ATOMIC STRUCTURE

Physical Constants

Constant and Symbol[b]		SI Value	Gaussian Value
Speed of light in vaccum	c	2.99×10^{8} m/s	2.99×10^{10} cm/s
Proton & electron charge	e	1.60×10^{-19} C	4.8×10^{-10} statC
Permitivity of vaccum	ε_0	8.85×10^{-12} C^2/N-m^2	
Avogadro constant	N_A	6.02×10^{23} mol^{-1}	6.02×10^{23} mol^{-1}
Electron rest mass (0.000548 amu)	m_e	9.10×10^{-31} kg	9.10×10^{-28} g
Proton rest mass (1.00757 amu)	m_P	1.67×10^{-27} kg	1.67×10^{-24} g
Neutron rest mass (1.00893 amu)	m_n	1.67×10^{-27} kg	1.67×10^{-24} g
Planck constant	h	6.62×10^{-34} J s	6.62×10^{-27} erg s
Permeability of vaccum	μ_0	$4\pi \times 10^{-7}$ NC^{-2} s^2	
Bohr radius	a_0	5.29×10^{-11} m	0.529×10^{-8} cm
Bohr's velocity		$2.188 \times 10^{6} \times \frac{Z}{n}$ m/sec.	$2.188 \times 10^{8} \times \frac{Z}{n}$ cm/sec.
Bohr's energy (–13.6 eV/atom)		$-21.8 \times 10^{-19} \frac{Z^2}{n^2}$ J/atom	-21.8×10^{-12} erg/atom
Bohr magneton (BM)	β_e	9.27×10^{-24} J/T	
Gas constant	R	8.3145 J/mol-K	8.3145×10^{7} erg/mol-K
Boltzmann constant	k	1.38×10^{-23} J/K	1.30×10^{-16} erg/K
Gravitational constant	G	6.67×10^{-11} m^3/kg -s^2	6.67×10^{-8} cm^3/g-s^2

Energy Conversion Factors

1 erg = 10^{-7} J

1 cal = 4.184 J

1 eV = 1.602177×10^{-19} J = 1.602177×10^{-12} erg = 23.0605 kcal/mol

Greek Alphabet

Alpha	Α	α	Beta	Β	β
Gamma	Γ	γ	Delta	Δ	δ
Epsilon	Ε	ε	Zeta	Ζ	ζ
Eta	Η	η	Theta	Θ	θ
Iota	Ι	ι	Kappa	Κ	κ
Lambda	Λ	λ	Mu	Μ	μ
Nu	Ν	ν	Xi	Ξ	ξ
Omicron	Ο	ο	Pi	Π	π
Rho	Ρ	ρ	Sigma	Σ	σ
Tau	Τ	τ	Upsilon	Υ	υ
Phi	Φ	φ	Chi	Χ	χ
Psi	Ψ	ψ	Omega	Ω	ω

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 $^{40}_{20}Ca$.

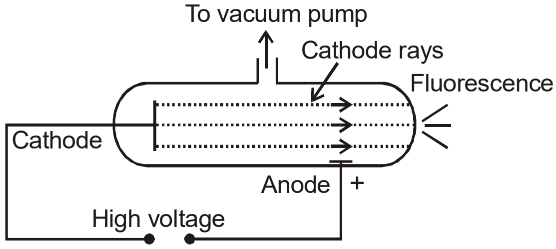
(f) Atoms of elements combines to form molecules.

(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.

(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.

(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.

(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 :

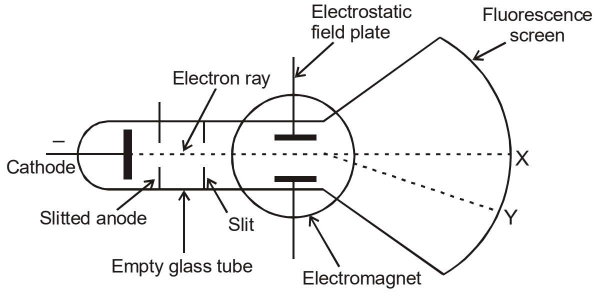
J.J. Thomson (1897) proved through experiments that.

(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 ${}_{-1}e^0$.



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 $= \dfrac{1.6725 \times 10^{-24}}{1.66 \times 10^{-24}} = 1.00757$ 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 DISCOVERY & THEIR DISCOVERERS

	Name of Particles	Scientist	Mass	Charge
1.	Electron	J.J. Thomson	9.1×10^{-31} kg	-1.6×10^{-19} cb
2.	Proton	Goldstein	1.673×10^{-27} kg	$+ 1.6 \times 10^{-19}$ cb
3.	Neutron	Chadwick	1.675×10^{-27} kg	Zero
4.	Positron	C.D. Anderson	(same as electron)	same as proton
5.	Anti Proton	Sugri	(same as proton)	electron
6.	Neutrino	Pauling	Negligible	Zero
7.	Meson	Yukawa	(200 times than electron)	(+ , –, zero)
8.	Isotopes	Soddy		
9.	Isobar	Aston		
10.	Cathode Ray	William Crooke's		
11.	Anode Ray	GoldStein		
12.	Neucleus	Rutherford		
13.	Atomic No.	Moseley		
14.	Nomenclature of e^-	Stoney		
15.	Charge of e^-	Millikan		
16.	Specific charge on e^- (e/m)	J.J. Thomson		

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. Some important definations

1. Atomic Number :

It is represented by Z. The number of protons present in the nucleus is called atomic number.of an element. It is also known as nuclear charge.

For neutral atom : Number of proton = Number of electron

For charged atom : Number of e^- = Z – (charge on atom)

Z= number of protons only

2. Mass Number :

It is represented by capital A The sum of number of Neutrons and protons is called the mass number. of the element. It is also known as number of nucleons becuase neutron & proton are present in nucleus.

A = number of protons + number of Neutrons

Note : It is always a whole number.

Atom

Inside the nucleus $[n, p^+]$ — Outside the nucleus $[e^-]$

An atom of the element is represented by $_ZX^A$

Where, X = Symbol of element

Z = Atomic number = no. of proton = no. of e^- (If atom is neutral)

A = Mass number = no.of neutron + Atomic no.

eg.

$_{11}Na^+$	$_9F^-$	$_6C^{12}$	$_8O^{16}$
$(p^+ \rightarrow 11)$	$(p^+ \rightarrow 9)$	$(p^+ \rightarrow 6)$	$(p^+ \rightarrow 8)$
$(e^- \rightarrow 10)$	$(e^- \rightarrow 9 + 1 = 10)$	$(e^- \rightarrow 6)$	$(e^- \rightarrow 8)$

eg.

$_6C^{12}$
$p^+ \rightarrow 6$
$n^0 \rightarrow 12 - 6 = 6$
$e^- \rightarrow 6$

$_8O^{16}$
$p^+ \rightarrow 8$
$n^0 \rightarrow 16 - 8 = 8$
$e^- \rightarrow 8$

Mass no. [A] and **atomic weight** (a.m.u.= atomic mass unit)

Mass of Proton (m_p)	Mass of Neutron (m_n)	Mass of electron (m_e)
1.673×10^{-27} kg	1.675×10^{-27} kg.	9.1×10^{-31} kg.
1.673×10^{-24} grams	1.675×10^{-24} g.	9.1×10^{-28} kg.
1.00750 a.m.u.	1.00850 a.m.u.	0.000549 a.m.u.

$[m_p \simeq m_n]$ $\qquad$ $[m_n > m_p]$

Method for Analysis of atomic weight $\rightarrow$

eg. $\quad {}_{6}C^{12}$

$p^+ \rightarrow 6$ $\quad$ Weight of Proton $= 6 \times 1.00750$

$n^0 \rightarrow 6$ $\quad$ Weight of Neutron $= 6 \times 1.00850$

$e^- \rightarrow 6$ $\quad \dfrac{\text{Weight of electron} = 6 \times 0.000549}{\text{Weight of C atom} = 12.011 \text{ a.m.u.}}$

Mass no. of C atom = 12 [p^+ and n]

Note : Mass no. of atom is always a whole no. but atomic weight may be in decimal.

Q. If no. of protons in X^{-2} is 16. then no. of e^- in X^{+2} will be–

(1) 14 $\qquad$ (2) 16 $\qquad$ (3) 18 $\qquad$ (4) None

Sol. $\because$ No. of proton in X^{-2} is = 16

$\therefore$ No. of electron in X^{+2} is = 14

Q. In C^{12} atom if mass of e^- is doubled and mass of proton is halved, then calculate the percentage change in mass no. of C^{12}.

Sol.

${}_{6}C^{12}$

e^-	p^+	n°	
6	6	6	$A \rightarrow 12$
12	3	6	$A \rightarrow 9$

$$\% \text{ change} = \frac{3}{12} \times 100 = 25\%$$

Q. Assuming that atomic weight of C^{12} is 150 unit from atomic table, then according to this assumption, the weight of O^{16} will be :-

Sol. $\because$ 12 amu = 150

$\therefore$ 1 amu $= \dfrac{150}{12}$

$\therefore$ 16 amu $= \dfrac{150}{12} \times 16 = 200$ Unit

3. **Isotopes** : (Given by Soddy) They are atoms of a given element which have the same atomic no. but differ in their mass no.

eg. $\quad {}_{6}C^{12}, {}_{6}C^{13}, {}_{6}C^{14}$

${}_{8}O^{16}, {}_{18}O^{17}, {}_{18}O^{18}$

${}_{1}H^{1}, {}_{1}H^{2}, {}_{1}H^{3}$

${}_{6}C^{12}\,{}_{6}C^{13} \quad {}_{6}C^{14}$

$p^+ \rightarrow 6$	6	6
$e^- \rightarrow 6$	6	6
$n^\circ \rightarrow 6$	7	8

Note : Isotopes have the same nuclear charge but differ in the number of neutrons in the nucleus.

$_1H^1$ Protium (H)	$_1H^2$ Deuterium (D)	$_1H^3$ (Radioactive element) Tritium (T)
$p^+ \rightarrow 1$	1	1
$e^- \rightarrow 1$	1	1
$n^\circ \rightarrow 0$	1	2

* Neutron is not available in Protium

* No. of Neucleon = No. of Neutron + No. of Proton

$= n + p^+$

Que.– There are 3 Neucleons in which Isotopes of the Hydrogen.

Sol. – Neucleons $= n + p^+ = 3$, so Isotope will be Tritium

4. Atomic Weight : The atomic weight of an element is the average of weights of all the isotopes of that element.

* An element have there isotopes y_1, y_2 and y_3 and their isotopic weights are w_1, w_2, w_3 and their percentage/ possibility/probability/ratio of occurance in nature are x_1, x_2, x_3 respectively, then the ave. atomic weight of element is –

$$\text{ave. wt} = \frac{w_1x_1 + w_2x_2 + w_3x_3}{x_1 + x_2 + x_3}$$

eg.	Cl^{35}		Cl^{37}
Probability ratio	75%		25%
	3	:	1

$$\frac{35 \times 3 + 37 \times 1}{3 + 1} = \frac{142}{4} = 35.5$$

Q. $_{35}Br^{79}$: $_{35}Br^{81}$

1 : 1 (Ratio of occurance)

$$\frac{79 \times 1 + 81 \times 1}{1 + 1} = \frac{160}{2} = 80$$

Probability ratio of Hydrogen	$_1H^1$		$_1H^2$		$_1H^3$
	99.99%	:	0.001	:	10^{-15}%

Q. An element have three isotopes and their isotopic weight are 11, 12 , 13 unit and their percentage of occurance in nature is 85, 10, 5 respectively then calculate the ave. atomic weight of element.

Sol. Average Atomic weight $= \frac{11 \times 85 + 12 \times 10 + 13 \times 5}{85 + 10 + 5}$

$$= \frac{935 + 120 + 65}{100}$$

$$\text{Avg. wt.} = \frac{1120}{100} = 11.2$$

Q. Ave. atomic weight of an element M is 51.7. If two isotopes of M, M^{50} and M^{52} are present then calculate the percentage of occurance of M^{50} in nature.

Sol. M^{50} $\quad M^{52}$

$x_1 + x_2 = 100\%$

$x_2 = (100 - x_1)$

$$wt = \frac{w_1x_1 + w_2x_2}{x_1 + x_2}$$

$$51.7 = \frac{50 \times x_1 + 52 \times x_2}{x_1 + x_2}$$

$$51.7 = \frac{50x_1 + 52(100 - x_1)}{x_1 + (100 - x_1)}$$

$5170 = 50\,x_1 + 5200 - 52x_1 \Rightarrow\ 5170 = -\,2x_1 + 5200$

$2x_1 = 30 \ \Rightarrow\ x_1 = 15$

$M^{50} = 15\%\ \ M^{52} = 85\%$

Q. Calculate the precentage of Deuterium in heavy water.

Sol. D_2O

$({}_1H^2)_2\ O^{16}$

$4 + 16 = 20$ (Moleculer weight)

$\frac{4}{20} \times 100 \Rightarrow$ Ans = 20%

5. Isobar's :

They are atoms with the same mass number but different atomic numbers.

eg. ${}_6C^{14}$, ${}_7N^{14}$

${}_{18}Ar^{40}$, ${}_{20}Ca^{40}$

${}_{14}Si^{30}$, ${}_{15}P^{30}$

6. Isotones/Isoneutronic /Isotonic :

They are atoms of different element which have the same number of neutrons

Ex.1 ${}_6C^{13}$ $\quad {}_7N^{14}$

$n^0 \rightarrow$ 7 $\quad$ 7

Ex.2 ${}_{14}Si^{30}$ $\quad {}_{15}P^{31}$ $\quad {}_{16}S^{32}$

$n^0 \rightarrow$ 16 $\quad$ 16 $\quad$ 16

7. Isodiaphers :

They are the atoms of different element which have the same difference of the number of Neutrons & protons.

Ex.1 ${}_5B^{11}$ $\quad {}_6C^{13}$

$$\left.\begin{array}{l} p = 5 \\ n = 6 \\ e^- = 5 \end{array}\right] \ n - p = 1 \qquad \left.\begin{array}{l} p = 6 \\ n = 7 \\ e^- = 6 \end{array}\right] \ n - p = 1$$

Ex.2 ${}_7N^{15}$ $\quad {}_9F^{19}$

$$\left.\begin{array}{l} p = 7 \\ n = 8 \\ e^- = 7 \end{array}\right] \ n - p = 1 \qquad \left.\begin{array}{l} p = 9 \\ n = 10 \\ e^- = 9 \end{array}\right] \ n - p = 1$$

8. Isosters :

They are the molecules which have the same number of atoms & electrons.

Ex.1		CO_2	N_2O
	Atoms	= 1 + 2	= 2 + 1
		= 3	= 3
	Electrons	= 6 + 8 × 2	= 7 × 2 + 8
		= 22 e^-	= 22e^-
Ex.2		CaO	KF
	Atoms	2	2
	Electrons	20 + 8	19 + 9
		28 e^-	28 e^-
Ex.3		OF_2	HClO
	Atoms	= 3	3
	Electrons	= 8 + 18	1 + 17 + 8
		= 26 e^-	26 e^-

9. Isoelectronic :

They are the atoms/molecules/ions possesing the same no. of electrons, called iso electronic.

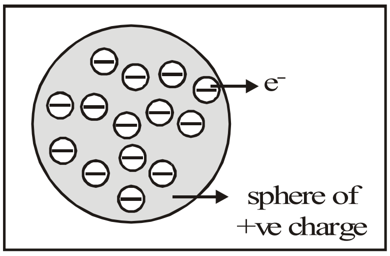
Ex.1		$_{10}Ne$	$_{11}Na^+$		
	e^-	10	10		
Ex.2		NH_4^+	CH_4	Mg^{+2}	H_2O
	e^-	7 + 4 – 1	6 + 4	12 – 2	2 + 8
		= 10e^-	= 10e^-	= 10 e^-	= 10 e^-
Ex.3		Ne	NH_3	NH_2^-	
		10e^-	10e^-	1 + 7 + 2 = 10e^-	

11. THOMSON'S MODEL OF ATOM [1904]

* Thomson was the first to propose a detailed model of the atom.

* Thomson proposed that an atom consists of a uniform sphere of positive charge in which the electrons are present at some places.

* This model of atom is known as "Plum-Pudding model" or "Raisin Pudding Model" or "Water Melon Model".



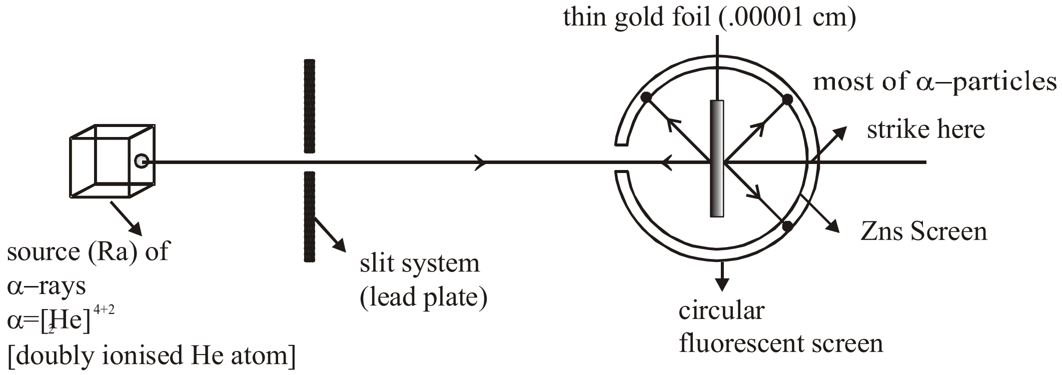
Drawbacks :

* An important drawback of this model is that the mass of the atoms is considered to be evenly spread over that atom.

* It is a static model. It does not reflect the movement of electron.

12. RUTHERFORD's α- SCATTERING EXPERIMENT

α-scattering experiment

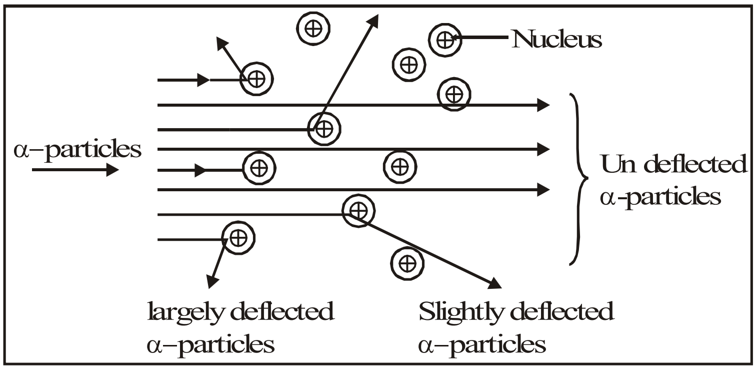


Rutherford observed that -

(i) Most of the α-particles (nearly 99.9%) went straight without suffering any deflection.

(ii) A few of them got deflected through small angles.

(iii) A very few α-particles (about one in 20,000) did not pass through the foil at all but suffered large deflections (more than 90°) or even come back in the direction from which they have come i.e. a deflection of 180°.



Following conclusions were drawn from the above observations -

(1) Since few of the α-particles were deflected from their original path through moderate angles; it was concluded that whole of the +ve charge is concentrated and the space occupied by this positive charge is very small in the atom.

* Whenever a-particles come closer to this point, they suffer a force of repulsion and deviate from their paths.

* The positively charged heavy mass which occupies only a small volume in an atom is called **nucleus**. It is supposed to be present at the centre of the atom.

(2) Since most of the a-particle went straight through the metal foil undeflected, it means that there must be very large empty space within the atom.

(3) A very few of the a-particles suffered strong deflections or even returned on their path indicating that the nucleus is rigid and a-particles recoil due to direct collision with the heavy positively charged mass.

(4) The relation between number of deflected particles and deflection angle q is

$$\mu \propto \frac{1}{\sin^4 \frac{\theta}{2}}$$ [q increases, m decreases] where m = deflected particles q = deflection angle

* As atomic number increases, the number of protons increases which increases the repulsion and so deflection angle q increases.

12.1 APPLICATIONS OF RUTHERFORD MODEL

On the basis of scattering experiments, Rutherford proposed the model of an atom, which is known as nuclear atomic model. According to this model -

(i) An atom consists of a heavy positively charged nucleus where all the protons are present.

(ii) The volume of the nucleus is very small and is only a minute fraction of the total volume of the atom. Nucleus has a radius of the order of 10^{-13} cm and the atom has a radius of the order of 10^{-8} cm

$$\frac{r_A}{r_N} = \frac{\text{radius of the atom}}{\text{radius of the nucleus}} = \frac{10^{-8}}{10^{-13}} = 10^5 \quad , \quad r_A = 10^5\, r_N$$

Thus radius (size) of the atom is 10^5 times the radius of the nucleus.

* The radius of a nucleus is proportional to the cube root of the no. of nucleons within it.

$R \propto A^{1/3} \quad \Rightarrow \quad R = R_0A^{1/3}$ cm

Where $R_0 = 1.33 \times 10^{-13}$(a constant) an, A = mass number (p + n)

R = radius of the nucleus. $\Rightarrow \quad R = 1.33 \times 10^{-13} A^{1/3}$ cm

(iii) There is an empty space around the nucleus called extra nuclear part. In this part electrons are present. The no. of electrons in an atom is always equal to no. of protons present in the nucleus. As the nuclear part of atom is responsible for the mass of the atom, the extra nuclear part is responsible for its volume. The volume of the atom is about 10^{15} times the volume of the nucleus.

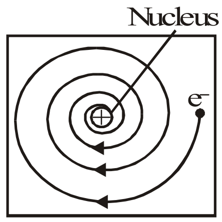
$$\frac{\text{vol. of the atom}}{\text{vol. of the nucleus}} = \frac{\left(\frac{4}{3}\pi r_A^3\right)}{\left(\frac{4}{3}\pi r_N^3\right)} = \frac{\left(10^{-8}\right)^3}{\left(10^{-13}\right)^3} = 10^{15}$$

(iv) Electrons revolve round the nucleus in closed orbits with high speeds and the elctrostatic force of attraction between electrons and nucleus is balanced by centrifugal force acting on electrons.

* This model was similar to the solar system, the nucleus representing the sun and revolving electrons as planets and so also known as Rutherford planetary model of atom.

Drawbacks of Rutherford model -

(1) This theory could **not** explain the stability of an atom. According to Maxwell electron loses it's energy continuously in the form of electromagnetic radiations. As a result of this, the e^- should loss energy at every turn and move closer and closer to the nucleus following a spiral path. The ultimate result will be that it will fall into the nucleus, thereby making the atom unstable.



(2) If the electrons loss energy continuously, the observed spectrum should be continuous but the actual observed spectrum consists of well defined lines of definite frequencies (discontinuous). Hence, the loss of energy by electron is **not** continuous in an atom.

13. ELECTROMAGNETIC WAVES (EM WAVES) OR RADIANT ENERGY/ELECTROMAGNETIC RADIATION

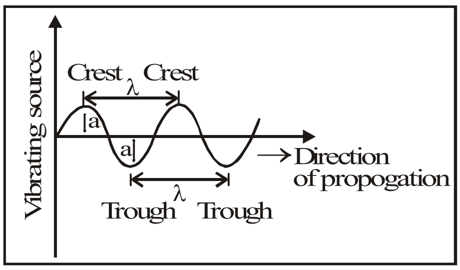
* It is the energy transmitted from one body to another in the form of waves and these waves travel in the space with the same speed as light (3×10^8 m/s) and these waves are known as Electromagnetic waves or radiant energy.
* The radiant Energy do not need any medium for propagation.

Ex : Radio waves, micro waves, Infra red rays, visible rays, ultraviolet rays, x–rays, gama rays and cosmic rays.

* The electro magnetic radiations have electric and magnetic field perpendicular to each other and to the direction of propagation of light.

A wave is characterized by following six characterstics

The upper most point of the wave is called crest and the lower most point is called trough. Some of the terms employed in dealing with the waves are described below.



1. Wavelength λ (Lambda) :

It is defined as the distance between two nearest crest or nearest trough.

It is measured in terms of a A° (Angstrom), pm (Picometre), nm (nanometer), cm(centimetre), m (metre)

1Å = 10^{-10} m, 1 pm = 10^{-12} m, 1nm = 10^{-9} m, 1cm = 10^{-2}m

2. Frequency (ν) (nu)→

Frequency of a wave is defined as the number of waves which pass through a point in 1 sec.

* It is measured in terms of Hertz (Hz), sec^{-1}, or cycle per second (cps)

1 Hertz = 1 sec^{-1}

3. **Time period (T) :** Time taken by a wave to pass through one point. $T = \frac{1}{\nu}$ sec.

4. **Velocity → (c)**

Velocity of a wave is defined as distance covered by a wave in 1 sec.

$c = \lambda / T = \lambda\nu$ $\quad \nu = c/\lambda$ $\quad c = \nu\ (sec^{-1}) \times \lambda\ (m)$ $\quad c = \nu\lambda\ (m\ sec^{-1})$

Since c is constants

i.e. frequency is inversely propotional to λ

5. **Wave number** → ($\bar{\nu}$) (nu bar) → It is the reciprocal of the wave length that is number of waves present in 1cm

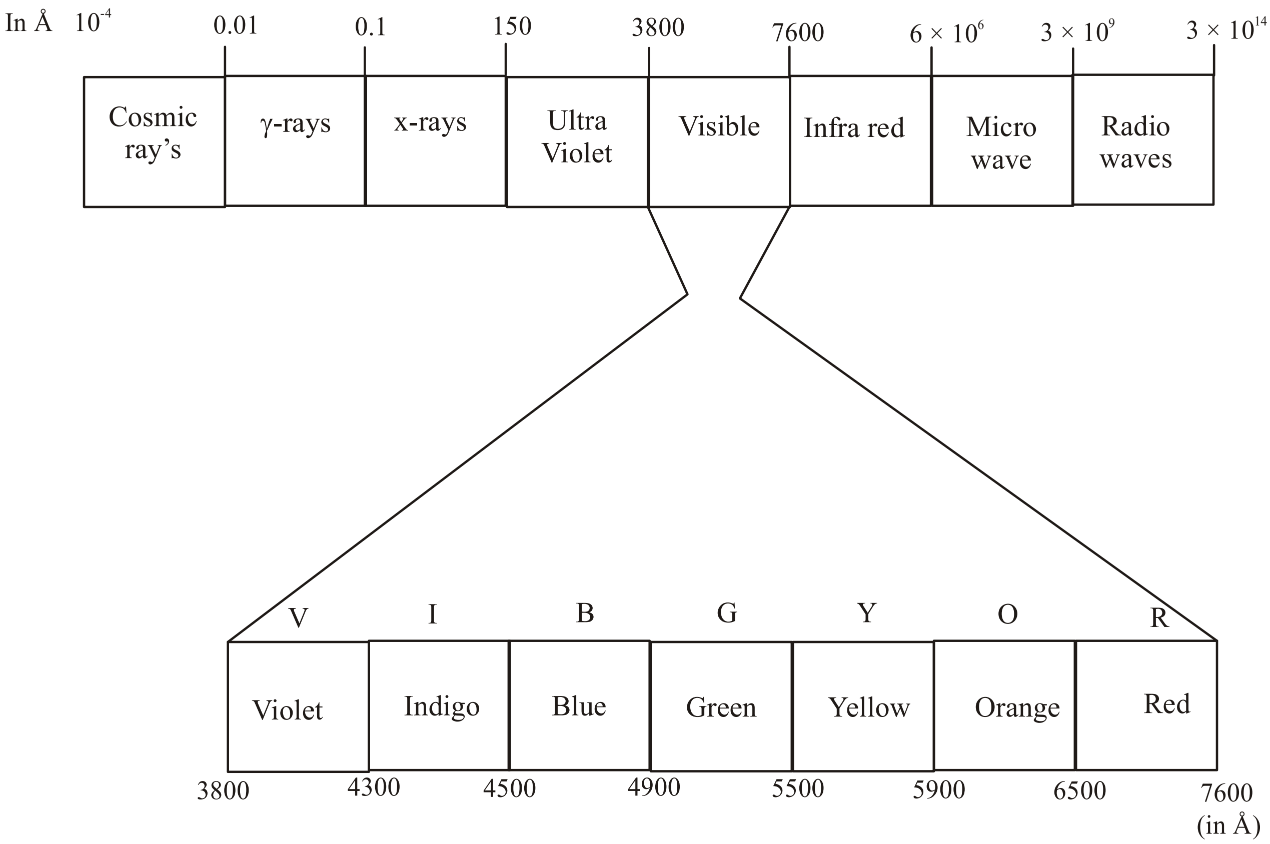
$\bar{\nu} = \frac{1}{\lambda}$ $\quad (1cm^{-1} = 100\ m^{-1})$

* It is measured in terms of cm^{-1}, m^{-1} etc,

6. **Amplitude → (a)**

The amplitude of a wave is defined as the height of crest or depth of trough.

THE ELECTROMAGNETIC SPECTRUM



14. PLANCK'S QUANTUM THEORY

According to planck's quantum theory :

1. The radiant energy is emitted or absorbed by a body not continuously but discontinuously in the form of small discrete packets of energy and these packets are called quantum.
2. In case of light, the smallest packet of energy is called as 'photon' but in general case the smallest packet of energy called as quantum.
3. The energy of each quantum is directly proportional to frequency of the radiation i.e.

$E \propto \nu \quad \Rightarrow \quad E = h\nu \quad$ or $\quad E = \frac{hc}{\lambda} \quad \left\{\because \nu = \frac{c}{\lambda}\right\}$

Proportionality constant or Plank's constant (h)

$h = 6.626 \times 10^{-37}$ kJ sec.

or 6.626×10^{-34} J sec $\quad (1\text{erg} = 10^{-7}\text{J})$

or 6.626×10^{-27} erg sec.

4. Total amount of energy transmited from one body to another will be some integral multiple of energy of a quantum.

$E = nh\nu$

Where n is an integer andn = number of quantum

$$E = h\nu = \frac{hc}{\lambda} = hc\overline{\nu}$$

15. Bohr's Atomic Model

Some Important formulae :

Coulombic force = $\frac{kq_1q_2}{r^2}$

Centrifugal force = $\frac{mv^2}{r}$

Angular momentum = mvr

* This model was based on quantum theory of radiation and Classical laws of physics.
* Bohr model is applicable only for single electron species like H, He^+, Li^{2+} etc.
* Bohr model is based on particle nature of electron.

The important postulates on which Bohr's Model is based are the following :

1st Postulate :

* Atom has a nucleus where all protons and neutrons are present.
* The size of nucleus is very small and it is present at the centre of the atom.

2nd Postulate :

* Negatively charged electron revolve around the nucleus in the same way as the planets revolve around the sun.
* The path of electron is circular.
* The attraction force (Coulombic or electrostatic force) between nucleus and electron is equal to the centrifugal force on electron.

i.e. Attraction force towards nucleus = centrifugal force away from nucleus.

3rd Postulate :

* Electrons can revolve only in those orbits in which angular momentum (mvr) of electron is integral multiple of $\frac{h}{2\pi}$

i.e. $\boxed{mvr = \frac{nh}{2\pi}} = n\hbar \qquad \hbar = \text{hash} = \frac{h}{2\pi}$

where : n = Whole number
h = Planck's constant,
p = Constant

* Angular momentum can have values such as $\frac{h}{2\pi}, 2\frac{h}{2\pi}, 3\frac{h}{2\pi}, 4\frac{h}{2\pi}, 5\frac{h}{2\pi}$ ....but can not have fractional values such as $1.5\frac{h}{2\pi}, 1.2\frac{h}{2\pi}, .5\frac{h}{2\pi}$ ....

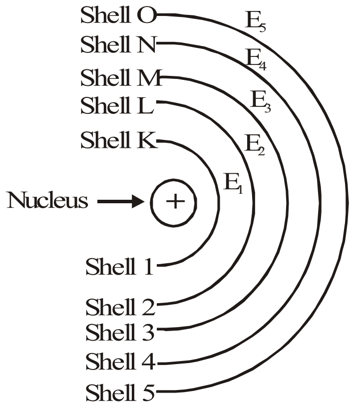
4th Postulate :

* The orbits in which electron can revolve are known as **stationary Orbits** because in these orbits energy of electron is always constant.

5^{th} Postulate :

* Each stationary orbit is associated with definite amount of energy therefore these orbits are also called as energy levels and are numbered as 1, 2, 3, 4, 5, ... or K, L, M, N, O, ... from the nucleus outwards.

6^{th} Postulate



* The emission or absorbtion of energy in the form of photon can only occur when electron jumps from one stationary state to another & it is

D E = $E_{\text{final state}} - E_{\text{initial state}}$ = Energy of a quantum

* Energy is absorbed when electron Jumps from inner to outer orbit and is emitted when electron moves from outer to inner orbit.

Important Definations :-

(i) Ionization energy : Minimum the energy required to liberate an electron from the ground state of an isolated atom is called the ionization energy.

(ii) Separation energy : Minimum energy required to remove an electron from its excited state is called as separation energy.

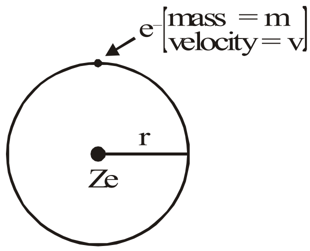
(iii) Excitation energy : Amount of energy required to shift an electron from ground state to any excited state.

Note : All these kinds of energy are always positive.

15.1 APPLICATION OF BOHR'S MODEL

when electron revolves in fixed circular orbit than electrostatic force of attraction and centrifugal force are equal.

Electrostatic force $= \dfrac{Kq_1q_2}{r^2} = \dfrac{K.Ze.e}{r^2} = \dfrac{KZe^2}{r^2}$



Where, constant $K = 9 \times 10^9\ Nm^2/C^2$ (MKS) = 1 (CGS)

Centrifugal force $= \dfrac{mv^2}{r}$

In balanced condition

Electrostatic force = Centrifugal force

$$\frac{KZe^2}{r^2} = \frac{mv^2}{r} \quad \text{or} \quad \frac{KZe^2}{r} = mv^2 \quad \text{or} \quad \frac{Ze^2}{r} = mv^2 \ \text{(CGS)} \quad ....(1)$$

1. RADIUS OF VARIOUS ORBITS (SHELL) :

According to Bohr model $mvr = \dfrac{nh}{2\pi}$

$$v = \frac{nh}{2\pi mr} \qquad ....(2)$$

Now putting the value of v from eq.(2) into eq.(1)

$$\frac{KZe^2}{r} = m\left(\frac{nh}{2\pi mr}\right)^2$$

$$\frac{KZe^2}{r} = \frac{mn^2h^2}{4\pi^2m^2r^2}$$

$$r = \frac{n^2h^2}{4\pi^2 mKZe^2} \qquad \text{or} \qquad r = \frac{n^2h^2}{4\pi^2 mZe^2} \quad (\text{CGS} \because K = 1) \qquad ....(3)$$

Putting the value of π, h, m, K, & e (Constants) in the above eq (3)

$$r = 0.529 \times 10^{-10} \times \frac{n^2}{Z} \text{ m } \{ \text{Å} = 10^{-10}\text{m} = 10^{-8} \text{ cm}\}$$

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

This formula is only applicable for hydrogen and hydrogen like species i.e. species containing single electron.

2. VELOCITY OF ELECTRON IN BOHR ORBIT :

A/c to Bohr postulate

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

$$v = \frac{nh}{2\pi mr} = \frac{nh}{2\pi m \times \dfrac{n^2h^2}{4\pi^2 mKZe^2}}$$

$$v = \frac{2\pi KZe^2}{nh} \text{ (MKS)} \qquad ....(4)$$

$$v = \frac{2\pi Ze^2}{nh} \quad \text{(CGS)}$$

Putting the value of p, h, K, & e (Constants) in the above eq (4)

$$\mathbf{v = 2.18 \times 10^6 \frac{Z}{n} m/s}$$

$$\mathbf{v = 2.18 \times 10^8 \frac{Z}{n} cm/s}$$

3. TOTAL ENERGY OF ELECTRON IN BOHR ORBIT :

Total energy of an electron is the sum of kinetic and potential Energy.

i.e. T.E. = K.E.+ P.E.

(i) Potential energy: $\text{P.E.} = -\frac{Kq_1q_2}{r}$

$$\text{P.E.} = -\frac{KZe^2}{r} = -\frac{Ze^2}{r}$$

(ii) Kinetic energy : $\text{K.E.} = \frac{1}{2}mv^2$

But $\frac{KZe^2}{r} = mv^2$ (By eq. 1)

$$\text{K.E.} = \frac{KZe^2}{2r}$$

(iii) Total energy : T.E. = K.E.+ P.E.

$$\text{T.E.} = \frac{KZe^2}{2r} - \frac{KZe^2}{r}$$

$$\text{T.E.} = -\frac{KZe^2}{2r}$$

now putting the value of r from eq. (3)

$$\text{T.E.} = -\frac{KZe^2 \times 4\pi^2 mKZe^2}{2n^2h^2}$$

$$T.E. = -\frac{2\pi^2 m \times K^2 Z^2 e^4}{n^2 h^2}$$

now putting the value of p, K, e, m, h, we get :

$$T.E. = -2.18 \times 10^{-18} \times \frac{Z^2}{n^2} \text{ J / atom} = -1312 \times \frac{Z^2}{n^2} \text{ kJ/mol}$$

$$= -2.18 \times 10^{-11} \times \frac{Z^2}{n^2} \text{ erg/atom} = -313.6 \times \frac{Z^2}{n^2} \text{ Kcal/mol}$$

$$= -13.6 \times \frac{Z^2}{n^2} \text{ eV/atom}$$

Some extra points :

(i) K.E $= \frac{KZe^2}{2r}$ i.e. K.E. μ $\frac{1}{r}$ On increasing radius, K.E. decreases.

(ii) P. E. $= -\frac{KZe^2}{r}$ i.e. P.E. μ $-\frac{1}{r}$ On increasing radius, P.E. increases.

(iii) T.E. $= -\frac{KZe^2}{2r}$ i.e. E. μ $-\frac{1}{r}$ On increasing radius, total energry increases.

(iv) Relation between T.E., P.E. and K.E.

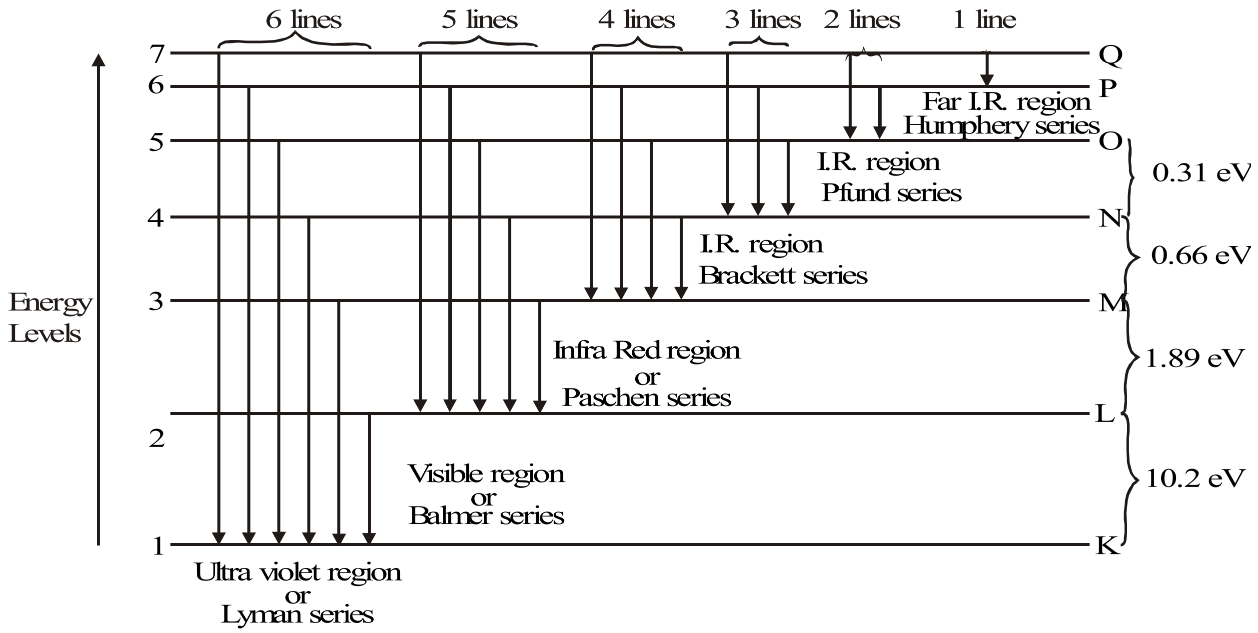
P.E = –2 KE

KE = –T.E.

P.E = 2 T.E.

16. Hydrogen line spectrum or Hydrogen spectrum :

When an electric excitation is applied on atomic hydrogen gas at Low pressure,a bluish light is emitted. when a ray of this light is passed through a prism, a spectrum of several isolated sharp lines is obtained.The wavelength of various lines show that spectrum lines lie in visible, Ultraviolet and Infra red region. These lines are grouped into different series.



Series	Discovered by	regions	$n_2 \rightarrow n_1$	Number of lines
Lyman	Lyman	U.V. region	$n_2 = 2,3,4 \ldots / n_1=1$	$n_2 - 1$
Balmer	Balmer	Visible region	$n_2 = 3,4,5 \ldots / n_1=2$	$n_2 - 2$
Paschen	Paschen	Infra red (I.R.)	$n_2 = 4,5,6 \ldots / n_1=3$	$n_2 - 3$
Brackett	Brackett	I.R. region	$n_2 = 5,6,7 \ldots / n_1=4$	$n_2 - 4$
Pfund	Pfund	I.R. region	$n_2 = 6,7,8 \ldots / n_1=5$	$n_2 - 5$
Humphery	Humphery	Far I.R. region	$n_2 = 7,8,9 \ldots / n_1=6$	$n_2 - 6$

SIMILAR WORDS

* First line / Starting line / Initial line ($\lambda_{max.}$ and ν_{min})

* Last line / limiting line / marginal line (λ_{min} and $\nu_{max.}$)

* First line of any series = α line

Second line of any series = β line

Third line of any series = γ line

* Total no. of lines = $\dfrac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$

For transition from n = any orbit to n = 1 , total no. of lines = $\dfrac{n(n-1)}{2}$

RYDBERG FORMULA

It an electron shows transition from n_2 to n_1 energy level then energy change DE will be.

$$\Delta E = En_2 - En_1$$

$$\Delta E = \frac{-2\pi^2 mK^2Z^2e^4}{n_2^2h^2} - \left[\frac{-2\pi^2 mK^2Z^2e^4}{n_1^2h^2}\right] = \frac{2\pi^2 mK^2Z^2e^4}{n_1^2h^2} - \frac{2\pi^2 mK^2Z^2e^4}{n_2^2h^2}$$

But $\Delta E = hn$

$$= \frac{hc}{\lambda}$$

$$\therefore \quad \frac{hc}{\lambda} = \frac{2\pi^2mK^2Z^2e^4}{h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \Rightarrow \frac{1}{\lambda} = \frac{2\pi^2mK^2e^4Z^2}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

where $\dfrac{2\pi^2mK^2e^4}{ch^3}$ is a constant called Rydberg constant (R).

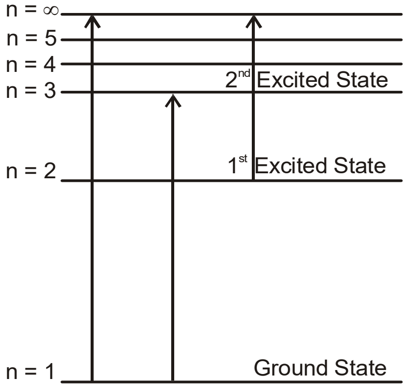
So, $$\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

value of R = 109678 cm^{-1} = 10967800m^{-1}

$\simeq$ 109700 $cm^{-1} \simeq$ 10970000 m^{-1}

$$\boxed{\frac{1}{R} = 912\,\text{Å}}$$

17. Quantization of electronic energy Levels



Quantum State	Energy	Separation Energy	Excitation Energy
$n = \infty$	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
$n = 3$ (2nd Excited State)	$-13.6/3^2 = 1.51$	1.51	12.1
$n = 2$ (1st Excited State)	$-13.6/2^2 = -3.4$	3.4	10.2
$n = 1$ (Ground State)	–13.6 eV	13.6 eV	0 eV

Electronic energy levels of hydrogen atoms

18. Limitation of the Bohr's model :

1. Bohr's theory does not explain the spectrum of multi electron atom.
2. Why the Angular momentum of the revolving electron is equal to $\frac{nh}{2\pi}$, has not been explained by Bohr's theory.
3. Bohr inter-related quantum theory of radiation and classical law of physics without any theoretical explaination.This was the biggest drawback of this model.
4. Bohr's theory does not explain the fine structure of spectral lines. Fine structure of the spectral line is obtained when spectrum is viewed by a spectroscope of high resolving power.
5. Bohr's theory does not explain the splitting of spectral lines in the presence of magnetic field **(Zeeman effect)** or electric field **(Stark effect)**

19. Sommerfeld Extension of the Bohr's Model

* According to sommerfeld electron revolve around the nucleus in the Elliptical Orbits.
* Sommerfeld gave the idea of sub orbits or sub levels
* For any particular orbit **n**

  Total number of path = n

  Number of elliptical path = (n – 1)

  Number of circular path = 1

Note : In every atom, 1st orbit is always circular.

* According to sommerfeld, number of fine lines in a spectral line resulting from a transition between n_2 to n_1 energy level = $n_2 \times n_1$

  Ex. If electron shows transition from $n_2 = 4$ to $n_1 = 2$

  Then, number of fine lines in the spectral line = $n_2 \times n_1 = 4 \times 2 = 8$

20. De- Broglie concept (Dual nature of matter)

1. In 1923, a French physicist, Louis de Broglie suggested that, like light, matter also has dual character. It exhibits wave as well as particle nature.

2. According to de Broglie, the wavelength λ of an electron is inversely proportional to its momentum p.

$$\lambda \propto \frac{1}{p} \quad \text{or} \quad \lambda \propto \frac{1}{mv}$$

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

p = momentum of electron

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

3. **Derivation of de-Broglie Relation**

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

$E = mc^2$ (Einstein's equation) ....(i)

Where E is energy, m is mass of a body and c is its velocity.

$E = h\nu = h \times \frac{c}{\lambda}$ (Planck's equation) $\left(\nu = \frac{c}{\lambda}\right)$ ....(ii)

combining (i) and (ii)

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

$$\boxed{\lambda = \frac{h}{mv} \text{ or } \lambda = \frac{h}{p}}$$

4. It is clear from the above equation that the value of λ decreases on increasing either m or v or both. The wavelength of many fast-moving objects like an aeroplane or a cricket ball, is very low because of their high mass.

Important points concerned with de-Broglie concept

* **de-Broglie wavelength in terms of kinetic energy.**

Kinetic Energy (K.E.) = $\frac{1}{2}mv^2$

or $m \times K.E. = \frac{1}{2}m^2v^2$ or $m^2v^2 = 2m\,K.E.$ or $mv = \sqrt{2m\,K.E.}$

But $\lambda = \frac{h}{mv}$ $\therefore$ $\lambda = \frac{h}{\sqrt{2m\,K.E.}}$ $\left(\because mv = \sqrt{2m\,K.E.}\right)$

* **When a charged particle carrying Q coulomb is accelerated by applying potential difference V then**

K.E. = Q × V Joule

But $\lambda = \frac{h}{\sqrt{2m\,K.E.}}$ $\therefore$ $\lambda = \frac{h}{\sqrt{2mQV}}$ $\boxed{\text{For electron } \left(\lambda = \sqrt{\frac{150}{V}}\ \text{Å}\right)} = \frac{12.25}{\sqrt{V}}$ Å

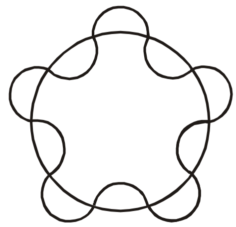
* The wave nature of electron was verified experimentally by Davisson and Germer.
* It is significant only for sub microscopic particles like electron, proton, neutron etc.
* We know that according to Bohr theory, $mvr = \frac{nh}{2\pi}$

or $2\pi r = \frac{nh}{mv}$ ($\because$ mv = p momentum) or $2\pi r = \frac{nh}{p}$ $\left(\because \frac{h}{p} = \lambda \text{ de-Broglie equation}\right)$

$\therefore$ $2\pi r = n\lambda$ (where n = total number of waves 1, 2, 3, 4, 5, ...$\propto$ and λ = Wavelength

* When an electron revolves in a particular Bohr orbit then number of waves made by electron = orbit number.

For ex. If electron revolves in fifth Bohr orbit then number of waves made by electron = 5



n = 5

$\therefore$ Waves made = 5

* Frequency of matter waves

$$\nu = \frac{v}{\lambda}$$

$$\nu = \frac{v}{h/p} = \frac{vp}{h} = \frac{mv^2}{h} = \frac{2K.E.}{h}$$

* Matter waves differ from electromagnetic waves in having :

(i) Lower velocity (ii) No electric and magnetic feilds.

(iii) Matter waves are not emitted by the particle under consideration.

21. Heisenberg uncertainity principle

Bohr's theory considers an electron as a material particle. Its position and momentum can be determined with accuracy. But, when an electron is considered in the form of wave as suggested by de-Broglie, it is not possible to ascertain simultaneously the exact position and velocity of the electon more precisely at a given instant since the wave extends throghout a region of space.

In 1927, Werner Heisenberg presented a principle known as Heisenberg uncertainity principle which states that : "It is impossible to measure simultaneously the exact position and exact momentum of a body as small as an electron."

The uncertainity in measurement of position, Δx, and the uncertainity in momentum Δp or $m\Delta v$, are related by Heisenberg's relationship as : ($p = mv$, $\Delta p = m\Delta v$)

$$\boxed{\Delta x.\Delta p \geq \frac{h}{4\pi}} \quad \text{or} \quad \Delta x\,.\,m\Delta v \geq \frac{h}{4\pi} \quad \text{or} \quad \boxed{\mathbf{\Delta x.\Delta v \geq \frac{h}{4\pi m}}}$$

where h is Planck's constant.

(i) When $\Delta x = 0, \Delta v = \infty$

Remember $\boxed{\frac{h}{4\pi} = 0.527 \times 10^{-34}\ \text{J sec}}$

(ii) When $\Delta v = 0, \Delta x = \infty$

So, if the position is known quite accurately, i.e., Δx is very small, Δv becomes large and vice-versa.

21.1 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

21.2 Schrondinger's Wave Equation

Quantum mechanics, as developed by Erwin Schrodinger in 1926, is based on the wave motion associated with the particles. For the wave motion of the electron in the three dimensional space around the nucleus, he put forward an equation, known after his name as Schrodinger wave equation,which is considered as the heart of quantum mechanics. The equation is :

$$\frac{\partial^2\Psi}{\partial x^2} + \frac{\partial^2\Psi}{\partial y^2} + \frac{\partial^2\Psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$

where ψ is the amplitube of the wave where the coordinates of the electron are (x, y, z), E is the total energy of the electron, V is its potential energy, m is the mass of the electron and h is Planck's constant. $\partial^2\Psi/dx^2$ represents second derivative of ψ w.r.t. x and so on.

21.3 Significance of the wave function :

The wave function ψ for an electron in an atom has no physical significance as such. However, just as in case of light or sound, the square of the amplitude of the wave at any point gives the intensity of the sound or light at that point, similarly the square of the amplitude of the electron wave, i.e. ψ^2 at any point gives the intensity of the electron wave at that point which in view of Heisenberg's uncertainty principle means **probability** of finding the electron at that point. Thus, ψ^2 at any point gives the probability of finding the electron at that point, i.e. electron density at that point. Since are region around the nucleus which represents the electron density at different points is called an orbital, that is why the wave function for an electron in an atom is called **orbital wave function** or simply **atomic orbital**. Since an electron can have many wave functions, therefore, there are many atomic orbitals in an atom.

22. 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.

22.1 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

22.2 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 $n + \ell = 7$?

[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.

22.3 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] $6d_{z^2}$ [2] $6d_{x^2-y^2}$ [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] $3d_{z^2}$ **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 $2p_z$.

22.4 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.

23. Rules for filling of orbitals

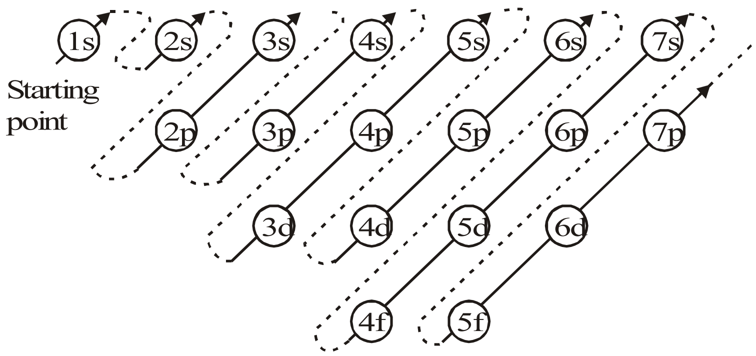
23.1. Aufbau Principle :

Aufbau is a German word and its meaning is **'Building up'**

Aufbau principle gives a sequence in which various subshell are filled up depending on the relative order of the energies of various subshells.

* Principle : The electrons are filled up in increasing order of the energy of the subshells. The subshell with minimum energy is filled up first and when this subshell obtains maximum quota of electrons then the next subshell of higher energy starts filling.

* The sequence in which various subshell are filled is the following.



$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^2, 5f^{14}, 6d^{10}, 7p^6$

23.2. (n+ℓ) rule :

According to it the sequence in which various subshell are filled up can also be determined with the help of (n +ℓ) value for a given subshell.

Principle of (n+ℓ) rule :

The subshell with lowest(n+ℓ) value is filled up first, but when two or more subshells have same (n+ℓ) value then the subshell with lowest value of n is filled up first.

Sub Shell	n	ℓ	n+ℓ	
1s	1	0	1	
2s	2	0	2	
2p	2	1	3	(1)
3s	3	0	3	(2)
3p	3	1	4	(1)
4s	4	0	4	(2)
3d	3	2	5	(1)
4p	4	1	5	(2)
5s	5	0	5	(3)
4d	4	2	6	(1)
5p	5	1	6	(2)
6s	6	0	6	(3)

23.3 Pauli's Exclusion principle :

In 1925, Pauli stated that no two electrons in an atom can have same values of all four quantum numbers.

$\boxed{\uparrow\downarrow}\surd$ $\boxed{\downarrow\uparrow}\surd$ $\boxed{\uparrow\uparrow}\times$ $\boxed{\downarrow\downarrow}\times$

* An orbital can accomodate maximum 2 electrons and that too with opposite spin.

23.4. Hund's Rule of Maximum Multiplicity :

(Multiplicity : Many of the same kind)

* According to Hund's rule electrons are distributed among the Orbitals of subshell in such a way as to give maximum number of unpaired electron with parallel spin. i.e. in a subshell pairing of electron will not start until and unless all the orbitals of that subshell will get one electron each with same spin.

Exceptions of Aufbau principal

In some cases it is observed that the actual electornic configuration of an element is slightly different from the arrangement given by aufbau principal. A simple resion behind this is that the half filled and full filled subshell have got extra stability due to symmetry and exchange energy.

* **Example**

Nb, Mo, Ru, Rh, Pd, Au, Ag, Pt,

La, Ce, Pm, Gd, Ac, Th, Pa, U, Np, Cm, Bk

<u>Electronic Configuration</u>

$_1H \rightarrow 1s^1$ $\quad$ $_2He \rightarrow 1s^2$

$_3Li \rightarrow 1s^2, 2s^1$ $\quad$ $_4Be \rightarrow 1s^2, 2s^2$

$_5B \rightarrow 1s^2, 2s^2, 2p^1$ $\quad$ $_6C \rightarrow 1s^2, 2s^2, 2p^2$

$_7N \rightarrow 1s^2, 2s^2, 2p^3$ $\quad$ $_8O \rightarrow 1s^2, 2s^2, 2p^4$

$_9F \rightarrow 1s^2, 2s^2, 2p^5$ $\quad$ $_{10}Ne \rightarrow 1s^2, 2s^2, 2p^6$

$_{11}Na \rightarrow 1s^2, 2s^2, 2p^6, 3s^1$ $\quad$ $_{12}Mg \rightarrow 1s^2, 2s^2, 2p^6, 3s^2$

$_{13}A\ell \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

$_{14}Si \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$

$_{15}p \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

$_{16}S \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$

$_{17}Cl \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

$_{18}Ar \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

$_{19}K \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$

$_{20}Ca \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$

$_{21}Sc \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^1$

$_{22}Ti \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^2$

$_{23}V \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^3$

$_{24}Cr \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^5$ [Exception]

$_{25}Mn \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^5$

$_{26}Fe \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^6$

$_{27}Co \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^7$

$_{28}Ni \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^8$

$_{29}Cu \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^{10}$ [Exception]

$_{30}Zn \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}$

$_{31}Ga \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^1$

$_{32}Ge \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^2$

$_{33}As \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^3$

$_{34}Se \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^4$

$_{35}Br \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^5$

$_{36}Kr \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6$

24. 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

25. 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%.

25.1 s-Orbital

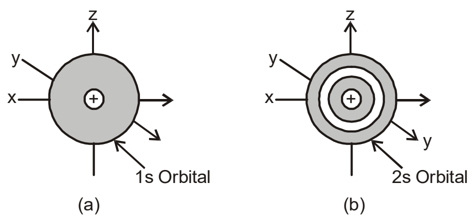
(a) Only one s-orbital is possible in an orbit because $l = 0$ and $m = 0$ for it.

(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

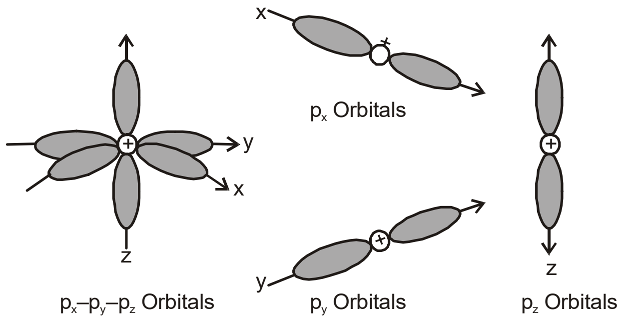
25.2 p-Orbital

(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

25.3 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 :

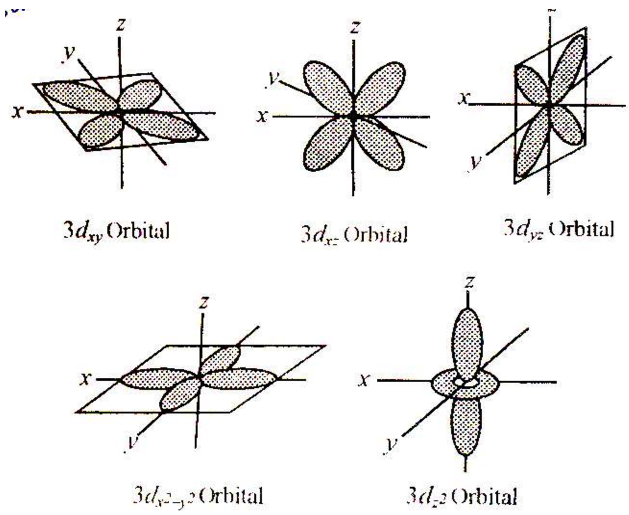
(i) The double dumbbell of d_{xy} orbital are situated between x and y axes.

(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.



Shape of d Orbitals

25.4 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.