Atomic Structure

1. ATOM & MOLECULES

(a) The smallest particle of a matter that takes part in a chemical reaction is called an atom. The atom of all gases except those of noble gases, cannot exist in free state. These exist in molecular form. The molecules of hydrogen, nitrogen, oxygen and halogens are diatomic (H_2, N_2). Phosphorus molecule is tetratomic and that of sulphur is octa atomic.

(b) The smallest particle of a matter that can exist in free state in nature, is known as a molecule.

(c) Some molecules are composed of homoatomic atom, e.g., H_2, O_2, N_2, Cl_2, O_3 etc., while the molecules of compounds are made up of two or more heteroatomic atoms e.g., HCl, NaOH, HNO_3, $CaCO_3$, etc.

2. DALTON'S ATOMIC THEORY

The concepts put forward by John Dalton regarding the composition of matter are known as Dalton's atomic theory. Its important points are as follows.

(a) Every matter is composed of very minute particles, called atoms that take part in chemical reactions.

(b) Atoms cannot be further subdivided.

(c) The atoms of different elements differ from each other in their properties and masses, while the atoms of the same element are identical in all respects.

(d) The atoms of different elements can combine in simple ratio to form compounds. The masses of combining elements represent the masses of combining atoms.

(e) Atom can neither be created nor destroyed.

2.1 Modern Concept :

Many of the concepts of Dalton's atomic theory cannot be explained. Therefore, foundation of modern atomic theory was laid down by the end of nineteenth century. The modern theory is substantiated by the existence of isotopes, radioactive disintegration, etc. The important points of the modern atomic theory are as follows.

(a) Prof. Henri Bacquerel discovered the phenomenon of radioactivity and found that an atom is divisible.

(b) An atom is mainly composed of three fundamental particles, viz. electron, proton and neutron.

(c) Apart from the aforesaid three fundamental particles, many others have also been identified, viz. positron, meson, neutrino, antiproton, etc.

(d) Soddy discovered the existence of isotopes, which were atom of the same element having different masses. For example, protium, deuterium and tritium are atoms of hydrogen having atomic masses 1, 2 and 3 a.m.u. respectively.

(e) Atoms having same mass may have different atomic numbers. These are known as isobars. For example, $^{40}_{18}Ar$ and $^{40}_{20}Ca$.

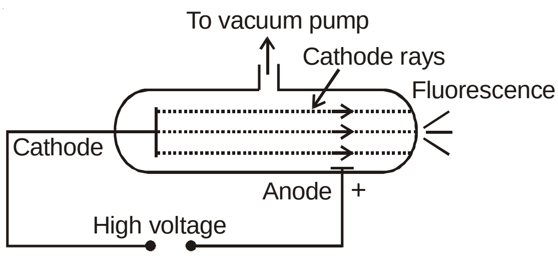
(f) Atoms of elements combines to form molecules.

(g) It is not necessary that the atoms should combine in simple ratio for the formation of compounds. The atoms in non-stoichiometric compounds are not present in simple ratio. For example, in ferrous sulphide crystals, iron and sulphur atoms are present in the ratio of 0.86 : 1.00.

(h) Atoms participate in chemical reactions.

3. CATHODE RAYS (DISCOVERY OF ELECTRON)

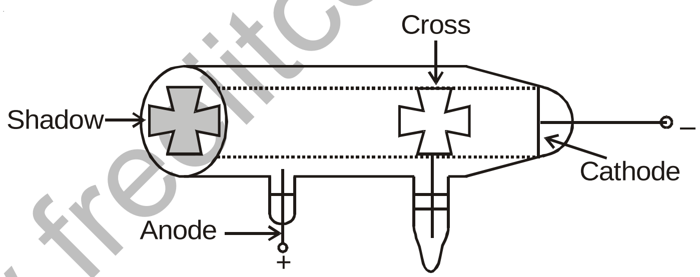
Dry gases are normally bad conductors of electricity. But under low pressure, i.e., 0.1 mm of mercury or lower, electric current can pass through the gases. Julius Plucker in 1859 found that a type of rays, called cathode rays, emit from the cathode when electricity is passed through a discharge tube. William Crookes (1879), J.J.Thomson and many other scientists studied the properties of cathode rays and came to the conclusion that the cathode rays of same properties are obtained using any gas or any cathode material.



Production of cathode rays

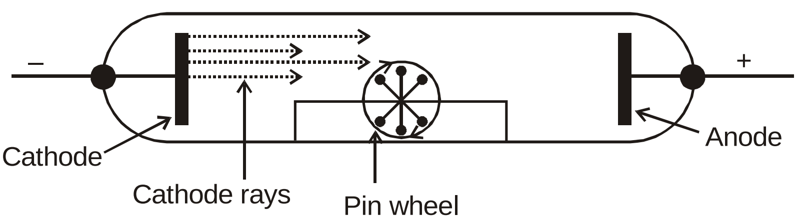
The salient features of cathode ray are as follows.

(a) Cathode rays travel in a straight line. This indicates that the formation of a shadow when an opaque object is placed in its path.



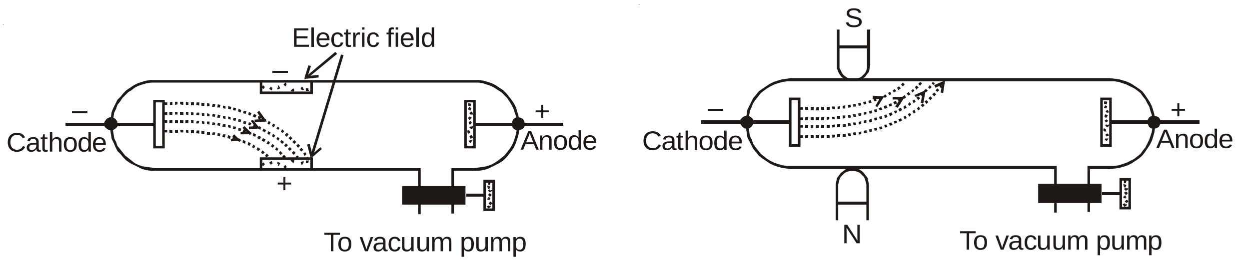
Cathode rays travel in a straight line

(b) If a light metal pinwheel is placed in the path of cathode rays, the wheel starts revolving. This proves that is cathode rays consist of tiny particles having momentum.



Cathode rays consist of tiny particles

(c) Cathode rays get deviated in electrical and magnetic fields. This proves that they are composed of charged particles. Their derivation towards anode indicates their negatively charged nature. The direction of their deviation in magnetic field depends on pole of the magnet which has been placed near the cathode ray tube.



(a) Electric Field (b) Magnetic Field

(d) Cathode rays produce green fluorescence on the walls of the glass tube.

(e) Cathode rays produce incandescence in at thin metal foil.

(f) Cathode rays effect the photographic plate.

(g) Cathode rays ionize gases proving that they are charged.

(h) Cathode rays penetrates across a thin metal foil.

(i) Cathode rays produce X-rays when they hit a piece of tungsten or any other metal having high melting point.

3.1 Nature of Cathode Rays :

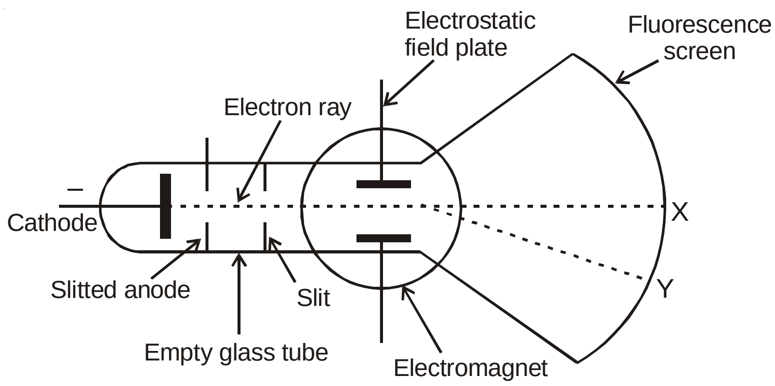
J.J. Thomson (1897) proved through experiments that.

(a) Cathode rays are composed of extremely tiny negatively charged particles (electrons).

(b) The ratio of negative charge (e) and mass (m) for cathode ray particle (electrons) is a constat. This ratio is independent of the material used in the preparation of the electrodes of the discharge tube or the gas filled in it. Thus, e/m of an electron is a universal constant.

$$\frac{\text{charge on electron}}{\text{mass of electron}} = \frac{e}{m} = 1.76 \times 10^{8} \text{ Coulomb/gm}$$

In addition to the above proofs, photoelectric effect, thermionic effect and emission of beta particles from radioactive elements also confirm that electron is an essential constituent of matter. These negatively charged tiny particles discovered by **Thomson**. It is denoted by e^- or ${}_{-1}e^0$.



Determination of $\frac{e}{m}$ of an electron

4. POSITIVE RAYS OR CANAL RAYS : DISCOVERY OF PROTON

Eugene Goldstein in 1886 found that a dim glow is visible behind the cathode when an electric discharge is passed through a perforated cathode in a discharge tube filled with a gas at low pressure. These new type of rays travel from anode to that cathode. Goldstein gave the name canal rays to these rays because these rays cross the canals of the cathode and reach the other side. **W.Wein** in 1897 proved through experiments that the canal rays consist of positively charged particles. **J.J. Thomson** gave the name **positive rays** to them because they are composed of positively charged particles.

4.1 Properties of Positive Rays :

(a) Positive rays travel in the direction opposite to that of cathode rays.

(b) Positive rays travel in straight line.

(c) Positive rays affect photographic plate.

(d) Positive rays are deviated in the electric and magnetic fields. The direction of their deviation proves the presence of positive charge on their particles.

(e) Positive rays pass across a very thin sheet of metal. But their penetrating power is less than that of cathode rays.

(f) Positive rays produce fluorescence and phosphorescence.

4.2 Nature of Positive Rays :

Thomson and Wein studied the nature of positive rays and proved with the help of experiments that.

(a) Positive rays are composed of positively charged particles.

(b) The ratio (e/m), of positive charge (e) and mass (m) for the particles of positive rays depends on the nature of the gas filled in the discharge tube. The value of e/m for the particles of positive rays obtained from different gases is different. The e/m value for positive rays is not a universal constant. Thomson and Wein found out through experiments that the maximum value of e/m is for particles of positive rays of hydrogen gas.

(c) Experiments proved that for a positively charged particle (H^+) of the positive rays of hydrogen gas $\frac{e}{m} = 9.578 \times 10^4$ coulomb per gram. If we suppose that the charge (e) of this particle is 1.602×10^{-19} coulomb unit positive charge, the mass (m) of the particle will be 1.6725×10^{-24} gram. The particle (H^+) of the positive rays of hydrogen gas having 1.602×10^{-19} coulomb positive charge and 1.6725×10^{-24} gram mass is called a proton.

5. PROTON

(a) Proton is a fundamental particle of an atom. It is an essential constituent of every matter.

(b) The credit for the discovery of proton goes to **Goldstein**.

(c) Proton bears one unit positive charge.

(d) **Thomson** and **Wein** estimated the value of e/m as 9.578×10^4 coulomb per gram for the positively charged particle proton.

(e) The amount of positive charge (e) on proton is 1.602×10^{-19} coulomb or 4.8×10^{-10} e.s.u.

(f) Mass of proton (m) $= 1.6725 \times 10^{-24}$ gram ; $= 1.6775 \times 10^{-17}$ kilogram
$= 1.6725 \times 10^{-29}$ quintal ; = 1837 times that of electron
= 1.00757 a.m.u. ; = Mass of hydrogen atom

Mass of proton (m) in a.m.u $= \frac{1.6725 \times 10^{-24}}{1.66 \times 10^{-24}} = 1.00757$ a.m.u.

(g) Mass of proton (m) multiplied by Avogadro number (6.023×10^{23}) gives molar mass of proton. Thus

Gram molecular mass of proton = $1.6725 \times 10^{-24} \times 6.023 \times 10^{23} = 1.008$ (Approx)

(h) Proton is present in the nucleus of an atom.

(i) The number of electrons is equal to the number of protons in a neutral atom.

(j) The atomic number of an atom is equal to the number of protons present in the nucleus of that atom.

(k) Proton is the nucleus of protium i.e. the common hydrogen atom.

(l) Proton is ionized hydrogen atom, i.e. (H^+)

(m) Proton is obtained when the only one electron present in hydrogen atom is removed. Hydrogen atom consists of only one electron and one proton.

6. ELECTRON (e^- or $_{-1}e^0$)

(a) Electron is a fundamental particle of an atom, which is an essential constituent of every matter.

(b) The credit for discovery of cathode rays goes to **Sir William Crookes** while the credit for discovery of negatively charged electron goes to **J.J. Thomson**. The name 'electron' was first given by **Stony**.

(c) A unit negative charge is present on electron.

(d) The value of $\frac{e}{m}$ was found to be 1.76×10^8 coulomb/gram by **Thomson**.

(e) **R.A. Mulliken** calculated the charge on an electron by his famous **Oil Drop Experiment**. The value came out to be 1.6012×10^{-19} coulomb or 4.803×10^{-10} e.s.u.

(f) The value of e/m of an electron is known as its **specific charge**. With the help of this specific charge and the charge on the electron (determined by Mulliken), the mass of the electron could be calculated as follows.

$$\frac{e}{e/m} = \frac{1.6012 \times 10^{-19}\,\text{coulomb}}{1.76 \times 10^{8}\,\text{coulomb/gram}} = 9.1091 \times 10^{-28}\ \text{gram}$$

$$= 0.0005486\ \text{a.m.u.}$$

$$= 1/1837^{th}\ \text{of H atom}$$

(g) Molar mass of electron is obtained on multiplying mass of electron by Avogadro number (6.023×10^{23}). Therefore gram molecular mass of electron is as follows.

$$= 9.1091 \times 10^{-28} \times 6.023 \times 10^{23}$$

$$= 5.483 \times 10^{-4}$$

(h) Electron is very much lighter than an atom of the lightest element hydrogen. The gram molecular mass of hydrogen is 1.008. Therefore the ratio of gram molecular mass of hydrogen and that of electron is

$\frac{1.008}{5.483 \times 10^{-4}} = 1837$. In other words, an atom of hydrogen (or a proton) is 1837 times heavier than electron.

$$\frac{\text{Mass of H atom}}{\text{Mass of electron}} = \frac{1.67 \times 10^{-24}}{5.483 \times 10^{-28}} = 1837$$

(i) The mass of 1.1×10^{27} electrons is one gram.

(j) The mass of one mole of electrons is 0.5583 mg.

(k) The amount of charge on one mole of electrons is one faraday or 96500 coulomb.

(l) The mass of an electron at rest is called static electron mass and its value is 9.1091×10^{-28} gram.

(m) The mass of an electron in motion is calculated with the help of the following expression.

$$\text{Mass of electron in motion (m)} = \frac{\text{Rest mass of electron}}{\sqrt{\left[1-\left(\frac{v}{c}\right)^2\right]}}$$

where v is velocity of electron and c is velocity of light.

When v = c, the mass of the electron in motion becomes infinity.

Therefore the mass of an electron increases with increase in its velocity due to which specific charge e/m on it decreases.

(n) Electron, being the fundamental particle of an atom, takes part in chemical combination.

(o) The physical and chemical properties of an element depend on the distribution of electrons in its outermost energy level.

7. DISCOVERY OF NEUTRON

Penetrating rays are emitted on bombarding α-particles on the elements like beryllium, boron and aluminium. **James Chadwick** in 1932 studied the nature of these radiation and came to the conclusion that these rays are composed of very tiny electro neutral particles. The mass of these particles is almost equal to that of the hydrogen atom. This particle is called neutron and is denoted by the symbol, ${}_0n^1$.

7.1 Neutron (${}_0n^1$)

(a) It is a fundamental particle of atom that is present in the nuclei of all atoms except hydrogen or protium.

(b) It was discovered by James Chadwick in 1932.

(c) It is an electro neutral particle, i.e. it does not have any positive or negative charge on it.

(d) The mass of a neutron is almost equal to that of a proton. Actually it is a little bit heavier than proton. Its mass (m) is as fallows :

Mass (m) of a neutron = 1.6748×10^{-24} gram = Approximately mass of a proton

(e) Neutron is relatively heavier out of the three fundamental particles of an atom.

(f) Molar mass of a neutron is obtained by multiplying the mass (m) of a neutron with Avogadro number (6.023×10^{23}). Therefore the gram molecular mass of a neutron is $1.6748 \times 10^{-24} \times 6.23 \times 10^{23} = 1.00893$.

(g) The atomic mass is equal to the total mass of all the protons and neutrons present in the atom.

(h) Isotopes are formed as a result of difference in the number of only neutrons in the nuclei of atoms.

(i) It is assumed that a neutron is a result to joining together of an electron and a proton. A neutron, being unstable, decays as fallows :

$${}_0n^1 \longrightarrow {}_{+1}P^{1-} + {}_{-1}e^0 + {}_0q^0 \text{ (antineutrino)}$$

Its half-life is 20 minutes.

(j) The density of neutrons is of the order of 1×10^{12} Kg/c.c.

8. OTHER PARTICLES OF ATOM

(a) **Positron :** It was discovered by **C.D. Anderson** in 1932. It beards a unit positive charge and its mass is equal to that of an electron. Thus its mass regarded as negligible. It merges with an electron and emit electromagnetic radiations. It is denoted by e^+.

(b) **Meson : Yukawa** in 1935 discovered this particle. Different types of meson particles are possible in the atom. These are called meson family.

(c) **Neutrino : Pauling** discovered these particles in 1927. They do not bear any charge, i.e. they are electro neutral particle.

(d) **Antiproton : Segre** discovered this particle in 1956. It bears a unit negative charge and its mass is equal to that of a proton.

9. CLASSIFICATION OF ATOMIC PARTICLES

9.1 Stable Particles

Properties of Stable Fundamental Particles

	Particle	Symbol	Charge	Mass*	Mass**	Spin***
1.	Proton	p	+	1.00758	1, 836	$\frac{1}{2}$
2.	Electron	e^-, β^-	–	0.0005486	1	$\frac{1}{2}$
3.	Positron	e^+, β^+	+	0.0005486	1	$\frac{1}{2}$
4.	Neutrino	v	0	0.000022	0.04	$\frac{1}{2}$
5.	Antiproton	p^-	–	1.00758	1, 836	$\frac{1}{2}$
6.	Graviton	G	0	0	0	2
7.	Photon	γ	0	0	0	1

*Physical atomic weight unit $_8O^{16}$ = 16,00,000

**Mass with respect to e, where e = 9.11×10^{-28} gram

*** $\frac{h}{2\pi}$ unit

9.2 Unstable Particles

Properties of Some Unstable Fundamental Particles

	Particle	Symbol	Charge	Mass*	Mass**	Spin***
1.	Neutron	n	0	1.00893	1, 836	$\frac{1}{2}$
2.	Negative μ meson	μ^-	–	0.1152	210	$\frac{1}{2}$
3.	Positive μ meson	μ^+	+	0.1152	210	$\frac{1}{2}$
4.	Neutral π meson	π^0	0	0.1454	265	0
5.	Negative π meson	π^-	–	0.1514	276	0
6.	Positive π meson	π^+	+	0.1514	276	0

9.3 Distinction of α, β and γ Rays

	Property	α Ray	β Ray	γ Ray
1.	Velocity	2×10^9 cm/sec	2.8×10^{10} cm/sec	Equal to velocity of light. 3×10^{10} cm/sec
2.	Penetration power	Very low	About 10 times to that of α rays	About 1000 times to that that of α rays
3.	Charge and mass	2 unit positive charge and 4 unit mass	1 unit negative charge and zero mass	Magnetic radiations of very high frequency
4.	Effect of ZnS plate	Produce fluoresence	No effect	No effect
5.	Nature	He^{+2} or helium nuclei, denoted by the symbol $_2He^4$.	Denoted by the symbol $_1e^0$ or electron	denoted by the symbol $_0\gamma^0$
6.	Megnetic field	Deviation towards cathode	Deviation towards anode	No effect
7.	Nature of the product	$_4A^9 \xrightarrow{\alpha^-} {}_2A^5$	$_4A^9 \xrightarrow{\beta^-} {}_5A^9$	$_4A^9 \xrightarrow{\gamma^-} {}_4A^9$

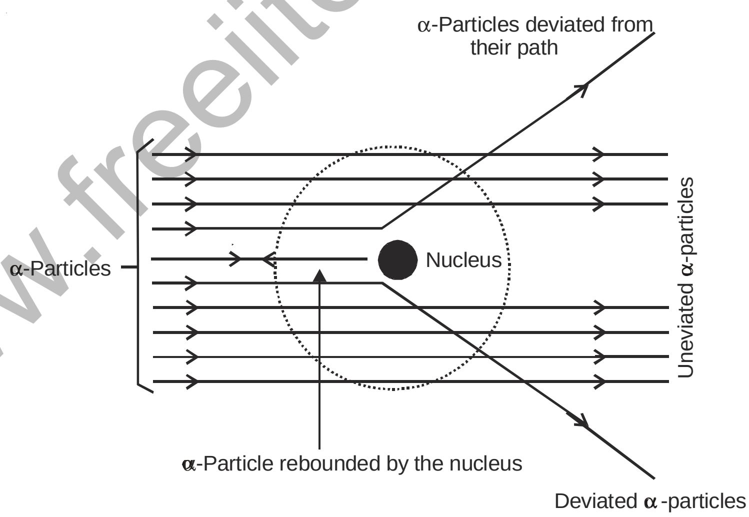
10. NUCLEUS

(i) **Rutherford** discovered the nucleus in an atom by **α-particle scattering experiment**. He showered α-particles, $_2He^4$ (obtained from radium) on a 0.01 mm thin gold film and allowed them to collide with a screen coated with zinc sulphide and placed behind the gold film. He observed fluorescence on the screen.

(a) Most of the α-particle passed through the gold film without deviating from their path.

(b) Some particles got deviated from their path on colliding with the gold film.

(c) A very small number of particles rebounded after colliding with the gold film.



(ii) The following are the inferences derived from the above experiment.

(a) Most of the α-particles pass through the gold foil without deviation in their path, showing that most of the part if an atom is vacant.

(b) Whole of the mass of an atom is confined to its nucleus, which consists of positively charged protons and neutral neutrons. These together are termed as nucleons.

(c) It has been found on the basis of calculation that the radius of the atomic nucleus is 1×10^{-13} to 1×10^{-12} cm or 1×10^{-15} to 1×10^{-14} meter, while radius of an atom is 1×10^{-8} cm.

(d) Magnitude of atomic nucleus = $\dfrac{\text{Radius of atom}}{\text{Radius of atomic nucleus}}$

(e) Nuclear density Density (D) = $\dfrac{\text{Mass (M)}}{\text{Volume (V)}}$

Since, the shape of atom is regarded as spherical, therefore, if radius of the nucleus is r, then

$$\text{Volume of nucleus} = \frac{4}{3}\pi r^3$$

Ex.1 Calculate the density of fluorine nucleus supposing that the shape of the nucleus is spherical and its radius is 5×10^{-13}. (Mass of F = 19 amu)

Sol. $\therefore$ Mass of the nucleus of F atom $= 19 \times 1.66 \times 10^{-24}$ gm

(1 amu = 1.660×10^{-24} gm)

Volume of the nucleus of F atom $= \dfrac{4}{3}\pi r^3$

$= \dfrac{4}{3} \times 3.14\,(5 \times 10^{-13})^3$

$= 525 \times 10^{-39}$ cm^3

Density of the nucleus of F atom $= \dfrac{\text{Mass}}{\text{Volume}}$

$= \dfrac{19 \times 1.66 \times 10^{-24}\,\text{gm}}{525 \times 10^{-39}\,\text{cm}^3}$

$= 6.0 \times 10^{13}$ gm cm^{-3}

11. NUCLEAR CHARGE AND ATOMIC NUMBER

Positive charge on the nucleus of an atom is equal to the atomic number of that atom. A scientist named **Mosley** studied the frequency of X-rays emitted by showering high velocity electrons on a metal and established the following relationship.

$$\sqrt{\nu} = a\,(z - b)$$

where ν = frequency of X-rays

z = atomic number or nuclear charge

a and b are constants.

Thus nuclear charge of an atom is equal to the atomic number of that atom. Since an atom is electro neutral, the number of positively charged protons in its nucleus is equal to the negatively charged electrons moving around the nucleus in the atom. Thus

Atomic number = number of protons in the atom or

number of electrons in the atom

12. ATOMIC WEIGHT OR MASS NUMBER

The value of mass number of an atom (in amu) is always a whole number.

Mass number of an atom is the sum of number of protons and number of neutrons present in that atom.

Mass number = Number of protons (Z) + Number of neutrons (n)

= Atomic number + Number of neutrons

For example	$_8O^{16}$	$_7N^{14}$	$_{11}Na^{23}$
Protons	8	7	11
Neutrons	8	7	12
Atomic weight	16	14	23

(a) The protons and neutrons present in the nucleus are known as nucleons.

(b) The weight of electrons is neglected during calculation of the atomic weight, because the mass of an electron is negligible in comparison to that of a proton or a neutron.

(c) In the nucleus of an electro neutral atom, the number of positively charged protons is equal to that of negatively charged electrons.

Particle	$_8O^{16}$	$_7N^{14}$	$_{11}Na^{23}$	$_6C^{12}$	$_9F^{17}$
Protons	8	7	11	6	9
Neutrons	8	7	12	6	8
Atomic weight	16	14	23	12	17
Electrons	8	7	11	6	9

(d) The number of protons present in an atom is called atomic number of that atom.

For example	O	F	Ne
Protons	8	9	10
Atomic number	8	9	10

(e) **Kernel :** The group of all the electrons except those of the outermost energy level, is called that kernel of that atom and the electrons present in the kernel are known as electron of the kernel.

For example, if the electronic configuration of an atom is 2, 6, then the number of kernel electrons is 2.

If the electronic configuration of an atom is 2, 8, 8, then the number of kernel electrons is 10.

If the electronic configuration of an atom is 2, 8, 8, 8, then the number of kernel electrons is 18.

13. IONS

When an atom loses electron, it is converted into a cation, while it is converted into an anion on gaining electron.

(a) Number of electrons in a cation = Number of protons - charge present on the cation

(b) Number of electrons in a anion = Number of protons + Charge present on the anion

For example	Na^{+}	Mg^{+2}	Al^{+3}
Protons	11	12	13
Electrons	10	10	10
	Cl^{-}	O^{-2}	F^{-}
Protons	17	8	9
Electrons	18	10	10

Ex.2 What difference will appear in the mass number if the number of neutrons is halved number of electrons is doubled in $_8O^{16}$.

[1] 25% decrease [2] 50% increase [3] 150% increase [4] No difference **Ans.[1]**

Sol. $_8O^{16}$ Initial weight – final weight

Protons	8p	$\rightarrow$	8p
Neutrons	8n	$\rightarrow$	4n
Weight	16	$\rightarrow$	12

Thus decrease in mass number = 25%

Ex.3 If the atomic weight of Zn is 70 and its atomic number is 30, then what will be the atomic weight of Zn^{+2} ?

[1] 70 [2] 68 [3] 72 [4] 74 **Ans. [1]**

Sol. Two electrons are removed in the formation of Zn^{+2} from Zn. The numbers of protons and neutrons remain unchanged.

Ex.4 The mass of one mole electrons should be -

[1] 0.55 mg [2] 1.0008 gm [3] 1.000 gm [4] 0.184 gm **Ans. [1]**

Sol. The number of electrons in one mole = Avogadro number = 6.023×10^{23}

Mass of one electron = 9.1×10^{-28} gm

Therefore Mass of 6.023×10^{23} electrons

$= 9.1 \times 10^{-28} \times 6.023 \times 10^{23} = 0.55$ mg

Ex.5 The number of atoms present in 20 grams of calcium will be equal to the number of atoms present in

(20 gm Ca = $\frac{1}{2}$ mole Ca)

$$Ca = \frac{6.023\times10^{23}}{2} = 3.012 \times 10^{23}$$

[1] 12 gm C [2] 12.15 gm Mg [3] 24.0 gm C [4] 24.3 gm Mg **Ans. [2]**

Sol. 24.3 gm Mg = 1 mole, therefore 12.15 gm = $\frac{1}{2}$ mole

14. ISOTOPES

(a) The atoms of the same element having same atomic number but different atomic weights, are called isotopes.

(b) Isotopes of an element have same number of protons but different number of neutrons in their atoms. Hence their atomic weight are different. For example, oxygen has the following three isotopes.

	$_{8}O^{16}$	$_{8}O^{17}$	$_{8}O^{18}$
Protons	8	8	8
Neutrons	8	9	10
Atomic weights	16	17	18

(c) Hydrogen has the following three isotopes.

	$_{1}H^{1}$ (Protium)	$_{1}D^{2}$ (Deuterium)	$_{1}T^{3}$ (Tritium)
Protons	1	1	1
Neutrons	0	1	2
Atomic weights	1	2	3

(d) Chlorine has the following two isotopes.

$_{17}C^{35}$ and $_{17}Cl^{37}$

Ex.6 What should be the percentage of deuterium in heavy water ?

[1] 20% [2] 80% [3] 60% [4] 40% **Ans. [1]**

Sol. Deuterium in 20 parts of D_2O = 4 parts

Deuterium in 100 parts of D_2O = $\frac{4}{20} \times 100 = 20\%$

Ex.7 Which of the following pairs consists of molecules having same mass number ?

[1] H_2O and D_2O [2] H_2O and HTO [3] D_2O and HTO [4] D_2O and HCl **Ans. [3]**

Sol. Mass number of H_2O = 18

Mass number of D_2O = 20

Mass number of HTO = 20

Mass number of HCl = 36.5

Ex.8 The mass number of three isotopes of an element are 11, 12 and 13 units. Their percentage abundance is 80, 15 and 5 respectively. What should be the atomic weight of the element ?

[1] 11.25 [2] 20 [3] 16 [4] 10 **Ans. [1]**

Sol. 80 : 15 : 5

Thus the ratio is 16 : 3 : 1

Total = 16 + 3 + 1 = 20

Average weight = $\frac{11 \times 16 + 12 \times 3 + 13 \times 1}{20}$ = 11.25

Ex.9 If two neutrons are added to an element X, then it will get converted to its

[1] isotope [2] isotone [3] isobar [4] None of the above **Ans. [1]**

Sol. The number of neutrons are different in the isotopes of the same element.

15. ISOBARS

(a) Isobars are the atoms of different elements having same atomic weight.

(b) Isobars have different numbers of protons as well as neutrons.

(c) The sum of number of protons and neutrons in isobars is same. For example

Atomic weight of three elements $_{18}Ar^{40}$, $_{19}K^{40}$ and $_{20}Ca^{40}$ is 40.

(i)	Ar^{40}	K^{40}	Ca^{40}
Protons	18	19	20
Neutrons	22	21	20

(ii)	$_{32}Ge^{76}$	$_{34}Se^{76}$
Protons	32	34
Neutrons	44	42

16. ISOTONES

The atoms having same number of neutrons are called isoneutronic or isotones. For example

	$_{14}Si^{30}$	$_{15}P^{31}$	$_{14}S^{32}$
Protons	14	15	16
Neutrons	16	16	16
Atomic weight	30	31	32

Ex.10 In two elements $_{Z_1}A^{M_1}$ and $_{Z_2}B^{M_2}$, $M_1 \neq M_2$ and $Z_1 \neq Z_2$ but $M_1 - Z_1 = M_2 - Z_2$. These elements are

[1] isotonic [2] isotopic [3] isobaric [4] isoprotonic **Ans [1]**

Sol. M_1 = Atomic weight Z_1 = Atomic number

In isobars $M_1 = M_2$ and in isotopes $Z_1 = Z_2$

In isotones (isoneutronic elements) $M_1 - Z_1 = M_2 - Z_2$

Ex.11 Two nuclides A and B are isoneutronic. Their mass numbers are 76 and 77 respectively. If atomic number of A is 32, then the atomic number of B will be

[1] 33 [2] 34 [3] 32 [4] 30 **Ans. [1]**

Sol.

	$_{32}A^{76}$		$_{P}B^{77}$
Protons	= 32	Protons + Neutrons	= 77
Protons + Neutrons	= 76	Neutrons	= 44
Neutrons	= 44	Protons	= 33

17. ISOELECTRONIC

The chemical species in which number of electrons is same are called isoelectronic. For example

(a)	Li^+	Be^{+2}	B^{+3}		
Electrons	2	2	2		
(b)	Na^+	Mg^{+2}	Al^{+3}	F^-	O^{-2}
Electrons	10	10	10	10	10
(c)	K^+	Ca^{+2}	Ar		
Electrons	18	18	18		

Ex.12 The isoelectronic pair of 32 electron is

[1] BO_3^{-3} and CO_3^{-2} [2] PO_4^{-3} and CO_3^{-2} [3] N_2 and CO [4] All of the above **Ans. [1]**

Sol. BO_3^{-3} CO_3^{-2}

5 + 24 + 3 = 32 6 + 24 + 2 = 32

Ex.13 The pair $NH_3 + BH_3$ is isoelectronic with

[1] B_2H_6 [2] C_2H_6 [3] C_2H_4 [4] CO_2 **Ans. [2]**

Sol. $NH_3 + BH_3$ C_2H_6

7 + 3 + 5 + 3 = 18 6 × 2 + 6 = 18

Ex.14 Which of the following is a one-electron species ?

[1] He [2] N [3] H_2 [4] N_2 **Ans. [4]**

Sol. There is only one electron in H_2^+

Ex.15 The molecular weight of an oxide of nitrogen is 30. What should be the number of electrons in it ?

[1] 15 [2] 30 [3] 45 [4] 20 **Ans. [1]**

Sol. The molecular weight of NO is 30. It will have 15 electrons.

Ex.16 A diapositive ion has 16 protons. What should be the number of electrons in its tetrapositive ion.

[1]16 [2] 14 [3] 12 [4] 10 **Ans. [3]**

Sol. X^{+2} has 16 protons, then In X – 16 protons and 16 electrons

In X^{+2} – 16 protons and 14 electrons In X^{+4} – 16 protons and 12 electrons

Ex.17 If atomic weights of C and Si are 12 and 28 respectively, then what is the ratio of numbers of neutrons in them

[1] 1 : 2 [2] 2 : 3 [3] 3 : 4 [4] 3 : 7 **Ans. [4]**

Sol. Number of neutrons in ${}_6C^{12}$ = 12 – 6 = 6 ; Number of neutrons in ${}_{14}Si^{28}$ = 28 – 14 = 14

The ratio of number of neutrons in C and Si is 6 : 14 or 3 : 7.

18. ATOMIC MODEL

18.1 Thomson's Model of an Atom

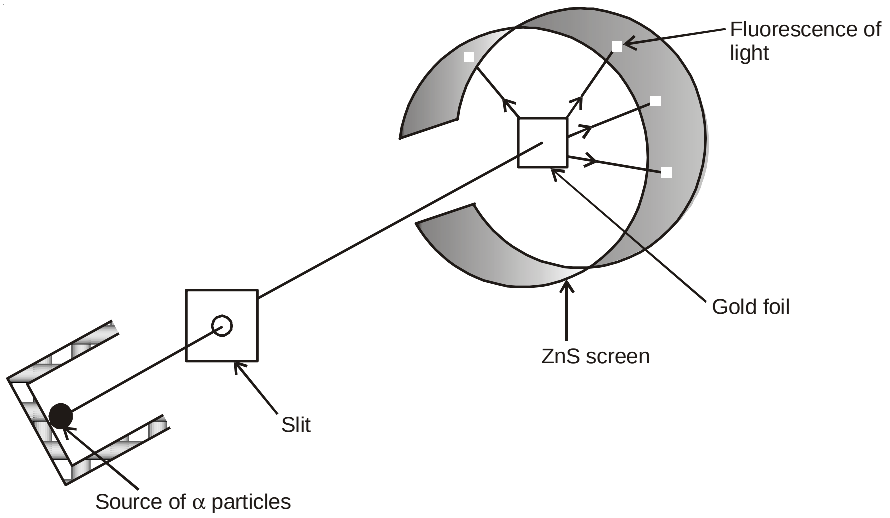
(i) Atom is a very minute, spherical, electro neutral particle that consists of positively and negatively charged matter.

(ii) The positively charged matter is uniformly distributed in the atom and the negatively charged electrons are embedded in it just as the seeds in water melon. Therefore, Thomson model of an atom is also called "water melon model".

(iii) Thomson's model of an atom failed to explain the production of the atomic spectrum. It cannot explain Rutherford's α particle scattering experiment also.

18.2 Rutherford's Model of an Atom

Ernest Rutherford in 1911put forward the **"nuclear model"** of atom on the basis of α particle scattering experiment. In this experiment, **Rutherford** showered α–particles (Helium nuclei, He^{+2}) on a thin gold foil and observed that most of the α-particles travelled straight without deviation in the direction of their path, some of them deviate from their path by different angles, while very few get rebounded after colliding with the foil. Rutherford gave the following nuclear model on the basis of the experiment.



Rutherford's α particle scattering experiment

(i) Atom is a very minute, spherical, electro neutral particle composed of the following two parts :

(a) Positively charged nucleus and **(b)** a vast extranuclear space in which electrons are present.

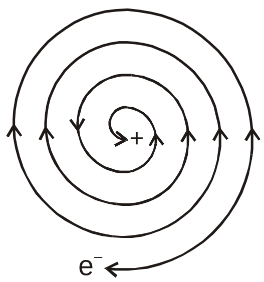
(ii) whole of the positive charge and almost all the mass of atom is confined to a very minute part at the centre of the atom, called the nucleus of the atom. The radius of nucleus is about 10^{-13} to 10^{-12} cm (or 10^{-15} to 10^{-14} meter), while the radius of atom is in the order of 10^{-8} cm.

(iii) The number of electrons in an atom is equal to the number of protons present in the nucleus. That is why an atom is electroneutral.

(iv) This model of an atom is also called "solar model" of "planetary model". This is because, the movement of electrons around the nucleus in this model has been compared to that of planets moving around the sun in the solar system.

18.2.1 Demerits of Rutherford's Model of an Atom

(i) According to Clark Maxwell's theory of electrodynamics, an electrically charged particle in motion continuously emits energy. This results in regular decrease in the energy of that particle. On the basis of this principle, it can be concluded that an electron moving around the nucleus will continuously emit the energy. This will result in decrease in the radius of the electron orbit, due to which the electron would ultimately plunge into the nucleus.



An electron emitting energy and pluging into nucleus

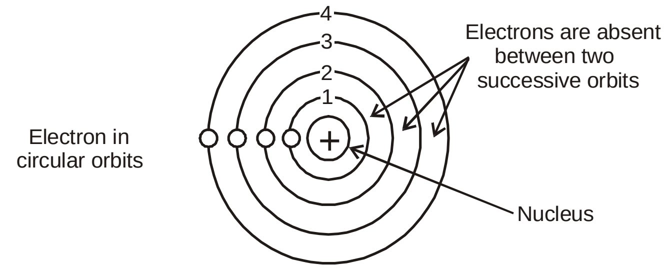
(ii) Plunging of an electron into the nucleus would definitely mean destruction of the atom or end of the existence of the atom. But we know that it never happens. Atom is a stable system. Therefore Rutherford model failed in explaining the stability of an atomic system.

(iii) If an electron moving around the nucleus continuously emits energy, then the atomic spectrum must be continuous, i.e. the spectrum should not have lines of definite frequency. However, the atomic spectrum is actually not continuous and possesses so many lines of definite frequency. Therefore, Rutherford model failed to explain the line spectrum of an atom.

18.3 Bohr's Model of an Atom

Neil Bohr in 1913 presented a quantum mechanical model of atomic structure.

(i) An electron moves around the nucleus in constant circular orbits.

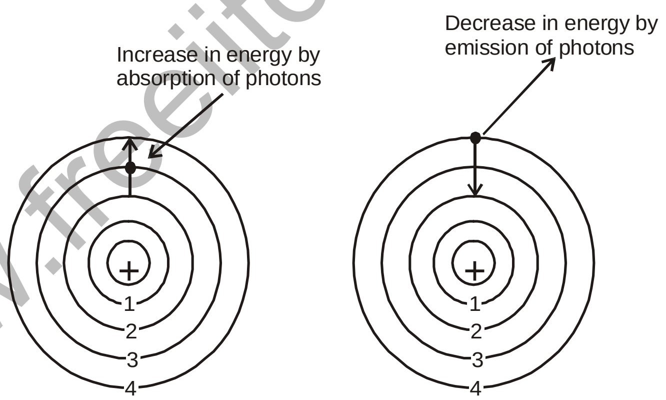


Electrons moving around in circular orbits

(ii) The electrons moving around the nucleus in only those circular orbits for which their angular momentum (mvr) is integral multiple of $\frac{h}{2\pi}$. This is called the condition of quantization. The angular momentum (mvr) of an electron is $\frac{nh}{2\pi}$ where m is the mass of electron. r is radius of its circular orbit, v is the velocity of electron, h is Planck's constant; n is a whole number whose value may be 1, 2, 3, 4 etc, : n is called principal quantum number.

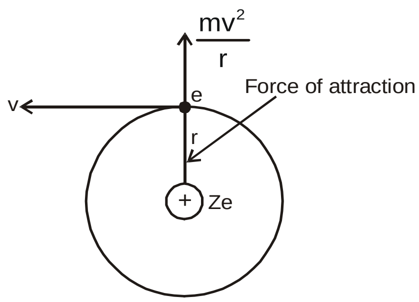
(iii) When energy is provided to an atom, its electrons get excited by absorption of energy and jumps to the orbits of higher energy.

(iv) When an electron in an atom falls from higher energy level to lower energy level, spectral lines are formed.



(a) Electron jumping from lower to higher energy level by absorption of energy
(b) Electron jumping from higher to lower energy level by emission of energy

(v) The force of attraction on electron by the nucleus is equal to the centrifugal force of that electron.



The electron moving in an orbit by various forces

Note : (a) $mvr = \frac{nh}{2\pi}$........[1] (b) $\frac{mv^2}{r} = \frac{Ze^2}{r^2}$........[2] (c) $E_{n_2} - E_{n_1} = h\nu$...........[3]

Ex.18 An electron has been excited from the first to the fourth energy state in an atom. Which of the following transitions are possible when the electron comes back to the ground state ?

[1] $4 \rightarrow 1$ [2] $4 \rightarrow 2, 2 \rightarrow 1$ [3] $4 \rightarrow 3, 3 \rightarrow 2m\ 2 \rightarrow 1$[4] All of the above **Ans. [4]**

Sol. Electron can undergo transition from higher state to all lower states by loss of energy.

Ex.19 How much total energy will be released when an electron present in hydrogen atom undergoes the following sequence of transition ?

$$n = 4 \rightarrow n = 2 \rightarrow n = 1$$

[1] One quantum [2] Two quantums [3] Three quantums [4] Four quantums **Ans. [2]**

Sol. One quantum of energy is released in each transition, i.e. one quantum in n = 4 to n = 2 and one quantum in n = 2 to n = 1 transition.

Ex.20 Which of the following is a fundamental particle

[1] Nucleus of He [2] Nucleus of H [3] A positive atom [4] None of these **Ans. [2]**

Sol. Fundamental particle H^+ is the nucleus of H

19. CALCULATION OF VELOCITY OF THE ELECTRON OF BOHR'S ORBIT

$$\frac{mv^2}{r} = \frac{Ze^2}{r^2} \quad \textbf{.............[1]}$$

From Bohr's postulate

$$mvr = \frac{nh}{2\pi} \quad \textbf{...............[2]}$$

Eq. [1] divided by [2]

$$\boxed{v = \frac{2\pi Ze^2}{nh}} \quad \text{or} \quad v = K\frac{z}{n}$$

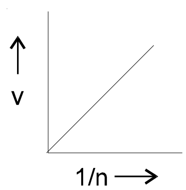
Here π , e and h are constants, therefore

Here $K = \frac{2\pi e^2}{h} = 2.188 \times 10^8$ cm/second

or $\boxed{v = \frac{Z}{n} \times 2.188 \times 10^8 \text{ cm/second}}$

(a) If Z is a constant, then

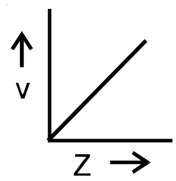
$\boxed{v \propto \frac{1}{n}}$



Therefore, velocity goes on decreasing with increase in the number of orbits.

Thus

$$\boxed{\frac{v_1}{v_2} = \frac{n_2}{n_1}}$$



(b) If n is a constant, then

$v \propto Z$

Therefore, velocity goes on increasing with increase in the atomic number.

$$\frac{v_1}{v_2} = \frac{z_1}{z_2}$$

(c) Time period T = $\frac{2\pi r}{V}$

$$= \frac{2\pi \times n^2h^2}{4\pi^2 mZe^2} \times \frac{nh}{2\pi Ze^2}$$

$$= \frac{n^3h^3}{4\pi^2 mZ^2e^4}$$

(d) Frequency $\frac{1}{T} = \frac{V}{2\pi r}$

Ex.21 The expression for calculation of velocity is

[1] $v = \left(\frac{Ze^2}{mr}\right)^{\frac{1}{2}}$ [2] $v = \frac{2\pi Ze^2}{nh}$ [3] $v = \frac{nh}{2\pi mr}$ [4] all of the above are correct

Ans. [4]

Sol. [1] $v^2 = \frac{Ze^2}{mr}$ $v = \left(\frac{Ze^2}{mr}\right)^{\frac{1}{2}}$

[2] $v = \frac{2\pi Ze^2}{nh}$

[3] $mvr = \frac{nh}{2\pi}$ $v = \frac{nh}{2\pi mr}$

Ex. 22 If the velocities of first, second, third and fourth orbits of hydrogen atom are v_1 , v_2 , v_3 and v_4 respectively, then which of the following should be their increasing order

[1] $v_1 > v_2 > v_3 > v_4$ [2] $v_4 < v_3 < v_2 < v_1$ [3] $v_1 > v_2 < v_3 > v_4$ [4] Equal for all **Ans. [2]**

Sol. Z is a constant, therefore $v \propto \frac{1}{n}$

i.e. $v_4 < v_3 < v_2 < v_1$

Ex.23 The ratio of velocities of electrons present in Na^{+10} and H should be

[1] 11 : 1 [2] 11:3 [3] 1 : 11 [4] 4 : 11

Sol. $Na^{11} \rightarrow Na^{+10} \rightarrow Is^1$

Thus, n is a constant

Therefore $\frac{v_1}{v_2} = \frac{Z_1}{Z_2} = \frac{11}{1}$

Ex.24 What should be the velocity of the electron present in the fourth orbit of hydrogen atom, if the velocity of the electron present in the third orbit is 7.29×10^7 cm per second ?

Sol. $\frac{v_3}{v_4} = \frac{n_4}{n_3}$

$$\frac{7.29 \times 10^7}{v_4} = \frac{4}{3}$$

$$v_4 = \frac{7.29 \times 10^7 \times 3}{4} = 5.46 \times 10^7 \text{ cm per second}$$

20. RADIUS OF n^th BOHR'S ORBIT

According to Bohr's hypothesis

put the value of v in $mvr = \frac{nh}{2\pi}$

$$\boxed{r = \frac{n^2h^2}{4\pi^2 mZe^2}} \quad \text{or} \quad \boxed{r = K\frac{n^2}{z}}$$

In the above expression h, π, m and e, all are constants. therefore

$$\left(K = \frac{h^2}{4\pi^2 me^2} = \text{constant} = 0.529 \text{Å}\right)$$

or $$\boxed{r = \frac{n^2}{Z} \times 0.529 \text{ Å}}$$

Note : (a) 1Å = 10^{-8} cm

(b) 1Å = 10^{-10} m

(c) 1 nm = 10^{-9} m

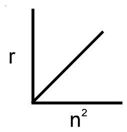
(d) 1 pm (picometer) = 10^{-10}cm

If Z is a constant, then

$r \propto n^2$

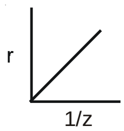
Thus, the radius of atoms goes on increasing as the number (n) of energy levels in the atoms goes on increasing as shown below.

$$\boxed{\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2}}$$



(b) If n is a constant, then

$$\boxed{\frac{r_1}{r_2} = \frac{Z_2}{Z_1}}$$



Ex.25 If the radius of first, second, third and fourth orbits of hydrogen atom are $r_{1,}$ $r_{2,}$ r_3 and r_4 respectively, then their correct increasing order will be

[1] $r_4 < r_3 < r_2 < r_1$ [2] $r_1 < r_2 < r_3 < r_4$ [3] $r_1 > r_2 > r_3 > r_4$ [4] Equal in all **Ans. [2]**

Sol. $r \propto n^2$

$$\boxed{r_1 < r_2 < r_3 < r_4}$$

Ex. 26 The ratio of radius of the fifth orbits of He^+ and Li^+ will be

[1] 2 : 3 [2] 3 : 2 [3] 4 : 1 [4] 5 : 3 **Ans. [2]**

Sol. Here n is a constant, therefore

$$\frac{r_1}{r_2} = \frac{Z_2}{Z_1} = \frac{3}{2} = 3 : 2$$

Ex. 27 Which of the following orbits of hydrogen atom should have the values of their radii in the ratio of 1 : 4 ?

[1] K and L [2] L and N [3] M and N [4] 1 and 2 both are correct

Ans. [4]

Sol. [1] Ratio of radii of orbits K and L

$$\frac{r_1}{r_2} = \frac{n_1^{\;2}}{n_2^{\;2}} = \frac{1^2}{2^2} = 1:4$$

[2] Ratio of radii of orbits L and N

$$\frac{r_1}{r_2} = \frac{n_1^{\;2}}{n_2^{\;2}} = \frac{2^2}{4^2} = 4 : 16 \text{ or } 1 : 4$$

Ex. 28 If $a = \frac{h}{4\pi^2 me^2}$, then the correct expression for calculation of the circumference of the first orbit of hydrogen atom should be

[1] $\sqrt{4h^2}\pi a$ [2] $2\pi r$ [3] $\sqrt{4}\pi ha$ [4] 1 and 3 both are correct

Ans. [4]

Sol. Circumference = $2\pi r$

$2 \times \pi \times \frac{n^2h^2}{4\pi^2 mZe^2}$, n = 1 , Z = 1 and $\frac{h}{4\pi^2 me^2} = a$

Thus $2 \times \pi \times h \times a$ or $\sqrt{4}\pi ha$ or $\sqrt{4h^2}\pi a$

21. ENERGY OF ELECTRON IN BOHR'S nth ORBIT

(a) The energy of an electron is negative because according to Bohr's hypothesis, the maximum energy of an electron at infinity is zero. Therefore, value of energy should be negative on moving towards lower side from infinity.

(b) The energy of electron at infinity is zero because attractive force between electron and the nucleus is minimum.

(c) Stability would increase as the electron in an atom moves from the infinity distance to a distance r from the nucleus, resulting in the value of the potential energy becoming negative. This is because of the fact that when two opposite charges attract each other, there is a decrease in the potential energy, as attractive forces = $\frac{Ze^2}{r^2}$

(d) Potential energy of electron is negative while kinetic energy is positive.

(e) Total energy is negative and the negative value shows that attractive forces are working between electron and nucleus. Therefore, work is to be done to remove the electron from this equilibrium state.

(f) Energies are of two types.

21.1 Kinetic Energy (E_K)

This energy is produced due to the velocity of electron. If mass is m, velocity is v and radius is r then

Kinetic energy = $\frac{1}{2}mv^2 = \frac{1}{2}\frac{Ze^2}{r}$

21.2 Potential Energy (E_P)

This energy is produced due to electrostatic attractive forces between electron and proton, and its value is negative. If atomic number is Z. charge is e and radius is r, then

Potential energy = $\frac{-Ze^2}{r}$

21.3 Total Energy (E_T)

Total energy = Kinetic energy + potential energy $\boxed{E_T = E_K + E_P}$ $\frac{1}{2}mv^2 + \frac{-Ze^2}{r}$

Total energy $E = -\frac{1}{2}\frac{Ze^2}{r}$

Formula

(i) Total energy = – Kinetic energy ($E_T = -E_K$)

(ii) Potential energy = 2 × Total energy) ($E_P = 2E_T$)

Ex. 29 What should be the kinetic energy and total energy of the electron present in hydrogen atom, if its potential energy is –5.02 eV

Sol. (a) Total energy $= \frac{\text{Potential energy}}{2} = \frac{-5.02}{2}$

Total energy = – 2.51 eV

(b) Kinetic energy = – Total energy

= – (–2.51 eV)

= + (2.51 eV)

21.4 Calculation of energy

Formula :- $\because$ $E = -\frac{1}{2}\frac{Ze^2}{r}$ put the value of r

$$E_T = -\frac{Z^2}{n^2}\times\frac{2\pi^2me^4}{h^2} \quad \text{or} \quad = -K\frac{Z^2}{n^2}$$

where $K = \frac{2\pi^2me^4}{h^2}$ = A constant, whose values can be depicted as follows

(a) = 13.60 eV per atom

(b) = 2.179×10^{-11} ergs per atom

(c) = 313.6 kilocalories per mole

(d) = 21.79×10^{-19} joules per atom

(e) = 1312.1 kilojoules per atom

Note : Units – (a) 1 erg = 10^{-7} joule

(b) 1 erg = 6.2419×10^{11} eV

(c) 1 eV = 23.06 kilocalories

(d) 1 eV = 1.602×10^{-12} ergs

(e) 1 joule = 6.2419×10^{18} eV

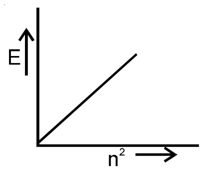
(f) 1 kilocalorie = 4.184 kilojoule

(i) $E = -\frac{Z^2}{n^2} \times 13.6$ eV

If Z is a constant, then $E \propto -\frac{1}{n^2}$

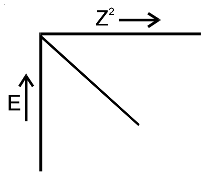
Therefore, the energy of electron goes on increasing as the number of orbits increases.

$$\frac{E_1}{E_2} = \frac{n_2^2}{n_1^2}$$



(ii) If n is a constant, then $E \propto -Z^3$

$$\frac{E_1}{E_2} = \frac{Z_1^2}{Z_2^2}$$



$$E = \frac{Z^2}{n^2} \times Rhc$$

[1] Kinetic energy = $\frac{Z^2}{n^2} \times Rhc$ [2] Potential energy = $2\left(-\frac{Z^2}{n^2} \times Rhc\right)$

Ex.30 What should be the order E_1, E_2, E_3 and E_4, if these are the respective energies of the first, second, third and fourth orbits of hydrogen atom ?

[1] $E_1 = E_2 = E_3 = E_4$ [2] $E_4 < E_3 < E_2 < E_1$ [3] $E_1 < E_2 < E_3 < E_4$ [4] $E_2 > E_3 < E_4 < E_1$ **Ans. [3]**

Sol. $E \propto -\frac{1}{n^2}$

Ex.31 What should be kinetic energy and potential energy, respectively, of the electron in the third orbit of hydrogen atom ?

[1] – 1.5 eV, 3.0 eV [2] 1.5 eV, – 3.0 eV [3] 1.5 eV, 3.0 eV [4] 3.0 eV, – 3.0 eV **Ans. [2]**

Sol. Total energy of the third orbit of H atom

$$E = -\frac{Z^2}{n^2} \times 13.6$$

$$= -\frac{1}{9} \times 13.6 = -1.5\text{eV}$$

[1] Kinetic energy = – Total energy

= – (–1.5) = + 1.5 eV

[2] Potential energy = 2 × Total energy

= 2 × –1.5 = –3.0 eV

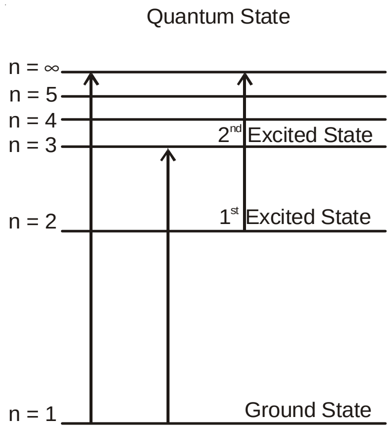
Ex.32 What should be the ratio of energies of the electrons of the first orbits of Na^{+10} and H ?

[1] 11:1 [2] 121 : 1 [3] 1 : 121 [4] 1 : 11 **Ans. [2]**

Sol. Here n is a constant, therefore

$$\frac{E_1}{E_2} = \frac{Z_1^2}{Z_1^2} = \frac{(11)^2}{(1)^2} = 121 : 1$$

21.5 Quantization of electronic energy Levels



Quantum State	Energy	Separation Energy	Excitation Energy
n = ∞	0	0	13.6
n = 5	$-13.6/5^2 = -0.54$	0.54	13.05
n = 4	$-13.6/4^2 = -0.65$	0.85	12.75
n = 3	$-13.6/3^2 = 1.51$	1.51	12.1
n = 2	$-13.6/2^2 = -3.4$	3.4	10.2
n = 1	–13.6 eV	13.6 eV	0 eV

Electronic energy levels of hydrogen atoms

21.6 Ground state

Ans atom in its lowest energy state or initial state is said to be in ground state. This is the most stable of an atom.

Ex.33 Which of the following should be the energy of an electron present in ground state of hydrogen atom ?

[1] – 13.6 eV [2] – 3.4 eV [3] – 1.5 eV [4] – 0.85 eV **Ans. [1]**

Sol. An electron in ground state is in n = 1 orbit. Therefore the energy of the electron = – 13.6 eV

21.7 Excited State

The states of higher energy than the ground state are said to be in excited state. For example, the electron of hydrogen atom in ground state is present in n = 1 orbit.

(a) Electron in n = 2 orbit is in first excited state

(b) Electron in n = 3 orbit is in second excited state

(c) Electron in n = 4 orbit is in third excited state

This means that the energy of n + 1 orbit is in first excited state, of n + 2 orbit in second excited state and of n + 3 orbit in third excited state, where n = the energy in ground state.

Ex.34 What should be the energy of the second excited state of Li^{+2} ?

[1] – 13.6 eV [2] – 30.6 eV [3] – 3.4 eV [4] – 1.5 eV **Ans. [1]**

Sol. Second excited state n = 3

$$E_n = -13.6 \times \frac{Z^2}{n^2}$$

$$= -13.6 \times \frac{3^2}{3^2}$$

$$= -13.6 \text{ eV}$$

21.8 Excitation Potential

(a) The energy required to excite an electron from ground sate to any excited state is known as excitation potential.

(b) Excitation potential has a positive value. For example,

First excitation potential of hydrogen atom = $E_2 - E_1$

Second excitation potential of hydrogen atom = $E_3 - E_1$

Third excitation potential of hydrogen atom = $E_4 - E_1$

Ex.35 How much minimum energy should be absorbed by a hydrogen atom in ground state to reach excites state ?

[1] + 10.2 eV [2] + 13.4 eV [3] + 3.4 eV [4] + 1.5 eV **Ans. [1]**

Sol. The electron has to go to the second orbit E_2 on excitation. Therefore

$E_2 - E_1 = -3.4 - (-13.6) = 13.6 - 3.4 = 10.2$ eV

21.9 Ionisation Energy or Ionisation Potential

The energy required to remove an electron from the outermost orbit of a gaseous atom in ground state is called ionisation energy or ionisation potential. Its value is positive.

Ex.36 The maximum energy absorbed by hydrogen atom in its ground state will be

[1] 13.6 eV [2] 3.4 eV [3] 10.2 eV [4] 0 eV **Ans. [1]**

Sol. $E_\infty - E_1$

$0 - (-13.6) = 13.6$ eV

Ex.37 The energy required in the process $He^{+2} \rightarrow He^{+3}$ will be

[1] 0 eV [2] + 13.6 eV [3] + 3.4 eV [4] + 1.5 eV **Ans. [1]**

Sol. He^{+2} does not have any electron, therefore the ionisation energy will be 0.

21.10 Separation Energy

The energy required to separate an electron from any excitation state of an atom is known as separation energy. For example, the first separation energy, i.e. the energy required to remove an electron from the first excited state in hydrogen is + 3.4 eV.

22. SPECTRAL EVIDENCE FOR QUANTIZATION IN BOHR'S THEORY

(a) When an electron undergoes transition from lower to higher orbit, there is absorption of energy and the spectrum obtained thereby is called absorption spectrum.

(b) When an electron undergoes transition from higher to lower orbit, there is emission of energy and the spectrum obtained thereby is called emission spectrum.

(c) A hydrogen atom has only one electron, yet a very large number of lines are visible in its spectrum.

(d) The wave number of spectrum can be find out using the following expression.

$$\bar{\nu} = \frac{1}{\lambda} = R \times Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where $\frac{1}{\lambda}$ is wave number

R = Rydberg constant,

n_1 = Number of lower energy level

n_2 = Number of higher energy level

Calculation of formula

$En_1 = -\frac{1}{2}\frac{Ze^2}{r_1}$ (r_1 = radius of the first orbit)

$En_2 = -\frac{1}{2}\frac{Ze^2}{r_2}$ (r_2 = radius of the second orbit)

$En_1 - En_2 = -\frac{1}{2}Ze^2\left(\frac{1}{r_1}-\frac{1}{r_2}\right)$

According to Bohr hypothesis

$En_1 - En_2 = h\nu$

$En_1 - En_2 = -h\nu$

Therefore $h\nu = \frac{2\pi^2 mZ^2e^4}{h^2}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$

$$\nu = \frac{2\pi^2 mZ^2e^4}{h^3}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

Here $\frac{2\pi^2 mZ^2e^4}{ch^3}$ is a constant, because for hydrogen atom Z = 1

Thus $R = \frac{2\pi^2 me^4}{ch^3}$ Value of R = 109678 cm^{-1}

If calculation, this value is 109700 cm^{-1}.

Formula = $\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$

Ex.38 What should be the value of wave number of emitted radiation with respect to R, when the electron present in hydrogen atom jumps from M orbit to K orbit ?

[1] $R \times \frac{8}{9}$ [2] $R \times \frac{5}{8}$ [3] $R \times \frac{3}{4}$ [4] $R \times \frac{5}{16}$ **Ans. [1]**

Sol. The electron jumps from M orbit (n = 3) to K orbit (n = 1). Therefore

$$\bar{\nu} = R\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$= R\left(\frac{1}{1}-\frac{1}{3^2}\right) = R\left(\frac{1}{1}-\frac{1}{9}\right) = R\left(\frac{9-1}{9}\right) = R \times \frac{8}{9}$$

Ex.39 What should be the energy of a photon whose wavelength is 4000 Å ?

[1] 4.06×10^{-19} joule [2] 4.96×10^{-19} joule [3] 3.0×10^{-12} joule [4] 2.4×10^{-19} joule

Ans. [2]

Sol. λ = 4000 Å i.e. 4000×10^{-8} cm 4×10^{-7} meter

$$E = h\nu = \frac{hc}{\lambda} = \frac{6.62\times10^{-34}\times3\times10^{8}}{4\times10^{-7}} = 4.96 \times 10^{-19} \text{ joule}$$

23. EMISSION SPECTRUM AND ABSORPTION SPECTRUM

When a beam of white light passes through a slit or an aperture and then falls on a prism, it gets spilt into many coloured bands. The image of colours so obtained is known as a spectrum. A spectrum is of mainly three types viz.

(i) Emission spectrum

(ii) absorption spectrum

and (iii) molecular spectrum

23.1 Emission Spectrum

When energy is provided to any substance, it starts emitting radiations. These radiations are passed through a spectroscope, they get split up into spectral lines producing emission spectrum. Normally a substance can be excited by any of the following ways.

(a) By heating the substance at high temperature

(b) By passing electric current through a discharge tube having gaseous substance at very low pressure.

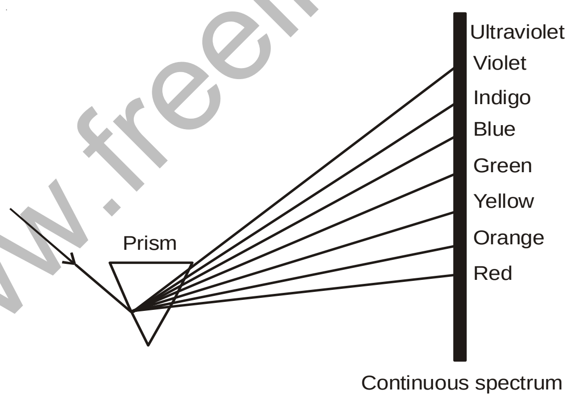
(c) By passing electric discharge through a metallic filament.

Emission spectra are of the following two types.

(i) Continuous spectrum and (ii) line spectrum or atomic spectrum

23.1.1 Continuous Spectrum

When sunlight or a glowing heat fluorescent substance like tungsten wire present in an electric bulb, is analysed with the help of a spectroscope, the spectrum obtained on a screen is observed as divided into bands of seven colours, which are in a continuous sequence. Such a spectrum is called a continuous spectrum.



Continuous spectrum of white light

23.1.2 Line spectrum or Atomic spectrum

When atoms of a substance is excited, it emits radiations. These radiations are analyzed with the help of a spectroscope, then many fine bright lines of specific colours in a sequence are seen in the spectrum, which is not continuous, i.e. there is dark zone in between any two lines. Such a spectrum is called a line spectrum or atomic spectrum. For example, neon single lamp, sodium vapour lamp, mercury vapour lamp, etc. emit light of different colours and they give specific line spectra.

23.2 Absorption Spectrum

When white light emitted by glowing heat fluorescent substance is passed through another substance lime sodium substance. This results in appearance of some black lines in the spectrum. These are present at those places where the line spectrum of the substance i.e. sodium vapour is formed. The spectrum so formed is known as absorption spectrum.

23.3 Molecular Spectrum

Molecular spectrum is given by molecules and it is also known as band spectrum. Three types of energy transitions are found in molecules. These are as follows.

(i) electronic transitions, (ii) vibrational transitions and (iii) rotational transitions.

Therefore, bands are obtained in the spectrum, which are actually groups of lines.

24. HYDROGEN SPECTRUM

Hydrogen atom gives line spectrum. When hydrogen gas is filled at low pressure in a discharge tube and electric discharge is passed through it, a pink coloured is produced in the visible region due to the formation of hydrogen atoms. On studying this light with the help of a spectroscope, series of lines of various wavelengths are obtained in the spectrum.

The frequency of spectral lines in the form of wave number can be calculated with the help of the following expression.

$$\bar{\nu} \text{ or } \frac{1}{\lambda} = R \times \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

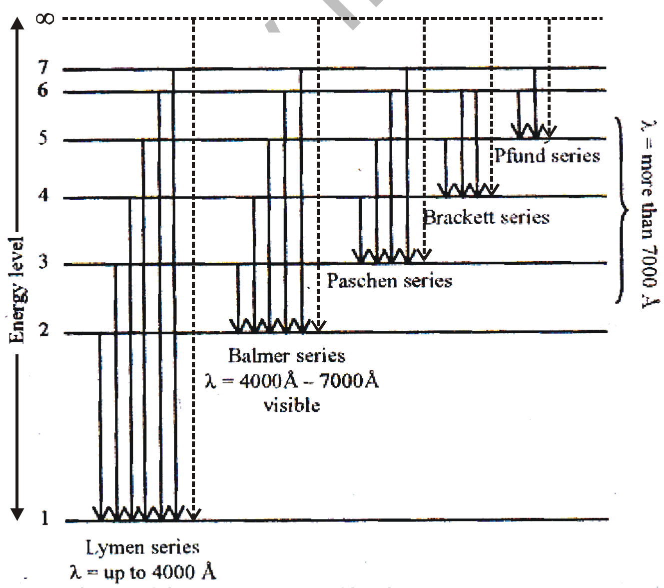
24.1 Series of Lines in Hydrogen Spectrum

24.1.1 Lymen Series

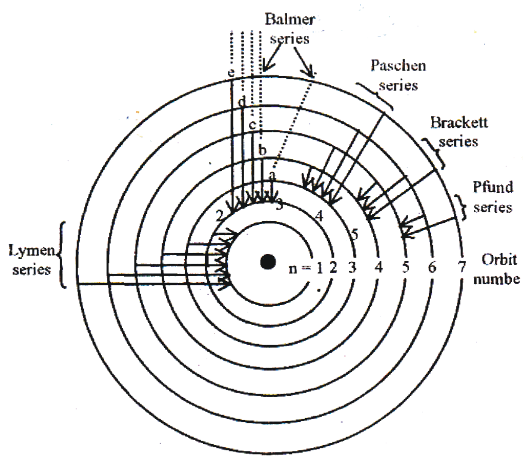
When an electron undergoes transition from a higher energy level (n_2), e.g. 2, 3, 4, 5, ... ∞ to ground state or lower energy level, the spectrum is said to belong to Lymen series. For this, n_1 = 1 and n_2 = 2, 3, 4, 5, 6, 7, 8 .... ∞.

24.1.2 Balmer Series

When an electron undergoes transition from a higher energy level (n_2), e.g. 3, 4, 5, 6, 7, 8 .... ∞ to the second energy level n_1 = 2, the spectrum is said to belong to Balmer series.



Regions of line spectrum of hydrogen atom



Explanation of Balmer series (line spectrum) on the basis of Bohr model

24.1.3 Paschen Series

When an electron falls from a higher energy level to third orbit (n = 3). It gives a spectrum that is associated with Paschen series. For this $n_1 = 3$ and $n_2 = 4, 5, 6, 7, 8 .... \infty$.

24.1.4 Brackett Series

When an electron falls from a higher energy level to the fourth orbit (n = 4), the spectrum obtained is associated with Brackett series. For this $n_1 = 4$ and $n_2 = 5, 6, 7, 8 ..... \infty$.

24.1.5 Pfund Series

When an electron falls from a higher energy level to the fifth orbit (n = 5), the spectrum obtained is associated with Pfund series. For this $n_1 = 5$ and $n_2 = 6, 7, 8, 9, 10 .... \infty$.

24.1.6 Humphry Series

When an electron falls from a higher energy level to the sixth orbit (n = 6), Humphrey series of the spectrum is obtained. For this $n_1 = 6$ and $n_2 = 7, 8, 9, 10, 11 .... \infty$.

S.No.	Series of lines	n_1	n_2	Spectral region	Wavelength
1.	Lymen series	1	2, 3, 4, 5 ...... ∞	Ultraviolet	< 4000Å
2.	Balmer series	2	3, 4, 5, 6 ...... ∞	Visible	4000Å to 7000Å
3.	Paschen series	3	4, 5, 6, 7 ...... ∞	Near infrared	> 7000Å
4.	Brackett series	4	5, 6, 7, 8 ..... ∞	Infrared	> 7000Å
5.	Pfund series	5	6, 7, 8, 9 ..... ∞	Far infrared	> 7000Å
6.	Humphrey series	6	7, 8, 9, 10 .... ∞	Far infrared	> 7000Å

For the given value of n (principal quantum number), the total number of spectral lines can be calculate by the by the expression $\frac{n(n-1)}{2}$.

Ex.40 How many emission spectral lines in all should be visible, if an electron is present in the third orbit of hydrogen atom ?

[1] 6 [2] 3 [3] 5 [4] 15 **Ans. [2]**

Sol. The expression of maximum number is $\frac{n(n-1)}{2} = \frac{3(3-1)}{2} = \frac{6}{2} = 3$

Ex.41 Which of the following should be the expression for the last line of Paschen series ?

[1] $\frac{1}{\lambda}=R\left(\frac{1}{9}-\frac{1}{\infty^2}\right)$ [2] $\frac{1}{\lambda}=R\left(\frac{1}{4}-\frac{1}{9}\right)$ [3] $\frac{1}{\lambda}=R\left(\frac{1}{9}-\frac{1}{16}\right)$ [4] $\frac{1}{\lambda}=R\left(\frac{1}{16}-\frac{1}{\infty}\right)$ **Ans. [1]**

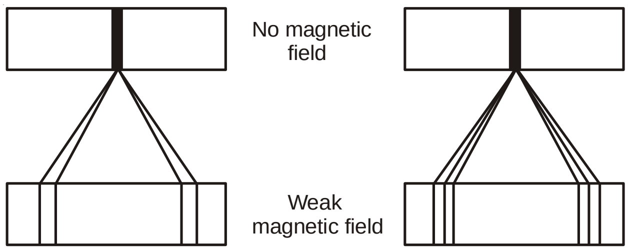
Sol. $\bar{\nu} = \frac{1}{\lambda} = R\left(\frac{1}{9}-\frac{1}{\infty}\right)$

25. FAILURES OF BOHR'S ATOMIC MODEL

(a) Bohr model cannot explain the elements having more than one electron. Only one-electron species, like hydrogen atom, He^{+1} ion, Li^{+2} ion, Be^{+3} ion, etc. can be explained with the help of Bohr model.

(b) Bohr model can explain only circular orbits in the atom and not the elliptical ones.

(c) Bohr model cannot explain splitting of spectral lines into finer lines in a magnetic field, which is known as **Zeeman effect**.



Representation of Zeeman effect

(d) Bohr model fails to explain the splitting of spectral lines into finer lines in an electric field, which is known as **Stark effect**.

(e) Bohr model fails to explain Hiesenberg uncertainty principle and it cannot be applied for giving any basis to classification of elements and periodicity in their properties.

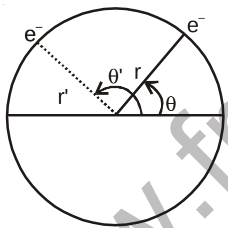
(f) Bohr model cannot be used for explaining finer structure of spectrum and calculating intensity of spectral lines.

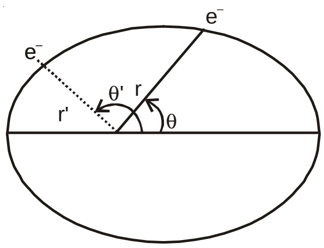
26. SOMMERFELD'S EXPANSION OF BOHR'S MODEL

(a) The aforesaid discovery proved that each principal quantum number (n) is composed of many suborbits.

(b) Sommerfeld suggested that electrons moves around the nucleus not only in circular orbits but also in elliptical orbits.

(c) When an electron travels in an elliptical orbit, its distance (r) from the nucleus and its angle of rotation both will change.





(d) In circular orbit, the distance r remains constant but angle of rotation will change. In elliptical orbit, the nucleus is regarded as situated at the focal point.

(e) A circular orbit is a particular situation of an elliptical orbit, in which the lengths of major axis is equal to that of the minor axis.

(f) In elliptical orbits, the orbital angular momentum is a sum of the following two vector number.

(i) towards the radius, which is called radial component P_r, and

(ii) in the perpendicular direction to radius, which is called azimuthal component P_k.

(g) The above two momenta are separately quantized, i.e. both are multiple of $\frac{h}{2\pi}$.

(h) Sommerfeld suggested that Bohr quantum number n is a sum of two quantum numbers, of which one is radial quantum number n_r and the other is azimuthal quantum number K, i.e.

$$n = n_r + K$$

27. THE WAVE THEORY OF LIGHT

Light, X-rays and radiation produced by a radioactive substance are some of the examples of radiation energy. In 1856 Clark Maxwell showed that energy of radiation is of wave nature, i.e. the energy is emitted in the form of a wave. Therefore, he called the emitted energy as electromagnetic wave or electromagnetic radiation. Since energy is a sort of wave, it is explained as wave motion. Following are the salient features of this wave motion.

[1] Wavelength (λ)

[2] Period (T)

[3] Fequcney (ν)

[4] Amplitude (A)

(5) Wave velocity (c or v)

The aforesaid properties of a wave have the following relationship

$$\nu = \frac{1}{T} \text{ and } c = \frac{\lambda}{T} \text{ or } c = \nu\lambda$$

27.1 Wave Length

The distance between any two successive crests (or troughs) is known as wavelength. This is expressed as λ (Lambda). The range of the wavelength associated with spectrum line is 10^8 to 10^6 cm. Its common units are as follows. Angstrom (Å).

27.2 Frequency

The number of vibrations produced in a unit time is called frequency. Here, the time is taken in seconds. The number of wavelengths passing forward in one second from a fixed point is called frequency.

27.3 Velocity of Light

The distance traveled by a light wave in a unit time (second) is called the velocity of that wave. It is represented by c and its unit is normally cm/second or m/second. Its value is definite. For example, for a light wave, the velocity c = 3×10^8 m/second or 3×10^{10} cm/second.

27.4 Amplitude

The maximum deviation of a wave from its equilibrium point is known as its amplitude.

27.5 Wave Number

The reciprocal of wavelength is called wave number. It is represented by $\bar{\upsilon}$.

$$\bar{\upsilon} = \frac{1}{\lambda}$$

Therefore, the unit of wave number is cm^{-1} or m^{-1}

$$\because \; c = \upsilon\lambda \text{ or } \lambda = \frac{c}{\upsilon} \text{ or } \upsilon = \frac{c}{\lambda} \text{ or } \upsilon = c\bar{\upsilon} \text{ or } \bar{\upsilon} = \frac{\upsilon}{c}$$

Ex.42 What should be the wavelength of an ultraviolet wave, if its frequency is 12×10^{16} cycles per seond and c = 3×10^8 m/second ?

[1] 25 Å [2] 2.5 Å [3] 0.25 Å [4] 0.025 Å **Ans.[1]**

Sol. $\because c = \nu\lambda$

Therefore $\lambda = \frac{c}{\nu} = \frac{3\times10^8}{12\times10^{16}} = 0.25 \times 10^{-8}$ m

or $\lambda = 2.5 \times 10^{-9}$ m or 25×10^{-10} m

λ = 25 Å

28. PLANCK'S QUANTUM THEORY

If a substance emits or absorbs energy, it does not do so continuously but does but does it in the form of discrete series of small packet or bundles, called quanta. This energy could be any of the quantum numbers 1, 2, 3, 4, 5 .... n but not in the form of fractional quantum number.

$\because \nu = \frac{c}{\lambda}$ Therefore $E = h \times \frac{c}{\lambda}$

29. THE DUEL NATURE OF MATTER (THE WAVE NATURE OF ELECTRON)

(a) In 1924, a French physicist, Luis De Breoglie suggested that if the nature of light is both that of a particle and of a wave, then this dual behaviour should be true for the matter also.

(b) The wave nature of light rays and X-rays is provided on the basis of their interference and diffraction and, many facts related to radiations can only be explained when the beam of light rays is regarded as composed of energy corpuscles or photons whose velocity is 3×10^{10} cm/second.

(c) According to ***De Broglie***, the wavelength λ of an electron is inversely proportional to its momentum p.

$\lambda \propto \frac{1}{p}$ or $\lambda = \frac{h}{p}$ Here h = Planck's constant

p = Momentum of electron

$\therefore$ Momentum (p) = Mass (m) × Velocity (c)

Therefore $\lambda = \frac{h}{mc}$ This is called De-Broglie equation

(d) The above relation can be confirmed as follows by using Einstein's equation, Planck's quantum theory and wave theory of light.

But according to Einstein's equation

$$E = mc^2 = h \times \frac{c}{\lambda} \text{ or } mc = \frac{h}{\lambda} \text{ or } p = \frac{h}{\lambda} \text{ or } \lambda = \frac{h}{p}$$

30. BOHR'S THEORY AND DE BROGLIE CONCEPT

(a) According to De Broglie, the nature of an electron moving around the nucleus is like a wave that flows in circular orbits around the nucleus.

(b) If an electron is regarded as a wave, the quantum condition as given by Bohr in his theory is readily fulfilled.

(c) If the radius of a circular orbit is r, its circumference will be $2\pi r$.

(d) We known that according to Bohr theory, $mvr = \frac{nh}{2\pi}$

or $2\pi r = \frac{nh}{mv}$ ($\because$ mv = p momentum) or $2\pi r = \frac{nh}{p}$ ($\because \frac{h}{p} = \lambda$ De Broglie equation)

(e) $\because 2\pi r = \frac{nh}{mv}$

or $mvr = \frac{nh}{2\pi}$ $\because$ mvr = Angular momentum

Thus mvr = Angular momentum, which is a integral multiple of $\frac{h}{2\pi}$

(f) It is clear from the above description that according to De Broglie there is similarity between wave theory and Bohr theory.

Ex.43 What should be the mass of the photon of sodium light if its wavelength is 5894 Å, the velocity of light is 3×10^8 metre/second and the value of h is 6.6252×10^{-34} Kg m^2/sec ?

[1] 3.746×10^{-26} [2] 3.746×10^{-30} [3] 3.746×10^{-34} [4] 3.746×10^{-36} **Ans. [4]**

Sol. $\lambda = \frac{h}{m \times c}$ or $\frac{h}{c\lambda}$

($\because \lambda$ = 5894 Å or 5894×10^{-10} m)

$m = \frac{6.652 \times 10^{-34}}{3 \times 10^8 \times 5894 \times 10^{-10}}$ or $\frac{6.652}{17682} \times 10^{-32}$

$= 3.746 \times 10^{-36}$ Kg

31. QUANTUM MECHANICAL THEORY OF ATOM

(a) The dual nature (particle and wave) of electron led to the use of a new system of mechanics called quantum mechanics. This system was first put forward by an Austrian physicist E. Schrodinger and a German physicist W. Heisenberg.

(b) The two fundamental principles of quantum mechanics are given below :

(i) Heisenberg's uncertainty principal and

(ii) Schrodinger's wave equation

31.1 Heisenberg's Uncertainty Principle

(a) According to this principle, it is impossible to experimentally determine together both exact position and actual momentum of a minute particle like an electron.

(b) This principal can be depicted mathematically as follows.

$$\Delta x \times \Delta p \geq \frac{h}{4\pi} \quad \text{or} \quad \Delta x \times m \times \Delta v \geq \frac{h}{4\pi}$$

Here Δx is uncertainty of position,

Δp is uncertainty of momentum and

h is Planck's constant

Ex.44 What should be the uncertainty in position if uncertainty in momentum is 1×10^{-2} gm cm/sec and value of h is 6.6253×10^{-34} Js ?

[1] 1.054×10^{-22}m [2] 1.054×10^{-25}m [3] 1.054×10^{-27} m [4] 1.054×10^{-32} m **Ans. [3]**

Sol. Given that

$\Delta p = 1 \times 10^{-2}$ gm cm/sec. $= 1 \times 10^{-7}$ Kg m/sec.

$h = 6.6252 \times 10^{-34}$ Js

$$\Delta x \times \Delta p = \frac{h}{2\pi} \quad \therefore \Delta x = \frac{h}{2\pi \times \Delta p}$$

or $$\Delta x = \frac{6.6252 \times 10^{-34}}{2 \times 3.14 \times 10^{-7}} \text{ m} = 1.054 \times 10^{-27} \text{m}$$

31.2 Schrondinger's Wave Equation

Schrondinger regarded electron as having wave nature and put forward the following complex differential equation.

$$\nabla^2 \psi + \frac{8\pi^2 m}{h^2} (E - v)\, \phi = 0$$

$$\nabla^2 = \frac{d^2}{dx^2} + \frac{d^2}{dy^2} + \frac{d^2}{dz^2}$$

where m = Mass of electron, h = Planck constant,

E = Total energy of electron, v = Potential energy of electron,

ψ = Wave function, ∇ = Laplacian Operator

32. QUANTUM NUMBERS

(a) The position of any electron in any atom can be ascertained with the help of quantum numbers.

(b) In an atom, the shell consists of sub-shells and the sub-shell consists of orbital can accommodate only two electrons, which are in opposite spins.

32.1 Principal Quantum Number (n)

(a) Principal quantum number indicates the shell or energy level or orbit.

(b) An atoms has K, L, M, N, O, P, Q, etc. shells.

(c) Principal quantum number also gives information about the radius of size.

(d) Principal quantum number also gives information about the distance of an electron from the nucleus in an atom

(e) Principal quantum number also given information about the energy of an electron.

(f) Principal quantum number also gives information about the velocity of an electron.

(g) In any orbit, the number of orbitals is given by n^2 and number of electrons is given by $2n^2$. This is called Bohr-Bury rule.

Ex.45 Which of the following is the principal quantum number for the last electron of $_{11}Na$?

[1] 3 [2] 2 [3] 4 [4] 1 **Ans. [1]**

Sol. $_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$

$n = 3$

Ex.46 Which of the following should have greater size ?

[1] 1s [2] 2s [3] 3s [4] 4s **Ans. [4]**

Sol. n = 4 for 4s

32.2 Azimuthal Quantum Number (*l*)

(a) Azimuthal quantum number gives information that a particular electron belongs to which sub-shell.

(b) In an atom the shells consist or sub-shells, which are indicated as *s*, *p*, *d* and *f*.

(c) Azimuthal quantum number determines the shape of an orbital.

(d) The value of n starts from 1, while that of l starts from 0. Therefore, the maximum value of l is n – 1.

(e) The values of n and l can never be equal.

Sub shell	s	p	d	f
l	0	1	2	3

(f) The number of orbitals in any sub orbit is determined by the expression $2l + 1$ and the number of electrons is determined by the expression $2(2l + 1)$.

(g) $l = 0$ $\rightarrow$ s Sub-shell $\rightarrow$ Spherical

$l = 1$ $\rightarrow$ p Sub-shell $\rightarrow$ Dumb-bell

$l = 2$ $\rightarrow$ d Sub-shell $\rightarrow$ Double dumb-bell

$l = 3$ $\rightarrow$ f Sub-shell $\rightarrow$ Complex

(h) The order of energy of various sub-shells present in any shell is s < p < d < f < g .... and so on.

(i) The value of orbital angular momentum, μ_i, of an electron can be determined with the help of azimuthal quantum number

$$\mu_i = \sqrt{l(l+1)} \times \frac{h}{2\pi}$$

Here ℓ = Azimuthal quantum number and h = Planck's constant

Ex.47 Which of the following should be the possible sub-shells, for n + ℓ = 7 ?

[1] 7s, 6p, 5d, 4f [2] 4f, 5p, 6s, 4d [3] 7s, 6p, 5d, 6d [4] 4s, 5d, 6p, 7s **Ans. [1]**

Sol. n + l = 7

7 + 0 = 7s ; 6 + 1 = 6p ; 5 + 2 = 5d ; 4 + 3 = 4f

Ex.48 What should be the maximum number of electron in the possible sub-shells, for n + l = 4 ?

[1] 8 [2] 6 [3] 12 [4] 16 **Ans. [1]**

Sol. n + l = 4 Maximum number of electrons

4 + 0 = 4s $\rightarrow$ 2

3 + 1 = 3p $\rightarrow$ 6

8

Ex.49 The sub-shell 2d is not possible because

[1] n $\neq$ l [2] l > n [3] n < l [4] None of these **Ans. [1]**

Sol. For sub-shell 2d, n = 2 and l = 2 and the values of n and l can never be equal.

Ex.50 What should be the maximum number of elements, if the elements above n = 4 do not exist in nature ?

[1] 40 [2] 60 [3] 44 [4] 108 **Ans. [2]**

Sol. Since, n = 1, 2, 3 and 4, therefore

$\frac{1s}{2}, \frac{2s, 2p}{8}, \frac{3s, 3p, 3d}{18}, \frac{4s, 4p, 4d, 4f}{32}$ Thus, total number of existent elements = 2 + 8 + 18 + 32 = 60

Ex.51 Which of the following orbitals should be nearest to the nucleus ?

[1] 5s [2] 6p [3] 3d [4] 4d **Ans. [3]**

Sol. n = 3 will be nearest to the nucleus.

32.3 Magnetic Quantum Number (m)

(a) Magnetic quantum number gives information about an orbital. It is depicted by the symbol m.

(b) Magnetic quantum number gives information about orientation of orbitals.

(c) The value of m ranges from $-\ell$ to $+\ell$.

(d) The total number of orbitals present in a sublevel is equal to the total values of magnetic quantum number. This can be find out by the following expression.

$m = 2l = 1$

where m is total value of magnetic quantum number and l is the value of azimuthal quantum number.

(i) For s sub-shell, l = 0. Thus, m = 2 × 0 + 1 = 1 and therefore s sub-shell consists of only one orbital called s orbital.

(ii) For p sub-shell, l = 1. Thus, m = 2 × 1 + 1 = 3 and therefore p sub-shell consists of three orbitals called p_x, p_y and p_z orbitals.

(iii) For d sub-shell, l = 2. Thus, m = 2 × 2 + 1 = 5 and therefore d sub-shell consists of five orbitals called d_{xy}, d_{yz}, d_{z^2}, d_{xz} and $d_{x^2-y^2}$ orbitals.

(i) For s sublevel, l = 0. Thus, for s orbital, the value of m is 0.

s

m = 0

(ii) For p sub-level, l = 1. Thus, the values of m for p orbitals are as follows.

p_x/p_y	p_z	p_z/p_y
–1	0	+1

(iii) For d sub-level, l = 2. Thus, the values of m for d orbitals are as follows.

d_{xy}	d_{yz}	d_{z^2}	d_{xz}	$d_{x^2-y^2}$
–2	–1	0	+1	+2

(iv) For f sub-level, l = 3. Thus, the values of m for f orbitals are as follows.

–3	–2	–1	0	+1	+2	+3

(e) The total number of orbitals present in an energy level is determined by the formula n^2 where n is principal quantum number.

Ex.52 What should be the total numbers of orbitals and electrons for m = 0, if there are 30 protons in an atom ?

[1] 7 orbitals, 14 electrons [2] 6 orbitals, 12 electrons

[3] 5 orbitals, 10 electrons [4] 3 orbitals, 6 electrons **Ans. [1]**

Sol. The configuration of the atom of atomic number 30 is $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2$. This will have 7 orbitals of m = 0.

Ex.53 The orbital having n = 6, l = 2 and m = 0 will be designated as

[1] $6\,d_{z^2}$ [2] $6\,d_{x^2-y^2}$ [3] $6d_{xy}$ [4] $6p_z$ **Ans. [1]**

Sol. For 6th of P energy level, l = 2 is for d sub-level, and m = 0 for d_{z^2} orbital

Ex.54 The orbital having n = 2, l = 1 and m = 0 is designated as

[1] $2p_z$ [2] $2p_x$ [3] $2p_y$ [4] $3\,d_{z^2}$ **Ans. [1]**

Sol. In the second or L energy level (n = 2) , l = 1 for p orbital, m = 0 for z axis, Hence, the orbital will be designated as $2p_z$.

32.4 Spin Quantum Number (s)

(a) Spin quantum number gives information about the spin of an electron.

(b) The value of s is 1/2 which depicts the direction of spin of the electron.

(c) If the electron spins in clockwise direction, s is denoted by $+\frac{1}{2}$ or a sign [↑]. Anticlockwise spin of the electron is denoted by $s = -\frac{1}{2}$ or [↓].

(d) One orbital can accommodate only two electrons, with opposite spins.

(e) One orbital can accommodate only two electrons, with opposite spins.

(f) The angular momentum of an electron is not only due its motion around the nucleus in an energy level but also due to its rotation along its own axis. The angular momentum that arises due to rotation of an electron along its axis, is called spin angular momentum and is depicted by the symbol μs. The value of μs can be found out with the help of the following expression.

$\mu s = \sqrt{s(s+1)} \times \frac{h}{2\pi}$ where s is spin quantum number. In this expression the value of s is always taken as $\frac{1}{2}$ and not $-\frac{1}{2}$.

Ex.55 If x is the number of electron in an atom, the configuration should be expressed as :

[1] l_x [2] nl^x [3] nm^x [4] None of these **Ans. [2]**

Sol. The electronic configuration of an atom is expressed by first writing principal quantum number (n), followed by azimuthal quantum number (l) and then writing number of electrons (x) as superscript.

Ex.56 What should be the atomic number of an element, if the quantum numbers of the highest energy electron of the element in ground state are n = 4, l = 1, m = – 1 s = $+\frac{1}{2}$?

[1] 31 [2] 35 [3] 30 [4] 32 **Ans. [1]**

Sol. The electronic configuration of the element will be $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^1$. Thus, the total number of electrons is 31 and hence the atomic number will be 31.

Ex.57 The orbital having m = – 2 should not be present in the following sub-shell

[1] d [2] f [3] g [4] p **Ans. [4]**

Sol. For p sub-shell, m = – 1, 0, + 1. Therefore, m = – 2 orbital will not be present in p sub-shell.

Ex.58 What should be the value of spin quantum number of the last electron for d^9 configuration ?

[1] 0 [2] $-\frac{1}{2}$ [3] $\frac{1}{2}$ [4] 1 **Ans. [2]**

Sol. The value of spin quantum number (s) can be $+\frac{1}{2}$ or $-\frac{1}{2}$, because an electron can rotate along its axis either in clockwise or in anticlockwise direction. But one quantum number depicts one electrons and thus its value will be $-\frac{1}{2}$ for d^9 configuration.

Ex.59 The all energy levels are called excited states when the value of principal quantum number is :

[1] n = 1 [2] n > 1 [3] n < 1 [4] n > – 1 **Ans. [2]**

Sol. All the energy states in which n is greater than 1 are called excited states.

33. AUFBAU PRINCIPLE

Aufbau is a German word that means to build up. Therefore, electrons are filled up in accordance with this principle.

(a) Pauli's exclusion principle should be followed during filling up of electrons, i.e. no two electrons should have same set of four quantum numbers. This means that maximum number of electrons to be filled in various sub-shells are 2 in s, 6 in p, 10 in d and 14 in f.

(b) Hund's rule should be followed during filling up of electrons i.e. the electrons are to be filled in the degenerate orbitals first in unpaired state.

(c) The electrons are filled in a sub-shell according in n + l rule.

33.1 Pauli's Exclusion Principle

(i) According to Pauli exclusion principle, any two electron cannot have same set of four quantum numbers.

For example :

(a) $6s^1$ and $6s^2$

$6s^1$	and	$6s^2$
n = 6		n = 6
l = 0		l = 0
m = 0		m = 0
$s = +\frac{1}{2}$		$s = -\frac{1}{2}$

(b) $4p^2$ and $4p^5$

$4p^2$	and	$4p^5$
n = 4		n = 4
l = 1		l = 1
m = 0		m = 0
$s = +\frac{1}{2}$		$s = -\frac{1}{2}$

In the above illustrations, the respectively values of n, l and m are same but that of s is different.

(ii) Pauli exclusion principle can be stated in other words as that **"only two electrons can be accommodated in the same orbital only when their spin quantum number is different".**

(iii) If the third electron enters in an orbital, the set of four quantum numbers becomes same for any two electron.

(iv) According to this rule, for any two electrons, a set of maximum three quantum numbers can be same, but the fourth has to be different. For example, two electrons can have same (n, l and m) or (l, m or s) or (n, m or s)

Example

$1s^1$	and	$1s^2$
n = 1		n = 1
l = 0		l = 0
m = 0		m = 0
$s = +\frac{1}{2}$		$s = -\frac{1}{2}$

(v) This rule does not apply for hydrogen atom because it contains only one electron.

Ex.60 Pauli exclusion principle applies to

[1] H [2] H^+ [3] H^- [4] None of the above **Ans. [3]**

Sol. Since, H has one electron and H^+ has no electron, therefore Paulie principal does not apply to them. However, H^- has two electrons, hence this principle applies on it.

Ex.61 Which of the following statements is true ?

[1] One orbit can accommodate a maximum of two electrons

[2] One sub-shell can accommodate a maximum of two electrons

[3] One orbital can accommodate a maximum of two electrons

[4] None of the above **Ans. [3]**

Sol. It is an orbital that can accommodate a maximum of two electrons having opposite spins

Ex.62 Which of the following is not according to Pauli exclusion principle ?

[1] $\boxed{\uparrow\uparrow}\boxed{}\boxed{}$ [2] $\boxed{\uparrow\downarrow\uparrow}\boxed{}\boxed{}$ [3] $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$ [4] 1 and 2 both **Ans. [4]**

Sol. The set of four quantum numbers are not same for the three electrons in answer 3. In answer 1 both of the electrons have same set of quantum numbers, while in answer 2 the first and third electrons have same set of quantum numbers.

Ex.63 Supposing that Pauli exclusion principle is not correct, if one orbital can accommodate three electrons, what should be the respective atomic number of the second member of alkali metal family and the first member of halogen family

[1] 16, 14 [2] 11, 9 [3] 16, 9 [4] 34, 17 **Ans. [1]**

Sol. (a) Sodium is the second member of alkali metal family

$Na^{11} = 1s^2, 2s^2, 2p^6, 3s^1$

We know that the inner orbitals of sodium are fully filled and the outer most orbit has one electron. If inner orbitals can accommodate three electrons each, the configuration will be as follows.

$1s^3, 2s^3, 2p^9, 3s^1$

Therefore, three will be 16 electrons in all. Hence the atomic number will be 16.

(b) The first member of halogen family is fluorine, F^9 whose configuration is $1s^2, 2s^2, 2p^5$

Halogen has one electron less than the next inert of noble gas. If inner orbitals can accommodate three electron each, the configuration will be as fallows.

$1s^3, 2s^3, 2p^8$

Therefore, total number of electrons will be 14 and thus the atomic number will also be 14.

Ex.64 Supposing that Pauli exclusion principle is nonexistent, which of the following should be the most unacceptable configuration of Li in ground state ?

[1] $1s^2, 2s^1$ [2] $1s^3$ [3] $1s^1, 2s^2$ [4] $1s^1, 2s^1, 2p^1$ **Ans. [4]**

Sol. As a matter of fact, the configuration given in 2, 3 and 4 are wrong, but configuration given in 4 is most unacceptable because there is one electron in each of the three orbitals and according to Paulis exclusion principle maximum two electrons can be occupier in a orbital.

33.2 Hund's Rule of Maximum Multiplicity

(a) Degenerate orbitals

The orbitals having same energy are called degenerate orbitals.

(b) s sub-shell consists of only one orbital. Thus, it cannot have degenerate orbital.

(c) According to Hund's rule, the degenerate orbitals get filled by electrons having parallel spin one by one to give an unpaired state.

(d) According to this rule, the degenerate orbitals are filled in such a way that there is a maximum number of unpaired electrons. For example, C^6 can possibly have the following two configurations of $2s^2\ 2p^2$.

2p

[1] $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{}\boxed{}$

[2] $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{}$

(e) The following two conditions have to be fulfilled for Hund's rule.

[1] The orbitals should be degenerate

[2] The member of electrons and the degenerate orbitals should be more than one

(f) Hund's rule is not applicable for H, He, Li and Be, because electrons in them go to s sub-shell, which does not have any degenerate orbital.

(g) Hund's rule is not applicable for B^5 also, because there is only one electrons in p orbital. Therefore, this rule is applicable from C^6 onwards.

(h) Hund's rule is not important for elements belonging to groups IA, IIA and IIIA.

Ex.65 Which of the following should be correct according to Hund's rule ?

[1] $C^6 = 1s^2, 2s^2$ $\boxed{\uparrow\uparrow}\boxed{}\boxed{}$ [2] $O^8 = 1s^2, 2s^2$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{}$

[3] $N^7 = 1s^2, 2s^2,$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$ [4] $F^9 = 1s^2, 2s^2$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$ **Ans. [4]**

Sol. Configuration of C^6 should be $2p_x^1\ 2p_y^1$ instead of $2p_x^2$

Configuration of O^8 should be $2p_x^2\ 2p_y^1\ 2p_z^1$ instead of $2p_x^2\ 2p_y^2$

Configuration of N^7 should be $2p_x^1\ 2p_y^1\ 2p_z^1$ instead of $2p_x^2\ 2p_y^1$

Configuration of F^9 $2p_x^2\ 2p_y^2\ 2p_z^1$ is correct because two out of the three degenerate p orbtials are fully-filled, one is half-filled and there is no unfilled p orbital.

33.3 n + *l* Rule

(a) n + *l* Rule gives information about the energy of various sub-shells.

(b) According to this rule, the sub-shells having higher value of n + *l* have higher energy.

(c) The sub-shells having lower value of n + *l* have lower energy.

(d) It two sub-shells have same value of n + *l*, then that sub-shell will have higher energy which has higher value of n.

Ex.66 If the value of n + *l* = 7, then what should be the increasing order of energy of the possible sub-shells ?

[1] 4*f* < 5d < 6p < 7s [2] 7s < 6p < 5d < 4f [3] 7s > 6p < 5d < 4p [4] 4f < 5d < 6p > 7s **Ans. [1]**

Sol. n + *l* = 7

7 + 0 = 7s Order of energy

6 + 1 = 6p 4f < 5d < 6p < 7s

5 + 2 = 5d

4 + 3 = 4f

Ex.67 Which of the following sub-shells will be filled by the electron after complete filling up of the orbital of the third principal shell ?

[1] 4s [2] 4f [3] 4d [4] 4p **Ans. [4]**

Sol. The electron goes to 4p after filling up to 3d.

Ex.68 Which of the following should be the basis of entry of an electron in 4s orbital before 3d orbital ?

[1] Energy level diagram [2] Hund's rule [3] Pauli's principle [4] Shielding constant **Ans. [1]**

Sol. n + *l* of 4s = 4 + 0 = 4 and that of 3d is 3 + 2 = 5. Therefore, energy of 4s is lower than that of 3d.

Increasing order of energy

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p$$

The maximum number of electrons that can be accommodated in s orbital is 2, that in p orbital is 6, that in d orbital is 10 and that in f orbital is 14.

Exceptions to n + *l* Rule

There are mainly two exceptions of n + *l* rule.

(a) La^{57} – $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 5d^1$

(b) Ac^{89} – $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6s^2, 7s^2, 6d^1$

Specific Electronic Configuration

	Element	Atomic number	Expected configuration	Atcual configration
1.	Cr	24	$[Ar]^{18}\ 3d^4\ 4s^2$	$[Ar]^{18}\ 3d^5\ 4s^1$
2.	Cu	29	$[Ar]^{18}\ 3d^9\ 4s^2$	$[Ar]^{18}\ 3d^{10}\ 4s^1$
3.	Mo	42	$[Kr]^{36}\ 4d^4\ 5s^2$	$[Kr]^{36}\ 4d^5\ 5s^1$
4.	Pd	46	$[Kr]^{36}\ 4d^8\ 5s^2$	$[Kr]^{36}\ 4d^{10}\ 5s^0$
5.	Ag	47	$[Kr]^{36}\ 4d^9\ 5s^2$	$[Kr]^{36}\ 4d^{10}\ 5s^1$
6.	W	74	$[Xe]^{54}\ 4f^{14}\ 5d^4\ 6s^2$	$[Xe]^{54}\ 4f^{14}\ 5d^5\ 6s^1$
7.	Pt	78	$[Xe]^{54}\ 4f^{14}\ 5d^8\ 6s^2$	$[Xe]^{54}\ 4f^{14}\ 5d^9\ 6s^1$
8.	Au	79	$[Xe]^{54}\ 4f^{14}\ 5d^9\ 6s^2$	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^1$

Due to greater stability of half-filled and fully-filled orbitals, the configurations $d^5\ ns^1$ and $d^{10}\ ns^1$ are written in place of $d^4\ ns^2$ and $d^9\ ns^2$ respectively.

33.4 Stability of Half-filled and Fully-Filled orbitals

The stability of half-filled orbitals (p^3, d^5 and f^7) and fully-filled orbitals (p^6, d^{10} and f^{14}) is higher than that in other states. This is due the following reasons.

(a) When a sub-shell is half–filled or fully-filled, it means that the distribution of electrons is symmetrical in the orbitals of equal energy. Unsymmetrical distribution of electrons results in lower stability.

(b) The electrons present in orbtials of equal energy in an atom can interchange their position, in this process energy is released, resulting stable system. The possibility of interchange of positions is highest in half filled and fully-filled states. This provides greater stability to the system.

(c) The exchange energy for half-filled and fully-filled orbitals is maximum. As the number of electrons increases, electron start pairing resulting in spin coupling. The energy liberated in the process of coupling is called coupling energy.

(d) The spin of electrons in a fully-filled orbital are opposite to each other or antiparallel. The energy of the system decreases due to neutralization of opposite spins. So fully-filled orbitals are more stable.

34. MODE OF FILLING UP TO ELECTRONS

Writing the configuration of ions

First of all, the configuration of the atom is written. Then, appropriate number of electrons are deducted from the outermost shell for the configuration of the cation. Similarly, appropriate number of electrons are added to the outermost shell for the configuration of the anion.

Ex.69 Which of the following should be the atomic number of an atom if its electronic configuration is $(n-2)s^2, (n-1)\ s^a p^b\ ns^a\ p^2$ where n = 3, a = 2 and b = 6 ?

[1] 14 [2] 12 [3] 16 [4] 15 **Ans. [1]**

Sol. $(3-2)s^2\ (3-1)s^2\ p^6\ 3s^2\ 3p^2$; $1s^2, 2s^2\ 2p^6, 3s^2, 3p^2 = 14$

The atomic number of the atom is 14

Ex.70 Which of the following should be the number of electrons present in X^{+2} on the basis of electronic configuration, if the ion X^{-3} has 14 protons ?

[1] 12 [2] 14 [3] 16 [4] 18 **Ans. [1]**

Sol. X^{-3} has 14 protons, i.e. X also has 14 protons and therefore 14 electrons.

$X = 14 = 1s^2, 2s^2\ 2p^6, 3s^3\ 3p^2$

$X^{+2} = 1s^2, 2s^2\ 2p^6, 3s^2 = 12$ electrons

Ex.71 Which of the following should be the electronic configuration of an atom in its first excited state if that atom is isoelectronic with O_2 ?

[1] [Ne] $3s^2\ 3p^4$ [2] [Ne] $3s^2\ 3p^3\ 3d^1$ [3] [Ne] $3s^1\ 3p^5$ [4] None of the above **Ans. [2]**

Sol. 16 electrons = $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^4$ (Two unpaired electrons)

Excited state = [Ne] $3s^2, 3p^3, 3d^1$ (Four unpaired electrons)

Ex.72 Which of the following should be the electronic configuration of P in H_3PO_4 ?

[1] [Ne] [2] [Ne] $3s^2\ 3p^6$ [3] [Ne] $3s^2$ [4] [Ne] $3s^2\ 3p^1$ **Ans. [2]**

35. DIFFERENCE BETWEEN ORBIT AND ORBITAL

S.No.	Orbit	Orbital
1.	It is depicted by n.	It is depicted by m
2.	It has maximum electron capacity of $2n^2$	It has maximum electron capacity of 2 in accordance with Pauli's principle
3.	It is bigger in size	It is smaller in size
4.	Orbit consist of suborbits	Sub-orbit consists of orbitals
5.	The path of an electron around the nucleus is called an orbit	The space around the nucleus where probability of finding an electron is maximum, is called an orbital

36. ORBITAL

(a) The space around the nucleus where probability of finding an electron is maximum, is called an orbital.

(b) An electron cloud is negatively charged and the nucleus is positively charged. Therefore, the probability of finding an electron is maximum around the nucleus.

(c) The probability of finding an electron is an orbital is 95% to 98%.

36.1 s-Orbital

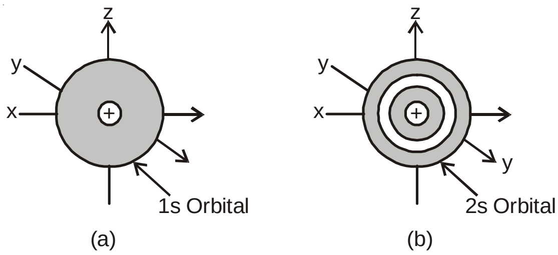
(a) Only one s-orbital is possible in an orbit because $l = 0$ and m = 0 for it.

(b) It is spherical in shape and thus the electron density is uniform in all directions.

(c) The size increases with increase in the value of n. There is vacant space between 1s orbital and 2s orbital, where the probability of finding electron is minimum, it is known as **nodal surface**.

(d) The nodal surface is missing inside 1s orbital because of its proximity with the nucleus.

(e) The number of nodel surfaces in an orbit is equal to (n – 1)



The shapes of s-orbital

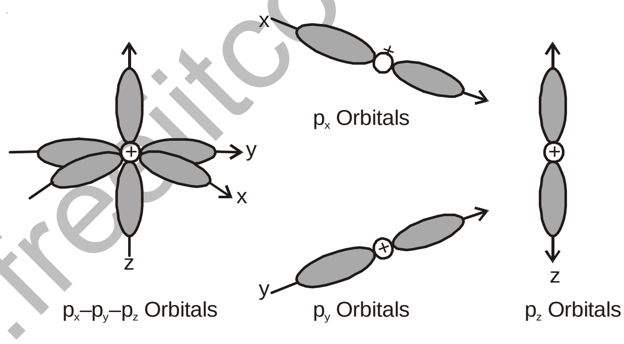
36.2 p-Orbital

(a) For p orbitals, l = 1 and m = – 1, 0, + 1. Thus, it can have three configurations, which are distributed in x, y and z axes. Therefore, there are three p-orbitals, which are dumbbell, shaped.

(b) Each p-orbital has two lobes and the probability of finding electron inside these two lobes is equal. The plane perpendicular to the axis two lobes and passing through the point where these two lobes join, is the nodal plane of p-orbital, because the probability of finding electron in this plane is negligible or minimum.

(c) The value of nodal planes for each of the p_x, p_y and p_z orbitals is same and these nodal planes are present in xy, yz and xz planes, respectively.

(d) The three p-orbitals of a particular orbit (p_x, p_y and p_z) have equal energy and therefore these are called degenerated orbitals.



Shape of p orbitals

36.3 d-Orbitals

(a) For d orbitals, l = 2 and m = – 2, – 1, 0, +1, + 2. Therefore, there are five orientations and thus five d-orbitals.

(b) Its shape is like a double dumbbell.

(c) The five orientations of d-orbitals are as follows :

(i) The double dumbbell of d_{xy} orbital are situated between x and y axes.

(ii) The double dumbbell of d_{yz} orbital are situated between y and z axes

(iii) The double dumbbell of d_{xz} orbital are situated between x and z axes.

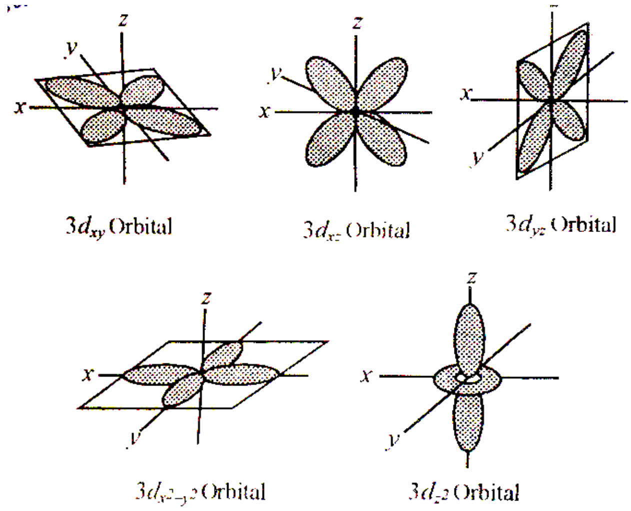
(iv) The double dumbbell of $d_{x^2-y^2}$ orbital are directed on x and y axes

(v) d_{z^2} orbital is composed of one dumbbell and one ring. The dumbbell is situated on z axis and the ring is present on its middle part.

The aforesaid five d orbtials can be classified into the following two categories.

(i) t_{2g} Orbitals (d_{xy}, d_{xz} and d_{yz}) – In these, the electron density is concentrated in-between the axes. These are also called grade orbitals.

(ii) e_g Orbitals ($d_{x^2-y^2}$ and d_{z^2}) – In these, the electron density is concentrated on the axes.



Shape of d Orbitals

36.4 f-orbitals

(i) They have complex shapes,

(ii) For these, l = 3 and m = – 3, –2, –1, 0, +1, +2, +3

(iii) These have seven orientations.

37. CONTRIBUTION OF SOME SCIENTISTS

[1] Atom

John Dalton in 1880 said that matter is composed of very minute and indivisible particles, called atoms.

[2] Electrons

(a) Discoverer – J.J. Thomoson

(b) Weight – $\frac{1}{1837}$ th of the weight of H atom

(c) Mass – 9.1×10^{-28} gram

(d) Amount of charge – (a) 4.8×10^{-10} e.s.u. (electrostatic unit) (b) 1.6012×10^{-19} coulomb

(e) Rest mass – Mass of electron = 9.1×10^{-28} grams

1 a.m.u. = 1.6×10^{-24} gm

$$\text{Rest mass} = \frac{9.1\times10^{-28}}{1.6\times10^{-24}} = 5.51 \times 10^{-4} \text{ a.m.u.}$$

The mass of an electron 9.1×10^{-28} gram is also called its rest mass.

(f) Discovery – With the help of cathode rays.

[3] Proton

(a) Discoverer – Goldstein

(b) Discovery – With the help of anode rays

(c) Mass – 1.6748×10^{-24} gram or 1.00757 a.m.u.

(d) Charge – Unit positive charge

[4] Neutron

(a) Discoverer – James Chadwick (1932)

(b) Charge – Zero (i.e. a neutral particle)

(c) Mass – 1.67×10^{-24} gram or 1.6×10^{-27} Kg

(d) Density – 10^{-12} Kg/cm^3

(5) Nucleus

(a) Discoverer – Rutherford

(b) Size of nucleus – 10^{-13} to 10^{-12} cm i.e. 10^{-15} to 10^{-14} metre

(c) Size of atoms – 10^{-8} cm

(d) Atomic radius – $10^5 \times$ Radius of the nucleus

(6) Positron

(a) Discoverer – C.D. Anderson (1932)

(b) Symbol – e^{+1}

(c) Charge – Unit positive charge

(d) Mass – Negligible (like electron)

(e) Stable particle

(7) Meson

(a) Predicted by — Yukawa

(b) Discoverer – Neddermeyer and Anderson

(c) Charge – Positive, negative or zero

(d) Mass – In between proton and electron

(e) Unstable particle

(8) Neutrino

(a) Discoverer – Pauling

(b) Charge – Zero

(c) Mass – Negligible (less than that of electron)

(d) Stable particle

(9) Antiproton

(a) Discoverer – Segre

(b) Charge – Unit negative charge

(c) Mass – Equal to that of proton

(d) Stable particle

(10) Volume of atom $= \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\, 10^{-24}$ cm

(11) Volume of nucleus $= \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\, 10^{-39}$ cm $= \frac{10^{-39}}{10^{-24}} = 10^{-15}$ cm

Thus, the nucleus of an atom occupies 10^{-15} part of an atom

(12) Some discoverers

(a) Positive rays – **Goldstein**

(b) Cathode rays – **William Crookes**

(c) Atomic number – **Mosley**

(d) e/m of proton – **J.J. Thomson**

(e) chanrge on electron – **Milliken**

(f) e/m of electron – **J.J. Thomson**

(f) Radioactivity – **Henry Becquerel**

38. IMPORTANT POINTS

(a) Number of sub-shells in the shell = n

(b) Number of orbitals in the shell = n^2

(c) Number of electrons in the shell = $2n^2$

(d) No. of orbitals in the sub-shell = $2l + 1$

(e) Number of elliptical orbits according to Sommerfeld = n – 1

(f) Maximum number of spectral lines = $\frac{n(n-1)}{2}$

(g) Number of nodal surfaces = n – 1

(h) K. E. = – T.E.

(i) T.E. = P.E./2

(j) Nodal Point : The nucleus of an atom called **Nodal Point**.

(k) Isodiapheres : The elements which have same value of (n – p) is called **isodiapheres**.

(l) Isomorphous : The two different type of compound which contain same crystalline structure called **isomorphous** and this property called **isomorphism**.

(m) Substance which have same number of electron and atoms called **Isosteres**.

(n) Core : The outer most shell of an atom called **Core** and the number of electron present to that shell is called **Core electron**.

(o) Promotion : The transfer of electron between subshells in an orbit is called promotion. While the transfer of one energy level to another is called transition. After the completion of promotion the transition process is occurred.

S.No.	Radius	Velocity	Energy	Wavelength
1.	$r = \frac{n^2h^2}{4\pi^2mZe^2}$	$V = \frac{2\pi Ze^2}{nh}$	$E = -\frac{Ze^2}{2r}$	$R = \frac{2\pi^2me^4}{Ch^3}$
2.	$r = \frac{n^2}{Z} \times 0.529Å$	$V = \left(\frac{Ze^2}{rm}\right)^{1/2}$	$E = -\frac{2\pi^2mZ^2e^4}{n^2h^2}$	$\frac{1}{\lambda} = R \times Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$
3.	$r = \frac{n^2}{Z} \times 0.0529$ nm	$V = \frac{nh}{2\pi mr}$	$E = -Rch \times \frac{z^2}{n^2}$	$E = h\upsilon$
4.	$r \propto n^2$ (Z const)	$V \propto \frac{1}{n}$ (Z const)	$E = -\frac{z^2}{n^2} \times 313.6$ Kcal	$\lambda = \frac{h}{mc}$
5.	$\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2}$ (Z const)	$\frac{V_1}{V_2} = \frac{n_2}{n_1}$ (Z const)	$E \propto -Z^2$ (n const)	c = v/t
6.	$r \propto 1/Z$ (n const)	Time period $T = \frac{2\pi r}{V}$	$\frac{E_1}{E_2} = \frac{Z_1^2}{Z_2^2}$ (n const)	$E = mc^2$