b) $^{207}_{82}Pb$ (c) $^{209}_{82}Pb$ (d) $^{210}_{83}Bi$

39. $4n+2$ series is known as:
(a) actinium series (b) thorium series
(c) uranium series (d) neptunium series

40. A radioactive element A on disintegration gives two elements B and C. If B is helium and C is the element of atomic number 90 and atomic mass 234, the element A is:
(a) $^{238}_{92}U$ (b) $^{234}_{88}Ra$ (c) $^{234}_{90}Th$ (d) $^{234}_{91}Pa$

41. Group displacement law was given by:
(a) Becquerel (b) Rutherford
(c) Mendeleeff (d) Soddy and Fajan

42. ^{234}U has 92 protons and 234 nucleons total in its nucleus. It decays by emitting an alpha particle. After the decay it becomes: (VITEEE 2008; DUMET 2010)
(a) ^{232}U (b) ^{232}Pa (c) ^{230}Th (d) ^{230}Ra

43. Starting from radium, the radioactive disintegration process terminates when the following is obtained:
(a) lead (b) radon (c) radium A (d) radium B

44. The only, most stable nucleus formed by bombarding either $^{27}_{13}Al$ by neutrons or $^{23}_{11}Na$ by deutrons is: [CET (J&K) 2007]
(a) $^{30}_{15}P$ (b) $^{30}_{14}Si$ (c) $^{24}_{12}Mg$ (d) $^{137}_{56}Ba$

45. Quantity of radioactive material which undergoes 10^6 disintegrations per second is called:
(a) Becquerel (b) Rutherford
(c) Curie (d) Faraday

46. The number of α-particles emitted per second by 1 g of ^{226}Ra is 3.7×10^{10}. The decay constant is:
(a) 1.39×10^{-11} sec^{-1} (b) 13.9×10^{-11} sec^{-1}
(c) 139×10^{-10} sec^{-1} (d) 13.9×10^{-10} sec^{-1}

[Hint: $\dfrac{\text{No. of atoms disintegrating per second}}{\text{Total number of atoms present}} = \lambda$

or $\dfrac{3.7 \times 10^{10}}{\dfrac{6.02 \times 10^{23}}{226}} = \dfrac{226 \times 3.7 \times 10^{10}}{6.02 \times 10^{23}} = \lambda$]

47. The decay constant of ^{226}Ra is 1.37×10^{-11} sec^{-1}. A sample of ^{226}Ra having an activity of 1.5 millicurie will contain ...... atoms:
(a) 4.05×10^{18} (b) 3.7×10^{17} (c) 2.05×10^{15} (d) 4.7×10^{10}

[Hint: 1 millicurie $= 3.7 \times 10^7$ disintegrations per sec

1.5 millicurie $= 5.55 \times 10^7$ disintegrations per sec

$$\frac{5.55 \times 10^7}{N_0} = \lambda = 1.37 \times 10^{-11}]$$

48. One curie of activity is equivalent to:
(a) 3.7×10^{17} disintegrations per sec
(b) 3.7×10^{10} disintegrations per sec
(c) 3.7×10^{14} disintegrations per sec
(d) 3.7×10^{3} disintegrations per sec

49. A sample of $^{40}_{19}K$ contains invariably $^{40}_{18}Ar$. This is because $^{40}_{19}K$ has tendency to undergo: **[JEE (Orissa) 2006]**
(a) α-decay (b) positronium decay
(c) β-decay (d) γ-decay

50. The value of disintegration constant of a radioactive isotope:
(a) decreases with increasing temperature
(b) decreases with increasing pressure
(c) increases with increasing concentration
(d) is independent of temperature, pressure and concentration

51. If the amount of a radioactive substance is increased three times, the number of atoms disintegrating per unit time would:
(a) be double
(b) not be change
(c) be triple
(d) be $\frac{1}{3}$rd of the original number of atoms

52. The half life of a radioactive element depends upon:
(a) the amount of the element
(b) the temperature
(c) the pressure
(d) none of these

53. The decay constant of a radioactive sample is λ. The half life and mean life of the sample are respectively: **(MLNR 1990)**
(a) $1/\lambda, \ln 2/\lambda$ (b) $\ln 2/\lambda, 1/\lambda$
(c) $\lambda \ln 2, 1/\lambda$ (d) $\lambda / Pn_2, 1/\lambda$

54. Average life of a radioactive substance is:
(a) 0.44 times of half life (b) 2.44 times of half life
(c) 1.44 times of half life (d) 0.693 times of half life

55. Radium has atomic mass 226 and half life of 1600 years. The number of disintegrations per second per gram are: **(BHU 1990)**
(a) 4.8×10^{10} (b) 3.7×10^{8}
(c) 9.2×10^{6} (d) 3.7×10^{10}

[Hint: $$\frac{\text{No. of disintegrations per sec}}{\text{Total no. of atoms in one gram of Ra}} = \lambda$$
$$= \frac{0.693}{1600 \times 365 \times 24 \times 60 \times 60}$$
or No. of disintegrations per sec
$$= \frac{0.693 \times 6.023 \times 10^{23}}{1600 \times 365 \times 24 \times 60 \times 60 \times 226}]$$

56. A radioactive sample has a half life 1500 years. A sealed tube containing 1 g of the sample will contain after 3000 years: **(MLNR 1994)**
(a) 1 g of the sample (b) 0.5 g of the sample
(c) 0.25 g of the sample (d) 0.01 g of the sample

57. ^{14}C has a half life of 5760 years. 100 mg of the sample containing ^{14}C is reduced to 25 mg in: **[PET (Raj.) 2006]**
(a) 11520 years (b) 2880 years
(c) 1440 years (d) 17280 years

58. If 3/4 quantity of radioactive substance disintegrates in 2 hours, its half life period will be: **(BHU 2006)**
(a) 15 minutes (b) 30 minutes
(c) 60 minutes (d) 90 minutes

59. Initial mass of a radioactive element is 40 g. How many grams of it would be left after 24 years if its half life period is of 8 years?
(a) 2 (b) 5 (c) 10 (d) 20

60. Half life of radium is 1580 years. It remains 1/16 after the years.......: **(VMMC 2007)**
(a) 1580 yrs (b) 3160 yrs (c) 4740 yrs (d) 6320 yrs

61. If half life period of radium is 1600 years, its average life period will be:
(a) 2304 years (b) 4608 years
(c) 230.4 years (d) 23040 years

62. A radioactive isotope having a half life of 3 days was received after 12 days. It was found that there were 3 g of the isotope in the container. The initial mass of the isotope when packed was:
(a) 48 g (b) 36 g (c) 24 g (d) 12 g

63. Radioactivity of a radioactive element remains 1/10 of the original radioactivity after 2.303 seconds. The half life period is:
(a) 2.303 (b) 0.2303 (c) 0.693 (d) 0.0693

[Hint: $$\lambda = \frac{2.303}{t} \log_{10} \frac{a}{(a-x)}$$
or $$\lambda = \frac{2.303}{2.303} \log_{10} \frac{1}{1/10} = 1, T = \frac{0.693}{\lambda} = 0.693]$$

64. A freshly prepared radioactive source of half life period 2 hours emits radiations of intensity which is 64 times the permissible safe level. The minimum time after which it would be possible to work with this source is:
(a) 6 hours (b) 12 hours
(c) 24 hours (d) 48 hours

65. A radio isotope has a half life of 10 days. If today there is 125 g of it left, what was its mass 40 days earlier? **(EAMCET 1991)**
(a) 600 g (b) 1000 g (c) 1250 g (d) 2000 g

66. The half life period of four isotopes is given below:
(i) 7.6 years (ii) 4000 years
(iii) 6000 years (iv) 3.2×10^5 years
Which of the above isotopes is most stable?
(a) (iv) (b) (iii) (c) (ii) (d) (i)

67. The first indication that a stable nucleus can be broken down was afforded by:
(a) Rutherford (b) Madam Curie
(c) Soddy (d) Schmidt

68. The first stable isotope which was transmuted by artificial means was:
(a) $^{16}_{8}O$ (b) $^{14}_{7}N$ (c) $^{12}_{6}C$ (d) $^{9}_{4}Be$

69. The instability of a nucleus is due to: (AIIMS 1999)
(a) high, proton : electron ratio
(b) high, proton : neutron ratio
(c) low, proton : electron ratio
(d) low, proton : neutron ratio

70. When $^{27}_{13}Al$ is bombarded with α-particles, a radioactive isotope of phosphorus $^{30}_{15}P$ with the emission of . . . is formed. [CET (Gujarat) 2006]
(a) neutrons (b) protons
(c) positrons (d) electrons

71. Nuclear reaction accompanied with emission of neutron(s) is: [PMT (MP) 1991]
(a) $^{27}_{13}Al + ^{4}_{2}He \longrightarrow ^{30}_{15}P + ^{1}_{0}n$
(b) $^{12}_{6}C + ^{1}_{1}H \longrightarrow ^{13}_{7}N$
(c) $^{30}_{15}P \longrightarrow ^{30}_{14}Si + ^{0}_{1}e$
(d) $^{240}_{96}Am + ^{4}_{2}He \longrightarrow ^{244}_{97}Bk + ^{0}_{1}e$

72. Which of the following transformations is not correct?
(a) $^{75}_{33}As + ^{4}_{2}He \longrightarrow ^{78}_{35}Br + ^{1}_{0}n$
(b) $^{7}_{3}Li + ^{1}_{1}H \longrightarrow ^{7}_{4}Be + ^{1}_{0}n$
(c) $^{45}_{21}Sc + ^{1}_{0}n \longrightarrow ^{45}_{20}Ca + ^{1}_{0}n$
(d) $^{209}_{83}Bi + ^{2}_{1}H \longrightarrow ^{210}_{84}Po + ^{1}_{0}n$

73. The reaction, $^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{140}_{56}Ba + ^{93}_{36}Kr + 3\,^{1}_{0}n$ represents:
(a) artificial radioactivity
(b) nuclear fission
(c) nuclear fusion
(d) none of these

74. $^{14}_{6}C$ in upper atmosphere is generated by the nuclear reaction: [PET (MP) 1993]
(a) $^{14}_{7}N + ^{1}_{1}H \longrightarrow ^{14}_{6}C + ^{0}_{+1}e + ^{1}_{1}H$
(b) $^{14}_{7}N \longrightarrow ^{14}_{6}C + ^{0}_{+1}e$
(c) $^{14}_{7}N + ^{1}_{0}n \longrightarrow ^{14}_{6}C + ^{1}_{1}H$
(d) $^{14}_{7}N + ^{1}_{1}H \longrightarrow ^{11}_{6}C + ^{4}_{2}He$

75. In the transformation of $^{238}_{92}U$ to $^{234}_{92}U$, if one emission is an α-particle, what should be the other emission(s)? (AIEEE 2006)
(a) two β^- (b) two β^- and one β^+
(c) one β^- and one γ (d) one β^+ and one β^-
[Hint: $^{238}_{92}U \longrightarrow ^{234}_{92}U + ^{4}_{2}He + 2\,^{0}_{-1}e$]

76. The reaction, $^{2}_{1}H + ^{2}_{1}H \longrightarrow ^{3}_{2}He + ^{1}_{0}n$ is called: (CPMT 1990)
(a) fusion (b) fission
(c) endothermic reaction (d) spontaneous reaction

77. When the nucleus of uranium is bombarded with neutrons, it breaks up into two nuclei of nearly equal mass. This process is called:
(a) nuclear fission (b) nuclear fusion
(c) physical change (d) artificial radioactivity

78. Which one of the following is an artificial fuel for nuclear reactors?
(a) ^{238}U (b) ^{239}Pu (c) ^{235}U (d) ^{232}Th

79. A positron is emitted from $^{23}_{11}Na$. The ratio of the atomic mass and atomic number of the resulting nuclide is: (IIT 2007)
(a) 22/10 (b) 22/11 (c) 23/10 (d) 23/12
[Hint: $^{1}_{1}H \longrightarrow ^{1}_{0}n + \underset{\text{Positron}}{^{0}_{+1}e}$
On positron emission, proton is converted to neutron, therefore, atomic number decreases by one unit but atomic mass remains constant.
$\therefore \frac{n}{p}$ ratio $= \frac{23}{10}$]

80. Hydrogen bomb is based on the principle of: (AIEEE 2005)
(a) nuclear fission (b) natural radioactivity
(c) nuclear fusion (d) artificial radioactivity

81. In nuclear reactors, the speed of neutrons is slowed down by:
(a) heavy water (b) ordinary water
(c) zinc rods (d) molten caustic soda

82. Which of the following is not a fissile material?
(a) ^{235}U (b) ^{238}U
(c) ^{233}U (d) ^{239}Pu

83. Which one of the following statements is wrong?
(a) An atom bomb is based on nuclear fission
(b) In atomic reactor, the chain reaction is carried out under control
(c) Fission reactions are the sources of sun's energy
(d) Hydrogen bomb is always associated with atomic bomb

84. The fuel in atomic pile is:
(a) carbon (b) sodium
(c) petroleum (d) uranium

85. Large energy released in atomic bomb explosion is mainly due to:
(a) conversion of heavier to lighter atoms
(b) products having lesser mass than initial substance
(c) release of neutrons
(d) release of electrons

86. One gram of mass is equal to:
(a) 5×10^{10} erg (b) 9×10^{20} erg
(c) 7×10^{5} erg (d) 11×10^{12} erg

87. If the energy released by burning 1 g of carbon is 3×10^{11} erg, then the amount of energy released by converting 1 g of carbon completely to nuclear energy would be equivalent to energy produced by burning ...... g of carbon.
(a) 10^{6} (b) 10^{8}
(c) 9×10^{20} (d) 3×10^{10}

88. Liquid sodium is used in nuclear reactors. Its function is:
(a) to collect the reaction products
(b) to act as heat exchanger
(c) to absorb the neutrons in order to control the chain reaction
(d) to act as moderator to slow down the neutrons

89. A sample of rock from moon contains equal number of atoms of uranium and lead ($t_{1/2}$ for U = 4.5×10^{9} years). The age of the rock would be: [UGET Manipal (Medical) 2006]
(a) 9.0×10^{9} years (b) 4.5×10^{9} years
(c) 13.5×10^{9} years (d) 2.25×10^{9} years

[Hint: $t = \frac{2.303}{\lambda} \log_{10} \left[1 + \frac{\text{No. of Pb atoms}}{\text{No. of U atoms}}\right]$

$= \frac{2.303 \times 4.5 \times 10^9}{0.693} \log_{10} (1 + 1)]$

90. In treatment of cancer, which of the following is used?
(a) $^{131}_{53}I$ (b) $^{32}_{15}P$ (c) $^{60}_{27}Co$ (d) $^{2}_{1}H$

91. In nuclear reactor, chain reaction is controlled by introducing: (AIIMS 1991)
(a) cadmium rod (b) iron rod
(c) platinum rod (d) graphite rod

92. Wooden artifact and freshly cut tree are 7.6 and 15.2 $min^{-1} g^{-1}$ of carbon ($t_{1/2}$ = 5760 years) respectively. The age of the artifact is:
(a) 5760 years (b) $5760 \times \frac{15.2}{7.6}$ years
(c) $5760 \times \frac{7.6}{15.2}$ years (d) $5760 \times (15.2 - 7.6)$ years

[Hint: $t = \frac{2.303}{\lambda} \cdot \log \frac{15.2}{7.6}$ or $t = \frac{2.303 \times 5760}{0.693} \log 2$]

93. The isotope used for dating archaeological finding is:
(a) $^{1}_{1}H$ (b) $^{18}_{8}O$ (c) $^{14}_{6}C$ (d) $^{235}_{92}U$

94. Which one of the following statements is wrong?
(a) Neutron was discovered by Chadwick
(b) Nuclear fission was discovered by Hahn and Strassmann
(c) Polonium was discovered by Madam Curie
(d) Nuclear fusion was discovered by Fermi

95. Neutrons are more effective projectiles than protons because they:
(a) are attracted by nuclei (b) are not repelled by nuclei
(c) travel with high speed (d) none of these

96. The source of enormous energy of sun is:
(a) fusion of hydrogen to form helium
(b) fission of uranium
(c) fusion of deuterium and tritium
(d) fusion of tritium to form helium

97. In the neutron-induced fission of $^{235}_{92}U$; one of the products is $^{90}_{37}Rb$. In this mode, another nuclide and two neutrons are also produced. The other nuclide is: [PMT (HP) 2006]
(a) $^{144}_{54}Xe$ (b) $^{144}_{55}Co$ (c) $^{145}_{55}Co$ (d) $^{143}_{54}Xe$

98. $^{228}_{88}X - 3\alpha - \beta \longrightarrow Y$. The element Y is: [JEE (Orissa) 2008]
(a) $^{216}_{82}Pb$ (b) $^{217}_{82}Pb$
(c) $^{218}_{83}Bi$ (d) $^{216}_{83}Bi$

99. Which radioactive isotope is used to detect tumours?
(a) ^{74}As (b) ^{24}Na (c) ^{131}I (d) ^{60}Co

100. Natural uranium consists of ^{235}U:
(a) 99% (b) 50%
(c) 10% (d) 0.7%

101. In the nuclear reaction, $^{14}_{7}N + ^{4}_{2}He \longrightarrow ^{p}_{q}X + ^{1}_{1}H$ the nucleus X is: (MLNR 1995)
(a) nitrogen of mass 16 (b) nitrogen of mass 17
(c) oxygen of mass 16 (d) oxygen of mass 17

102. The radioactive decay of $^{88}_{35}X$ by a beta emission produces an unstable nucleus which spontaneously emits a neutron. The final product is: (MLNR 1995)
(a) $^{88}_{37}X$ (b) $^{89}_{35}Y$
(c) $^{88}_{34}Z$ (d) $^{87}_{36}W$

103. $^{27}_{13}Al$ is a stable isotope. $^{29}_{13}Al$ is expected to disintegrate by: (IIT 1996)
(a) α-emission (b) β-emission
(c) positron emission (d) proton emission

104. The mass defect of the nuclear reaction $^{8}_{5}B \rightarrow ^{8}_{4}Be + ^{0}_{1}e$ is: (JIPMER 1999)
(a) Δm = atomic mass of ($^{8}_{4}Be - ^{8}_{5}B$)
(b) Δm = atomic mass of ($^{8}_{4}Be - ^{8}_{5}B$) + mass of one electron
(c) Δm = atomic mass of ($^{8}_{4}Be - ^{8}_{5}B$) + mass of the positron
(d) Δm = atomic mass of ($^{8}_{4}Be - ^{8}_{5}B$) + mass of two electrons

105. Which of the following is the man-made radioactive disintegration series?
(a) Thorium series (b) Neptunium series
(c) Uranium series (d) Actinium series

106. The density of nucleus is of the order of:
(a) 10^5 kg m^{-3} (b) 10^{10} kg m^{-3}
(c) 10^{17} kg m^{-3} (d) 10^{25} kg m^{-3}

107. A radioactive isotope having a half life of 3 days was received after 12 days. It was found that there were 3 g of the isotope in the container. The initial weight of the isotope when packed was:
(a) 12 g (b) 24 g
(c) 36 g (d) 48 g

108. A radioactive substance is decaying with $t_{1/2}$ = 30 days. On being separated into two fractions, one of the fractions, immediately after separation, decays with $t_{1/2}$ = 2 days. The other fraction, immediately after separation would show:
(a) constant activity (b) increasing activity
(c) decay with $t_{1/2}$ = 30 days (d) decay with $t_{1/2}$ = 28 days

109. A radioactive substance has a constant activity of 2000 disintegrations per minute. The material is separated into two fractions, one of which has an initial activity of 1000 disintegrations per second while the other fraction decays with $t_{1/2}$ = 24 hours. To the total activity in both samples after 48 hours of separation is:
(a) 1500 (b) 1000 (c) 1250 (d) 2000

110. A radioactive element X has an atomic number of 100. It decays directly into an element Y which decays directly into the element Z. In both processes a charged particle is emitted. Which of the following statements would be true?
(a) Y has an atomic number of 102
(b) Y has an atomic number of 101
(c) Z has an atomic number of 100
(d) Z has an atomic number of 99

111. Three isotopes of an element have mass numbers M, $(M + 1)$ and $(M + 2)$. If the mean mass number is $(M + - 0.5)$, then which of the following ratios may be accepted for M, $(M + 1)$, $(M + 2)$ in that order?
(a) 1 : 1 : 1 (b) 4 : 1 : 1 (c) 3 : 2 : 1 (d) 2 : 1 : 1

112. Enrichment of uranium is made by:
(a) distillation (b) diffusion
(c) evaporation (d) bleaching

113. Let us consider emission of α-particle from uranium nucleus:

$$_{92}U^{235} - {}_2He^4 \rightarrow {}_{90}Th^{231}$$

$e = 92$	$e = 0$	$e = 90$
$p = 92$	$p = 2$	$p = 90$
$n = 143$	$n = 2$	$n = 141$

Shortage of two electrons in thorium is due to:
(a) conversion of electron to positron
(b) combination with positron to evolve energy
(c) annihilation
(d) absorption in the nucleus

114. Artificial radioactive elements are present in:
(a) s-block (b) p-block (c) d-block (d) f-block

115. Half life of $_6C^{14}$, if its λ is 2.13×10^{-4} yrs, is: (CBSE 1999)
(a) 3.5×10^4 years (b) 3×10^3 years
(c) 2×10^2 years (d) 4×10^3 years

116. The ^{60}Co isotope decays with a half life of 5.3 years. How long would it take for 7/8 of a sample of 500 mg of ^{60}Co to disintegrate? (SCRA 2007)
(a) 21.2 years (b) 15.9 years
(c) 10.6 years (d) 5.3 years

117. Isotope of uranium used in atomic bomb is : [PET (MP) 2008]
(a) $^{237}_{92}U$ (b) $^{238}_{92}U$ (c) $^{239}_{92}U$ (d) $^{235}_{92}U$

118. Which among the following is wrong about isodiapheres?
(a) They have the same difference of neutrons and protons or same isotopic number
(b) Nuclide and its decay product after α-emission are isodiapheres
(c) $_Z A^M \longrightarrow {}_{Z-2}B^{M-4} + {}_2He^4$
'A' and 'B' are isodiapheres
(d) All are correct

119. At radioactive equilibrium, the ratio of two atoms A and B are $3.1 \times 10^9 : 1$. If half life of 'A' is 2×10^{10} yrs, what is half life of 'B'?
(a) 6.45 yrs (b) 4.65 yrs (c) 5.46 yrs (d) 5.64 yrs

120. The decay constant for an α-decay of Th^{232} is 1.58×10^{-10} s^{-1}. How many α-decays occur from 1 g sample in 365 days?
(a) 2.89×10^{19} (b) 1.298×10^{19}
(c) 8.219×10^{19} (d) None of these

121. What percentage of decay takes place in the average life of a substance?
(a) 63.21% (b) 36.79% (c) 90% (d) 99%

122. SI unit of radioactive decay is: [PMT (HP) 2006]
(a) curie (b) rutherford
(c) becquerel (d) all of these

123. The number of neutrons accompanying the formation of $^{139}_{54}$Xe and $^{94}_{38}$Sr from the absorption of a slow neutron by $^{235}_{92}$U, followed by nuclear fission is: (IIT 1999)
(a) 0 (b) 2 (c) 1 (d) 3

124. Thiosulphate ion ($S_2O_3^{2-}$) on acidification changes to SO_2 along with precipitation of sulphur,

$$^{35}S^{32}SO_3^{2-} + 2H^+ \longrightarrow H_2O + SO_2 + S$$

which is the correct statement?
(a) S^{35} is in sulphur (b) S^{35} is in SO_2
(c) S^{35} is in both (d) S^{35} is in none

125. A radioactive element decays as,

$$X \xrightarrow[t_{1/2} = 30 \text{ min}]{\alpha \text{ decay}} Y \xrightarrow[t_{1/2} = 2 \text{ days}]{(-2\beta)\beta \text{ decay}} Z$$

which of the following statements about this decay process is incorrect?
(a) After two hours, less than 10% of the initial X is left
(b) Maximum amount of Y present at any time before 30 min is less than 50% of the initial amount of X
(c) Atomic number of X and Z are same
(d) The mass number of Y is greater than X

126. Among the following nuclides, the highest tendency to decay by (β^+) emission is:
(a) ^{59}Cu (b) ^{63}Cu (c) ^{67}Cu (d) ^{68}Cu

127. Identify [A] and [B] in the following:
$^{227}_{89}Ac \xrightarrow{-\beta} [A] \xrightarrow{-\alpha} [B] \xrightarrow{-\alpha} Rn$ [JEE (WB) 2010]
(a) Po, Rn (b) Th, Po (c) Ra, Th (d) Th, Ra
[**Hint:** $^{227}_{89}Ac \xrightarrow{-\beta} {}^{227}_{90}Th \xrightarrow{-\alpha} {}^{223}_{88}Ra$]

128. β-particle is emitted in radioactivity by: (AIEEE 2002)
(a) conversion of proton to neutron
(b) from outermost orbit
(c) conversion of neutron to proton
(d) β-particle is not emitted

129. The nuclear reaction,

$$^{63}_{29}Cu + {}^4_2He \longrightarrow {}^{37}_{17}Cl + 14\ {}^1_1H + 16\ {}^1_0n$$

is referred to as: [PET (MP) 2002]
(a) spallation reaction (b) fusion reaction
(c) fission reaction (d) chain reaction

130. ^{226}Ra disintegrates at such a rate that after 3160 yrs only one fourth of its original amount remains. Half life of ^{226}Ra will be: [PET (MP) 2002]
(a) 790 years (b) 3160 years
(c) 1580 years (d) 6230 years

131. $^{235}_{92}$U nucleus absorbs a neutron and disintegrates into $^{139}_{54}$Xe, $^{94}_{38}$Sr and 'x'. What will be the product x? [CBSE (PMT) 2002]
(a) 3-neutrons (b) 2-neutrons
(c) α-particles (d) β-particles

132. A radioisotope, tritium ($^3_1$H) has half life of 12.3 years. If the initial amount of tritium is 32 mg, how many milligrams of it would remain after 49.2 years? [CBSE (PMT) 2003]
(a) 1 mg (b) 2 mg
(c) 4 mg (d) 8 mg

133. The radio nuclide $^{234}_{90}$Th undergoes two successive β-decays followed by one α-decay. The atomic number and mass number of the resulting radio nuclide are: (AIEEE 2003)

(a) 92, 234 (b) 94, 230 (c) 90, 230 (d) 92, 230

134. The half life of a radioactive isotope is three hours. If the initial mass of isotope were 256 g, the mass of it remaining undecayed after 18 hours would be: (AIEEE 2003)

(a) 4 g (b) 8 g (c) 12 g (d) 16 g

135. Consider the following nuclear reactions:

$$^{238}_{92}M \rightarrow ^{X}_{Y}N + 2\,^{4}_{2}He;\quad ^{X}_{Y}N \longrightarrow ^{A}_{B}L + 2\beta^{+}$$

The number of neutrons in the element L is: (AIEEE 2004)

(a) 142 (b) 144 (c) 140 (d) 146

136. A radioactive element gets spilled over the floor of a room. Its half life period is 30 days. If initial rate is ten times the permissible value, after how many days will it be safe to enter the room? (AIEEE 2007)

(a) 100 days (b) 1000 days
(c) 300 days (d) 10 days

137. A photon of hard gamma radiation knocks a proton out of $^{24}_{12}Mg$ nucleus to form: (AIEEE 2005)

(a) the isotope of parent nucleus
(b) the isobar of parent nucleus
(c) the nuclide of $^{23}_{11}Na$
(d) the isobar of $^{23}_{11}Na$

138. The element $^{232}_{90}Th$ belongs to thorium series. Which of the following will act as the end product of the series? [BHU (Pre.) 2005]

(a) $^{208}_{82}Pb$ (b) $^{209}_{82}Bi$ (c) $^{206}_{82}Pb$ (d) $^{207}_{82}Pb$

139. $^{238}_{92}U$ emits 8α-particles and 6β-particles. The neutron/proton ratio in the product nucleus is: (AIIMS 2005)

(a) 60/41 (b) 61/40 (c) 62/41 (d) 61/42

140. Calculate the mass loss in the following:

$$^{2}_{1}H + ^{3}_{1}H \longrightarrow ^{4}_{2}He + ^{1}_{0}n$$

Given the masses: $^{2}_{1}H = 2.014$ amu, $^{3}_{1}H = 3.016$ amu; $^{4}_{2}He = 4.004$ amu, $^{1}_{0}n = 1.008$ amu. [PET (Kerala) 2005]

(a) 0.018 amu (b) 0.18 amu
(c) 0.0018 amu (d) 1.8 amu
(e) 18 amu

141. A nuclide of an alkaline earth metal undergoes radioactive decay by emission of the α-particles in succession. The group of the periodic table to which the resulting daughter element would belong is: [CBSE (PMT) 2005]

(a) 4th group (b) 6th group (c) 14th group (d) 16th group

142. In the reaction $^{2}_{1}H + ^{3}_{1}H \longrightarrow ^{4}_{2}He + ^{1}_{0}n$, if the binding energies of $^{2}_{1}H$, $^{3}_{1}H$ and $^{4}_{2}He$ are a, b and c (in MeV) respectively, then energy (in MeV) released in this reaction is: [CBSE-PMT (Physics) 2005]

(a) $a + b - c$ (b) $c + a - b$
(c) $c - a - b$ (d) $a + b + c$

143. Two radioactive elements X and Y have half lives 6 min and 15 min respectively. An experiment starts with 8 times as many atoms of X as Y. How long it takes for the number of atoms of X left to equal the number of atoms of Y left? [PET (Kerala) 2008]

(a) 6 min (b) 12 min (c) 48 min (d) 30 min
(e) 24 min

144. Which of the following has the highest value of radioactivity? (DPMT 2009)

(a) 1 g of Ra (b) 1 g of $RaSO_4$
(c) 1 g of $RaBr_2$ (d) 1 g of $Ra(HPO_4)$

145. An artificial transmutation was carried out on $^{14}_{7}N$ by an α-particle which resulted in an unstable nuclide and a proton. What is the ratio of the atomic mass to the atomic number of the unstable nuclide? (SCRA 2009)

(a) $\frac{17}{8}$ (b) $\frac{15}{7}$ (c) $\frac{17}{9}$ (d) $\frac{15}{8}$

[**Hint:** $^{14}_{7}N + ^{4}_{2}He \longrightarrow ^{17}_{8}O + ^{1}_{1}H$

$$\frac{\text{Mass Number}}{\text{Atomic Number}} = \frac{17}{8}]$$

146. If 0.4 curie be the activity of 1 gram of a radioactive sample whose atomic mass is 226, then what is the half-life period of the sample? (1 curie = 3.7×10^{10} dis sec^{-1}) (SCRA 2009)

(a) 1.2×10^{11} sec (b) 1.8×10^{11} sec
(c) 1.2×10^{10} sec (d) 1.8×10^{10} sec

[**Hint:** Rate of decay $= \frac{0.693}{t_{1/2}} \times \frac{w}{A_w} \times 6.023 \times 10^{23}$

$$0.4 \times 3.7 \times 10^{10} = \frac{0.693}{t_{1/2}} \times \frac{1}{226} \times 6.023 \times 10^{23}$$

$$t_{1/2} = 1.2 \times 10^{11} \text{ sec}]$$

147. The half-life period of uranium is 4.5 billion years. After 9.0 billion years, the number of moles of helium liberated from the following nuclear reaction will be:

$$^{238}_{92}U \longrightarrow ^{234}_{90}Th + ^{4}_{2}He$$

Initially there was 1 mole uranium. [PET (MP) 2010]

(a) 0.75 mol (b) 1.0 mol (c) 11.2 mol (d) 22.4 mol

Set-2: The questions given below may have more than one correct answers

1. Match the following radioactive series:

(A) $4n$ (i) Uranium series
(B) $4n + 1$ (ii) Neptunium series
(C) $4n + 2$ (iii) Actinium series
(D) $4n + 3$ (iv) Thorium series

	A	B	C	D
(a)	(i)	(ii)	(iii)	(iv)
(b)	(iv)	(ii)	(i)	(iii)
(c)	(iii)	(i)	(iv)	(ii)
(d)	(ii)	(iii)	(i)	(iv)

2. Match the following reactions:

(A) $_4Be^9 + _2He^4 \rightarrow _6C^{12} + \ldots$ (i) $_2He^4$
(B) $_6C^{12} + \ldots \rightarrow _5B^{10} + _2He^4$ (ii) $_0n^1$
(C) $_7N^{14} + \ldots \rightarrow _8O^{17} + _1H^1$ (iii) $_1D^2$
(D) $_{20}Ca^{40} + \ldots \rightarrow _{19}K^{37} + _2He^4$ (iv) $_1H^1$

	A	B	C	D
(a)	(i)	(ii)	(iii)	(iv)
(b)	(ii)	(iii)	(i)	(iv)
(c)	(iv)	(ii)	(iii)	(i)
(d)	(iii)	(ii)	(i)	(iv)

3. A radioactive element is present in VIII group of the periodic table. If it emits one α-particle, the new position of the nuclide will be:

(a) VI B (b) VIII
(c) VII B (d) I B

4. Which statement is true about n/p ratio?
(a) It increases by β-emission
(b) It increases by α-emission
(c) It increases by γ-emission
(d) None of the above

5. How many α and β-particles should be eliminated so that an isodiaphere is formed?
(a) $n\alpha, n\beta$ (b) $n\alpha, (n+1)\beta$
(c) $n\alpha$ (d) $n\beta$

6. Match the following:

Series	Particles emitted
(a) Thorium	(i) $8\alpha, 5\beta$
(b) Neptunium	(ii) $8\alpha, 6\beta$
(c) Actinium	(iii) $6\alpha, 4\beta$
(d) Uranium	(iv) $7\alpha, 4\beta$

	A	B	C	D
(a)	(iv)	(iii)	(ii)	(i)
(b)	(ii)	(i)	(iv)	(iii)
(c)	(iii)	(i)	(iv)	(ii)
(d)	(i)	(ii)	(iii)	(iv)

7. Which of the following are used as control rods in a nuclear reactor?
(a) Cadmium rods (b) Graphite rods
(c) Steel rods (d) All of these

8. Which of the following notations shows the product incorrectly?
(a) $_{96}Cm^{242}(\alpha, 2n)\,_{97}Bk^{243}$ (b) $_5B^{10}(\alpha, n)\,_7N^{13}$
(c) $_7N^{14}(n, p)\,_6C^{14}$ (d) $_{14}Si^{28}(d, n)\,_{15}P^{29}$

9. Which is true about decay constant (λ)?
(a) Unit is time^{-1}
(b) Value of λ is always less than 1
(c) λ is independent of temperature
(d) λ is defined as the ratio of no. of atoms disintegrating per unit time to the total no. of atoms present at that time

10. Which of the following is not correct? (EAMCET 2006)
(a) Nuclei of atoms participate in nuclear reactions
(b) $^{40}_{20}Ca$ and $^{40}_{18}Ar$ are isotones
(c) 1 amu of mass defect is approximately equal to 931.5 MeV
(d) Uranium (U^{238}) series is known as $(4n+2)$ series

11. Correct order of radioactivity is:
(a) $_1H^1 > {_1H^2} > {_1H^3}$ (b) $_1H^3 > {_1H^2} > {_1H^1}$
(c) $_1H^3 > {_1He^1} > {_1H^2}$ (d) $_1H^3 > {_1H^1} = {_1H^2}$

12. At radioactive equilibrium, the ratio between 2 atoms of radioactive elements A and B is $3 \times 10^9 : 1$. If $t_{1/2}$ of A is 10^{10} yrs what is $t_{1/2}$ of B?
(a) 30 yrs (b) 3 yrs
(c) 3.3 yrs (d) None of these

13. In the sequence of the following nuclear reaction,

$$X^{238}_{98} \xrightarrow{-\alpha} Y \xrightarrow{-\beta} Z \xrightarrow{-\beta} L \xrightarrow{n\alpha} {}_{90}M^{218}$$

what is the value of n?
(a) 3 (b) 4
(c) 5 (d) 6

14. ^{60}Co has $t_{1/2} = 5.3$ years. The time taken for 7/8 of the original sample to disintegrate will be:
(a) 4.6 yrs (b) 9.2 yrs
(c) 10.6 yrs (d) 15.9 yrs

15. Which of the following is/are correct?
(a) α-rays are more penetrating than β-rays
(b) α-rays have greater ionizing power than β-rays
(c) β-particles are not present in the nucleus, yet they are emitted from the nucleus
(d) γ-rays are not emitted simultaneously with α and β-rays

16. Select the wrong statement:
(a) Nuclear isomers contain the same number of protons and neutrons
(b) The decay constant is independent of the amount of the substance taken
(c) One curie = 3.7×10^{10} dis/minute
(d) Actinium series starts with U^{238}

17. In a nuclear reactor, heavy water is used to:
(a) provide high speed to neutrons
(b) reduce the speed of neutrons
(c) capture neutrons produced by nuclear fission
(d) transfer the heat from the nuclear reactor

18. The correct starting material and product of different disintegration series are:
(a) ^{232}Th, ^{208}Pb (b) ^{235}U, ^{206}Pb
(c) ^{238}U, ^{207}Pb (d) ^{237}Np, ^{209}Bi

19. Which of the following is/are not true?
(a) The most radioactive element present in pitchblende is uranium
(b) ^{32}P is used for the treatment of leukaemia
(c) CO_2 present in the air contains ^{12}C only
(d) Omission of γ-rays changes the mass number but not atomic number

20. Which of the following is/are correct?
(a) 1 Curie = 3.7×10^{10} d/s
(b) 1 Rutherford = 10^6 d/s
(c) 1 Becquerel = 1 d/s
(d) 1 Fermi = 10^3 d/s

21. Match the List-I and List-II and select the correct answer using the codes given below the lists:

List-I Nuclear reactor component	List-II Substance used
1. Moderator	A. Uranium
2. Control rods	B. Graphite
3. Fuel rods	C. Boron
4. Coolant	D. Lead
	E. Sodium

[PET (Kerala) 2005]

Codes:
(a) 1—B, 2—A, 3—C, 4—E
(b) 1—B, 2—C, 3—A, 4—E
(c) 1—C, 2—B, 3—A, 4—E
(d) 1—C, 2—D, 3—A, 4—B
(e) 1—D, 2—C, 3—B, 4—A

22. Match the List-I and List-II and select the correct answer using the codes given below the lists:

List-I Isotope	List-II Characteristics
A. $^{40}_{20}Ca$	1. Unstable, α-emitter
B. $^{133}_{53}I$	2. Unstable, β-emitter
C. $^{121}_{53}I$	3. Unstable, positron emitter
D. $^{232}_{90}Th$	4. Stable

Codes:	A	B	C	D
(a)	1	2	3	4
(b)	1	3	2	4
(c)	4	3	2	1
(d)	4	2	3	1

23. Match the List-I with List-II and select the correct answer using the codes given below the lists:

List-I Isotope	List-II Characteristics
A. ^{32}P	1. Location of tumour in brain
B. ^{24}Na	2. Location of blood clot and circulatory disorders
C. ^{60}Co	3. Radiotherapy
D. ^{131}I	4. Agriculture research

Codes:	A	B	C	D
(a)	4	1	2	3
(b)	4	3	2	1
(c)	4	2	3	1
(d)	3	1	2	4

24. Consider the following nuclear reactions:
1. $^{14}_{7}N + ^{4}_{2}He \longrightarrow ^{17}_{8}O + ^{1}_{1}H$
2. $^{9}_{4}Be + ^{1}_{1}H \longrightarrow ^{9}_{3}Li + ^{4}_{2}He$
3. $^{24}_{12}Mg + ^{4}_{2}He \longrightarrow ^{27}_{14}Si + ^{1}_{0}n$
4. $^{10}_{5}B + ^{4}_{2}He \longrightarrow ^{13}_{7}N + ^{1}_{0}n$

Examples of induced radioactivity would include the reactions:
(a) 3 and 4 (b) 1 and 2 (c) 1, 3 and 4 (d) 1, 2, 3 and 4

25. Match the Column-I Radio-isotope with Column-II Medicinal use and select correct matching:

Column-I	Column-II
(I) ^{60}Co	(a) Leucaemia
(II) ^{131}I	(b) Anaemia
(III) ^{59}Fe	(c) Cancerous tumours
(IV) ^{32}P	(d) Disorders of thyroid gland

(a) I—c; II—d; III—a; IV—b
(b) I—a; II—b; III—c, IV—d
(c) I—c; II—d; III—b; IV—a
(d) I—d; II—c; III—b; IV—a

26.

Column-I	Column-II
(I) $^{14}_{6}C$	(a) Unstable and β-emitter
(II) $^{24}_{11}Na$	(b) Stable
(III) $^{13}_{7}N$	(c) Unstable, positron emitter
(IV) $^{13}_{6}C$	(d) Unstable, α-emitter

Correct matching is/are:
(a) I only (b) III only
(c) II and IV (d) I and III

27. Which of the following statements is/are correct?
1. A nucleus in an excited state may give up its excitation energy and return to the ground state by emission of electromagnetic γ-radiation.
2. γ-radiations are emitted as secondary effect of α and β-emission.
3. The nuclear isomers produced by γ-ray bombardment have the same atomic and mass number but differ in their life-times (whatever their ground state may be).
4. X-ray and γ-ray are both electromagnetic.

(a) 1 and 2 (b) 1, 2 and 3 (c) 2 and 3 (d) 1, 2, 3 and 4

28. Which of the following statements is/are correct?
1. When an electron is emitted by an atom and its nucleus gets de-excited as a result, the process is called internal conversion.
2. Electron capture and positron emission are identical.
3. Neutrons are emitted in the electron capture process.
4. Pair production is a process which involves the creation of positron-electron pair by a photon of energy 1.02 MeV.

(a) 1 and 2 (b) 1, 2 and 4
(c) 2, 3 and 4 (d) All are wrong

29. A nuclide has mass number (A) and atomic number (Z). During a radioactive process if:
1. both A and Z decrease, the process is called α-decay.
2. A remains unchanged and Z decreases by one, the process is called β^+ or positron decay or K-electron capture.
3. both A and Z remain unchanged, the process is called γ-decay.
4. both A and Z increase, the process is called nuclear isomerism.

The correct answer is:
(a) 1, 2 and 3 (b) 2, 3 and 4
(c) 1, 3 and 4 (d) 1, 2 and 4

30. In the decay process:

$$A \xrightarrow{-\alpha} B \xrightarrow{-\alpha} C \xrightarrow{-\beta} D$$

1. A and B are isobars
2. A and D are isotopes
3. C and D are isobars
4. A and C are isotones

The correct answer is:
(a) 1 and 2 (b) 2 and 3
(c) 3 and 4 (d) 1 and 4

31. The nuclide X undergoes α-decay and another nuclide Y, β^--decay. Which of the following statements are correct?
1. The β^--particles emitted by Y may have widely different speeds.
2. The α-particles emitted by X may have widely different speeds.
3. The α-particles emitted by X will have almost same speed.
4. The β-particles emitted by Y will have the same speed.

(a) 1 and 3 are correct (b) 2 and 3 are correct
(c) 3 and 4 are correct (d) 1 and 4 are correct

32. Fill in the blank space with a suitable answer selected from the list below. Write only the letter (A, B, C, ..., etc) of the correct answer in the blanks.

	Answer
(i) $^{12}_{6}C + ^{1}_{1}H \longrightarrow ^{13}_{7}N$	......
(ii) $^{27}_{13}Al + ^{1}_{1}H \longrightarrow ^{24}_{12}Mg + ^{4}_{2}He$	......
(iii) $^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{140}_{56}Ba + ^{93}_{36}Kr + 3\,^{1}_{0}n$	......
(iv) $^{75}_{33}As + ^{2}_{1}H \longrightarrow ^{56}_{25}Mn + 9\,^{1}_{1}H + 12\,^{1}_{0}n$	......
(v) $^{2}_{1}H + ^{3}_{1}H \longrightarrow ^{4}_{2}He + ^{1}_{0}n$	......

Answers:
A: Projectile capture
B: Spallation
C: Fusion
D: Projectile capture and particle emission
E: Fission

Select the correct answers according to the given codes:

Codes:	(i)	(ii)	(iii)	(iv)	(v)
(a)	A	D	E	B	C
(b)	D	C	A	E	B
(c)	A	B	C	D	E
(d)	E	D	C	B	A

Assertion-Reason TYPE QUESTIONS

The questions given below consist of two statements each printed as **Assertion (A)** and **Reason (R).** While answering these questions you are required to choose any one of the following four:

(a) If both (A) and (R) are correct and (R) is the correct explanation for (A).
(b) If both (A) and (R) are correct and (R) is not the correct explanation for (A).
(c) If (A) is correct but (R) is incorrect.
(d) If both (A) and (R) are incorrect.
(e) If (A) is incorrect but (R) is correct.

1. (A) Mass numbers of most of the elements are fractional.
(R) Mass numbers are obtained by comparing with the mass number of carbon taken as 12.

2. (A) The activity of 1 g pure uranium-235 will be greater than the same amount present in U_3O_6.
(R) In the combined state, the activity of the radioactive element decreases.

3. (A) α-rays have greater ionising power than β.
(R) α-particles carry 2^+ charge while β-particles carry only 1^- charge.

4. (A) β-particles have greater penetrating power than α-rays but less than γ-rays.
(R) β-particles are lighter than α but heavier than γ.

5. (A) During β-decay, a new element with atomic number greater than one is obtained.
(R) Protons and neutrons keep on changing into one another through meson.

6. (A) The average life of a radioactive element is infinity.
(R) As a radioactive element disintegrates, more of it is formed in nature by itself.

7. (A) Hydrogen bomb is more powerful than atomic bomb.
(R) In hydrogen bomb, reaction is initiated.

8. (A) The archaeological studies are based on the radioactive decay of carbon-14 isotope.
(R) The ratio of C-14 to C-12 in the animals or plants is the same as that in the atmosphere.

9. (A) The reactions taking place in the sun are nuclear fusion reactions.
(R) The main reason for nuclear fusion reactions in the sun is that H_2 is present in the sun's atmosphere so that hydrogen nuclei can fuse to form helium.

10. (A) In a radioactive disintegration, an electron is emitted by the nucleus.
(R) Electrons are always present inside the nucleus.

11. (A) In radioactive disintegrations, $_2He^4$ nuclei can come out of the nucleus but lighter $_2He^3$ can't.
(R) Binding energy of $_2He^3$ is more than that of $_2He^4$.

12. (A) Protons are better projectiles than neutrons.
(R) The neutrons being neutral do not experience repulsion from positively charged nucleus.

13. (A) Enrichment of U^{235} from a mixture containing more abundant U^{238} is based on diffusion of UF_6.
(R) UF_6 is a gaseous compound under ordinary conditions.

14. (A) The nucleus emits β-particles though it doesn't contain any electron in it.
(R) The nucleus shows the transformation $_0n^1 \rightarrow p + \beta +$ anti-neutrino for β-emission.

15. (A) Any kind of exchange force helps the nucleus to be more destabilised.
(R) π-mesons are exchanged between nucleons incessantly.

16. (A) Nuclide $_{13}Al^{30}$ is less stable than $_{20}Ca^{40}$. (IIT 1998)
(R) Nuclides having odd number of protons and neutrons are generally unstable.

17. (A) During β-decay, a new element with atomic number greater than one is obtained.
(R) Protons and neutrons keep on changing into one another with the help of meson.

18. (A) The position of an element in periodic table after emission of one α and two β-particles remains unchanged.
(R) Emission of one α and two β-particles gives isotope of the parent element which acquires same position in the periodic table.

19. (A) Nuclear isomers have same atomic number and same mass number but with different radioactive properties.
(R) $U_{(A)}$ and $U_{(Z)}$ are nuclear isomers.

20. (A) The emission of α-particles results in the formation of isodiapheres of parent element.
(R) Isodiapheres have same isotopic number.

21. (A) $^{238}_{92}U\,(IIIB) \xrightarrow{-\alpha} A \xrightarrow{-\alpha} B \xrightarrow{-\beta} C$
(R) Element B will be of IIA group.

22. (A) β-particles are deflected more than α-particles in a given electric field.
(R) Charge on α-particles is larger than on β-particles.

23. (A) The nucleus of gold is stable even though there is a very strong coulombic repulsion among the protons.
(R) The inverse square coulomb force is exactly balanced by another inverse square force which is very powerful, *i.e.*, nuclear force.

24. (A) K-shell electron capture is detected by analysing the wavelength of X-ray emitted.
(R) The wavelength of the X-ray is characteristic of the daughter element and not the parent element.

25. (A) Half life of a radioactive isotope is the time required to decrease its mass number by half.
(R) Half life of radioactive isotopes is independent of initial amount of the isotope.

26. (A) In a nuclear fission process, the total mass of fragments is always greater than the mass of the original nucleus.
(R) Difference in the mass due to the fission of a heavy nucleus is converted into energy according to mass-energy conversion. (SCRA 2007)

Answers : OBJECTIVE QUESTIONS

• Set–1

1. (b)	**2.** (d)	**3.** (c)	**4.** (c)	**5.** (d)	**6.** (a)	**7.** (d)	**8.** (d)
9. (e)	**10.** (c)	**11.** (c)	**12.** (b)	**13.** (b)	**14.** (d)	**15.** (a)	**16.** (c)
17. (c)	**18.** (a)	**19.** (c)	**20.** (b)	**21.** (a)	**22.** (d)	**23.** (d)	**24.** (d)
25. (c)	**26.** (a)	**27.** (b)	**28.** (a)	**29.** (a)	**30.** (c)	**31.** (c)	**32.** (d)
33. (c)	**34.** (d)	**35.** (b)	**36.** (b)	**37.** (c)	**38.** (a)	**39.** (c)	**40.** (a)
41. (d)	**42.** (c)	**43.** (a)	**44.** (d)	**45.** (b)	**46.** (a)	**47.** (a)	**48.** (b)
49. (b)	**50.** (d)	**51.** (c)	**52.** (d)	**53.** (b)	**54.** (c)	**55.** (d)	**56.** (c)
57. (a)	**58.** (c)	**59.** (b)	**60.** (d)	**61.** (a)	**62.** (a)	**63.** (c)	**64.** (b)
65. (d)	**66.** (a)	**67.** (a)	**68.** (b)	**69.** (b)	**70.** (a)	**71.** (a)	**72.** (c)
73. (b)	**74.** (c)	**75.** (a)	**76.** (a)	**77.** (a)	**78.** (b)	**79.** (c)	**80.** (d)
81. (a)	**82.** (b)	**83.** (c)	**84.** (d)	**85.** (b)	**86.** (b)	**87.** (d)	**88.** (b)
89. (b)	**90.** (c)	**91.** (a)	**92.** (a)	**93.** (c)	**94.** (d)	**95.** (b)	**96.** (a)
97. (b)	**98.** (b)	**99.** (a)	**100.** (d)	**101.** (d)	**102.** (d)	**103.** (b)	**104.** (d)
105. (b)	**106.** (c)	**107.** (d)	**108.** (b)	**109.** (d)	**110.** (b,d)	**111.** (b)	**112.** (b)
113. (b,c)	**114.** (d)	**115.** (c)	**116.** (b)	**117.** (d)	**118.** (d)	**119.** (a)	**120.** (b)
121. (a)	**122.** (c)	**123.** (d)	**124.** (a)	**125.** (d)	**126.** (c)	**127.** (d)	**128.** (c)
129. (a)	**130.** (c)	**131.** (b)	**132.** (b)	**133.** (c)	**134.** (a)	**135.** (b)	**136.** (a)
137. (c)	**138.** (a)	**139.** (c)	**140.** (a)	**141.** (c)	**142.** (c)	**143.** (d)	**144.** (a)
145. (a)	**146.** (a)	**147.** (a)					

• Set–2

1. (b)	**2.** (b)	**3.** (a, b, c)	**4.** (b)	**5.** (c)	**6.** (c)	**7.** (a)	**8.** (a)
9. (c)	**10.** (b)	**11.** (b)	**12.** (c)	**13.** (b)	**14.** (d)	**15.** (b, c, d)	**16.** (c, d)
17. (b, d)	**18.** (a, d)	**19.** (a, d)	**20.** (a, b, c)	**21.** (b)	**22.** (d)	**23.** (c)	**24.** (d)
25. (c)	**26.** (d)	**27.** (d)	**28.** (b)	**29.** (a)	**30.** (b)	**31.** (a)	**32.** (a)

Answers : ASSERTION-REASON TYPE QUESTIONS

1. (d)	**2.** (d)	**3.** (b)	**4.** (a)	**5.** (b)	**6.** (c)	**7.** (b)	**8.** (a)
9. (c)	**10.** (c)	**11.** (c)	**12.** (d)	**13.** (a)	**14.** (a)	**15.** (d)	**16.** (a)
17. (b)	**18.** (a)	**19.** (a)	**20.** (c)	**21.** (b)	**22.** (a)	**23.** (c)	**24.** (b)
25. (d)	**26.** (e)						