Subatomic Particles:

Name	Electron	Proton	Neutron
Symbol	e-	p	n
Approximate relative mass	1/1836	1	1
Mass in kg	9.109×10^{-31}	1.673×10^{-27}	1.675×10^{-27}
Mass in amu	5.485×10^{-4}	1.007	1.008
Charge (coulomb)	1.602×10^{-19}	1.602×10^{-19}	0
Actual Charge (e.s.u)	4.8×10^{-10}	4.8×10^{-10}	0

Atomic Models:

Thomson's Atomic Model (Plum – pudding model):-

Postulate: -Atom is a sphere of positive charge in which number of electrons are embedded.

Limitations: - Could not satisfactorily explain the results of scattering experiment carried out by Rutherford.

Rutherford's Model:

Postulates:-

- Almost all the positive charge and mass of atom is present in its nucleus.
- Electrons revolve around the nucleus in circular orbits.
- There is strong electrostatic attraction between nucleus and electrons

Limitations: - Could not explain stability and electronic structure of atom.

Atomic Terms

Terms	Definition / Explanation
Atomic Number (Z)	Number of protons or electrons of neutral atom.
Mass Number (A)	Total number of protons and neutrons in an atom
Nucleons	Protons and neutrons are present in a nucleus. So, these fundamental particles are collectively known as nucleons

Isotopes	Atoms of the element with same atomic number but different mass number e.g. $_1H^1$, $_1H^2$, $_1H^3$.
Isobars	Atoms having the same mass number but different atomic numbers, e.g. $_{15}P^{32}$ and $_{16}S^{32}$
Isotones	Atoms having the same number of neutrons but different number of protons or mass number, e.g. $_6C^{14}$, $_8O^{16}$, $_7N^{15}$
Isoelectronic	Atoms, molecules or ions having same number of electrons e.g. N_2,CO, CN^-
Nuclear isomers	atoms with the same atomic number and same mass number but with different radioactive properties. Example of nuclear isomers is Uranium –X (half life 1.4 min) and Uranium –Z (half life 6.7 hours
Isosters	Molecules having same number of atoms and also same number of electrons are called isosters. E.g., N_2 and CO

Wave

Terms	Explanation
Wave length (λ)	Distance between two neighbouring troughs or crests.
Frequency (ν)	Number of times a wave passes through a given point in a medium in one second. $\nu = c/\lambda$
Velocity (c)	The distance travelled by the wave in one second. $c = \nu\lambda$
Wave number	Number of wavelengths per cm.
Amplitude (a)	Height of the crest or depth of the trough. Determines the intensity of the beam of light.

Electromagnetic Waves

Radiations	Wave length (Å)
Radio waves	3×10^{14} to 3×10^{7}
Micro waves	3×10^{9} to 3×10^{6}
Infrared (IR)	6×10^{6} to 7600
Visible	7600 to 3800
Ultra violet (UV)	3800 to 150
X–rays	150 to 0.1
Gamma rays	0.1 to 0.01

Gamma Rays	0.1 to 0.01

Atomic spectrum of hydrogen atom:

Where, R_H = Rydberg constant (108978 cm^{-1})

n_1 and n_2 have integral values as follows

n_1	n_2	Spectral Series	Spectral region
1	2,3,4…	Lyman	UV
2	3,4,5…	Balmer	Visible
3	4,5,6…	Pascher	IR
4	5,6,7…	Brackett	IR
5	6,7,8…	Pfund	IR

Photoelectric effect:

- Ejection of electrons takes place from the surface of metal when light of suitable frequency fall on it.
- Minimum frequency required for ejection of electron is called threshold frequency (v_o).
- Energy of the ejected electrons is directly proportional to the frequency of radiation.
- Number of electrons ejected per second depends on the intensity of radiation.
- $hv - hv_o = 1/2 m_e v^2$

Planck's quantum theory:

Substances radiate or absorb energy discontinuously in the form of energy packets

The smallest packet of energy is called quantum. In case of light the quantum is known as photon.

The energy of a quantum is directly proportional to the frequency of the radiation.

E = hv were v is the frequency of radiation and h is Planck's constant having the value 6.626×10^{-27} erg sec or 6.626×10^{-34} J sec.

A body can radiate or absorb energy in whole number multiples of quantum hn, 2hv,3hv……...nhv, where n is the positive integer.

Bohr's atomic model:

Electrons revolve around the nucleus in circular orbits of fixed energy.

Electron revolve only in those orbits whose angular momentum (mvr) is an integral multiple of h/2Π.

Electron absorbs energy in the form of EMR, when it jumps from lower energy level (ground state) to higher energy level (excited

state) and vice-versa.

Energy absorbed or released in an electron jump, (dE) is given by $dE = E_2 - E_1 = h\nu$

Energy of stationary state oh hydrogen atom $(E_n) = -R_H (1/n^2)$

For an hydrogen like species i.e. He^+, Li^{2+} with atomic number Z

Radius of n^{th} orbit $(r_n) = 52.9 \times n^2/z$ pm

Energy of n^{th} orbit $(E_n) = -2.18\times10^{-18}(Z^2/n^2) = -13.6 \times(Z^2/n^2)$ eV $= 313.6 \times(Z^2/n^2)$ kcal /mole

Velocity of electron (v) $= (2.18 \times10^8)$ z/n cms^{-1}

Where n = 1,2,3,4…

Limitations of Bohr's theory:

- Failed to explain the spectra of atoms having more than one electron.
- Failed to account for the splitting of spectral line source of a spectrum is placed in a strong magnetic or electric field.
- Dual nature of particle and the uncertainty principle was ignored in the Bohr's atomic model.

de- Broglie equation:

$\lambda = h/mv = h/p$

Heisenberg's uncertainty principle:

It is impossible to determine simultaneously, the exact position and exact momentum of an electron.

Quantum Numbers:

Principal quantum number (n):

- It tells the main shell in which the electron resides and the approximate distance of the electron from the nucleus.
- Maximum number of electrons a shell can accommodate is $2n^2$.

Azimuthal or angular momentum quantum number (l):

- It represents the number of subshells present in the main shell.
- These subsidiary orbits within a shell will be denoted as s,p,d,f…
- This tells the shape of the subshells.
- For a given value of n, there are n possible values of l starting from 0 to (n-1)

Value of l	0	1	2	3	4

Notation of sub shell	s	p	d	f	g

The magnetic quantum number (m):

- It determines the number of preferred orientations of the electron present in a subshell.
- For a given value of l, there are (2l+1) possible values of m starting from –l to +l.

Value of l	0	1	2	3
Notation of sub shell	s	p	d	f
Values of m	0	-1,0,1	-2,-2,0,1,2	-3,-2,-1,0,1,2,3

The spin quantum number (s)

- It determines the direction of spin of electron in an orbit.
- There are only two possible values for spin quantum number i.e. -1/2 ,+1/2.

Rules for filling of electrons in various orbitals :

Aufbau Principle:

- Electrons are added one by one to the various orbitals in order of their increasing energy starting with the orbital of lowest energy.
- In neutral isolated atom, the lower the value of (n+ l) for an orbital, lower is its energy
- The increasing order of energy of various orbital is : 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d ..

Pauli's Exclusion principle :-

An orbital can contain a maximum number of two electrons and these two electrons must be of opposite spin.

Hund's rule of maximum multiplicity :-

Electron pairing in p, d and f orbital cannot occur until each orbital of a given subshell contains one electron each or is singly occupied".

Exceptional Configurations

Stability of half filled and completely filled orbitals

Cu has 29 electrons. Its expected electronic configuration is

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^9$

But a shift of one electron from lower energy 4s orbital to higher energy 3d orbital will make the distribution of electron symmetrical and hence will impart more stability.

Thus the electronic configuration of Cu is

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^{10}$

Fully filled and half filled orbitals are more stable