MIND MAP

Atomic Structure

Dalton's Atomic Theory

1. Atom is the smallest particle of any substance which cannot be divided further.
2. Atoms can neither be created nor destroyed, only the rearrangement of atoms occur in a chemical reaction.
3. All the atoms are hard and dense.
4. All the atoms of an element are identical in all respect but atoms of different elements are different.

Result of Thomson experiment

1. They travel in a straight line in absence of Electric or Magnetic fields.
2. Electric and Magnetic factors cause deflection of these cathode rays.
3. These rays deflect like a negatively charged particle.
4. Deflection is dependent on strength of Electric charge or Magnetic fields.

$$\frac{\text{charge}}{\text{mass}} = 1.758820 \times 10\ 11\ \text{C Kg}^{-1}$$

Charge on Electron = 1.602 x 10−19 C

X-Rays

Given by Wilhelm Rontgen

When electrons strike a material in the cathode ray tube, they produce rays which can cause fluorescence in the material placed outside the cathode ray tube. These Radiations are called X-Rays.

Radioactivity

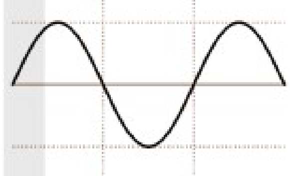
Given by Henri Becquerel, Marie Curie

Certain elements emit radiation on their own and this phenomenon is called Radioactivity.

Thomson's Model

1. Thomson proposed that an atom consists of a uniform sphere of positive charge in which the electrons are embedded.

Wave Characteristics



1) Wavelength- It is defined as the distance between two nearest crests or nearest troughs.

2) Wave number ($\bar{v}$) : It is defined as number of waves per unit length.

$1\ m \rightarrow 1/\lambda$ waves

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

3) Frequency (ν): Frequency of a wave is defined as the number of waves which pass through a point in 1 sec.

1 Hertz = 1 sec^{-1}

4) Time period (T): Time taken by a wave to pass through one point. Or time taken by a wave to complete distance of one wavelength

$$T = \frac{1}{v} \text{ sec}$$

5) Velocity $\rightarrow$ (c) Velocity of a wave is defined as distance covered by a wave in 1 sec.

6) Amplitude $\rightarrow$ (a) It is the height of the crest or depth of the trough of a wave

Conclusion of Rutherford Experiment

Most of the space inside the atom is empty because most of the a-particles passed through the gold foil without getting deflected.

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As a few alpha particles suffered minor deflections & a very few major deflections, this means that these must have met with some obstruction in their path.

A very small fraction of α-particles were deflected by 180o, indicating that all the positive charge and mass of the gold atom was concentrated in a very small volume within the atom.

Rutherford regarded this very small, dense & positively charged obstruction in an atom as nucleus (means centre).

The electrons with negligible mass & negative charge were supposed to be present in the portion around nucleus known as extra nuclear portion.

Rutherford Atomic Model

1. Like charges repel each other.
2. Mass of atoms was approximately coming out to be double the mass of protons.

Drawbacks of Rutherford model

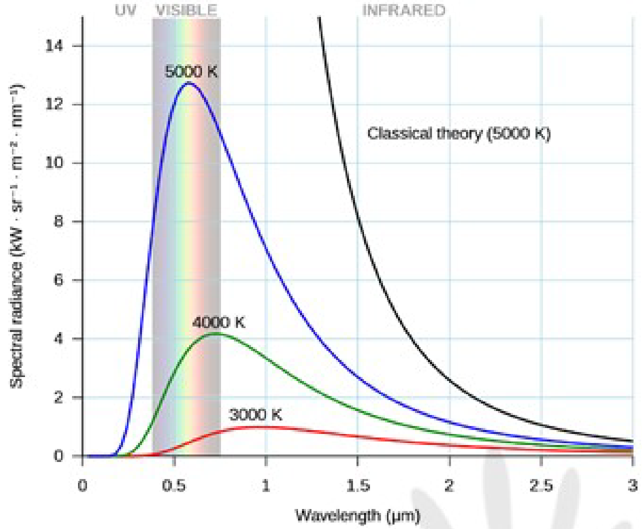
This theory could not explain the stability of an atom. According to Maxwell, accelerated electron loses its energy continuously in the form of electromagnetic radiations.

Electron (e) Negatively charged. Mass 9.1×10^{-31} Kg

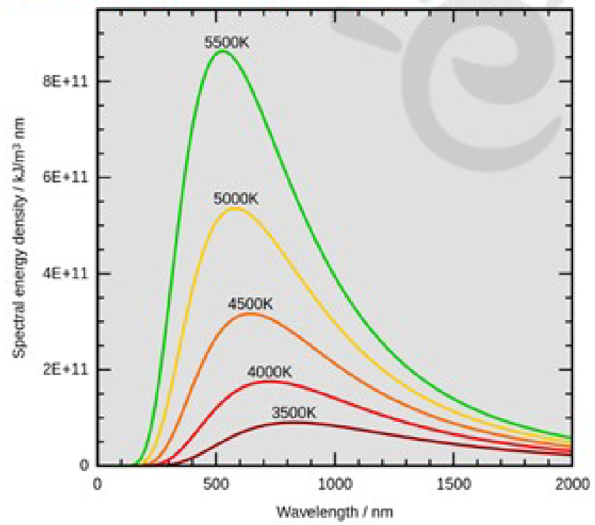
Proton (p) Positively charged. Mass 1.672×10^{-27} Kg

Neutron (n) Neutral. Mass 1.675×10^{-27} Kg

Black Body Radiation



Wein's Displacement Law



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$\lambda_{max} = \frac{b}{T}$ $\lambda T = b$

$b = 2.898 \times 10^{-3}$ m-K

$\lambda_1 T_1 = \lambda_2 T_2 = \lambda_3 T_3$

Area under the graph gives total energy

Higher temperature means more radiations

Planck's Theory

The radiant energy emitted or absorbed by a body is not radiated continuously but discontinuously in the form of small discrete packets of energy. These packets are called quantum.

$E \propto \nu$

$E = h\nu$

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

$h = 6.6 \times 10^{-34}$ Joule sec

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

$E = nh\nu$

Where n is an integer and n = number of quantum

$$E_\lambda = \frac{8\Pi hc^2}{\lambda^5} \frac{1}{e^{hc/\lambda kt} - 1}$$

Postulates of Bohr Model

Postulate 1

(a) Atom has a nucleus where all protons and neutrons are present.

Postulate 2

(a) Negatively charged electron are revolving around the nucleus in the same way as the planets are revolving around the sun. The attraction force (Coulombic or electrostatic force) between nucleus and electron is equal to the force on electron required for the circular motion

Postulate 3

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

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

Postulate 4

(a) The orbits in which electron can revolve are known as Stationary Orbits because in these orbits energy of electron is always constant.

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Atomic Structure

Postulate 5

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.

Postulate 6

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

$\Delta E = E_{\text{final state}} - E_{\text{initial state}}$

Radius of Bohr Orbits $r_n = 0.529 \frac{n^2}{Z} \text{Å}$

Energy Of Electrons

E = KE + PE

Total Energy = - KE = PE/2

$$E = -21.8 \times 10^{-19} \times \frac{Z^2}{n^2} \text{ J/atom}$$

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

Velocity Of Electron

$$v = 2.188 \times 10^6 \frac{Z}{n} \text{ m sec}^{-1}$$

Excitation Energy

Energy required to move an electron from ground state of the atom to any other state of the atom is called excitation energy of that state.

Ionisation Energy (IE)

Minimum energy required to move an electron from ground state to n = ? for hydrogen like species is called ionisation energy of the atom or ion.

Binding Energy

Energy required to move an electron from any state to n = ? is called binding energy of that state.

Emission Spectrum

When the radiation emitted is passed directly through the prism and then received on the screen then the obtained spectrum is called as emission spectrum.

Emission Line Spectrum

When any atomic gas, for example H gas, is given energy, it absorbs energy. It's electrons are excited to higher levels. When they come back to lower levels, they emit radiations. This results in formation of line spectrum

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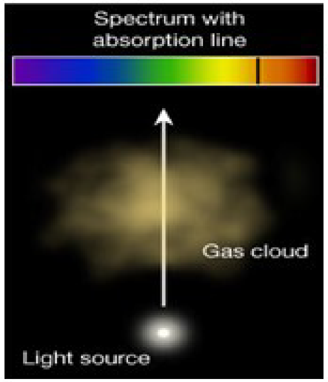
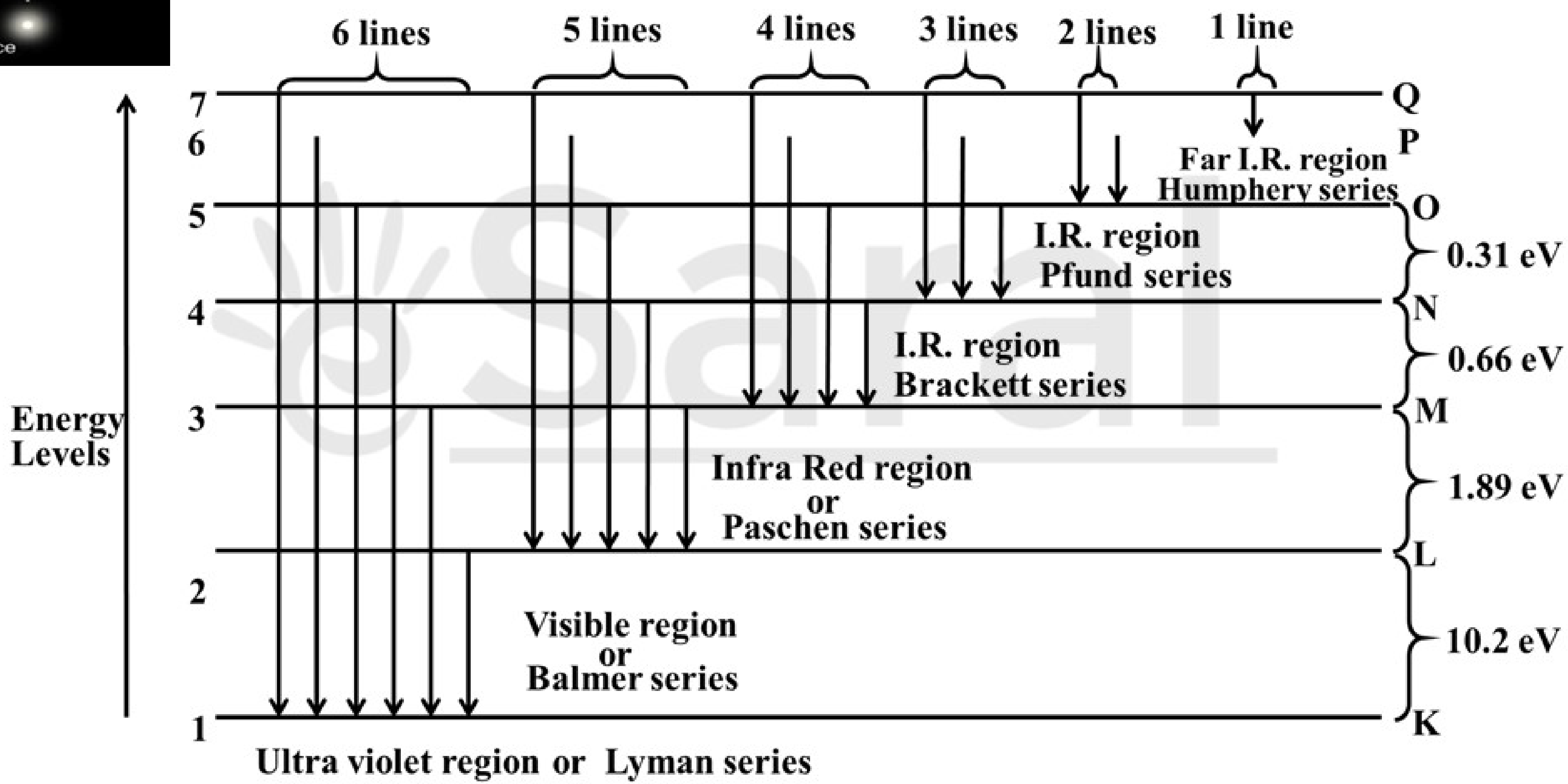
Atomic Structure



Absorption Line Spectrum



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Total no. of emission lines between

n_2 & $n_1 = \dfrac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$

$(n_2 > n_1)$

Energy During Transitions

$\Delta E = E n_2 - E n_1$

$$\frac{-2\pi^2 m K^2 Z^2 e^4}{n_2^2 h^2} - \left[\frac{-2\pi^2 m K^2 Z^2 e^4}{n_1^2 h^2}\right]$$

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

Rydberg Formula

$$\vec{v} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

value of R $= 109737.31\ cm^{-1}$

$= 10973731 m^{-1}$

$\approx 10973700\ m^{-1}$

De Broglie's Hypothesis

Like light, matter also has dual character. It exhibits wave as well as particle nature. The wavelength λ of a particle is inversely proportional to its momentum p.

$$\lambda = \frac{h}{p}$$

de-Broglie wavelength in terms of kinetic energy.

$$\lambda = \frac{h}{\sqrt{2m\,K.E.}}$$

de-Broglie wavelength when a charged particle carrying Q coulomb is accelerated by applying potential difference V then

$$\lambda = \frac{h}{\sqrt{2m\,QV}}$$

Heisenberg's Uncertainty Principle

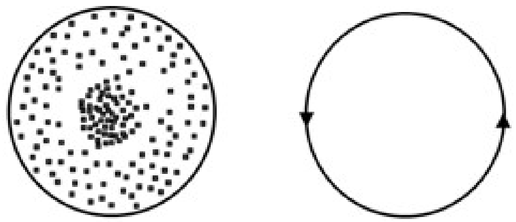
It is impossible to measure simultaneously the exact position and exact momentum of a body.

$$\Delta x . \Delta p \geq \frac{h}{4\pi}$$

$$\Delta x . m \Delta v \geq \frac{h}{4\pi}$$

$$\Delta x . \Delta v \geq \frac{h}{4\pi m}$$

These regions of space around the nucleus where probability of finding the electron is maximum are called orbitals.



Orbital
Electron-cloud representation of an orbital

Orbit
Bohr's orbit

Schrödinger's Equation

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

Where ψ = wave function of electron

m = mass of e^-

E = Total energy

V = Potential energy

ψ gives the complete description of state of system (like electron in an atom)

The ψ^2 gives probability density of finding a particle at a point in space.

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Quantum Numbers

1. Principal (n) 2. Azimuthal (l)

3. Magnetic (m) 4. Spin (s)

Principal Quantum Number (n)

It represents shell number/energy level.

It indicates the distance of an electron from the nucleus. It also determines the energy of the electron. In general higher the value of 'n', higher is the energy of a electron. Maximum electrons that a shell can accommodate $= 2n^2$

Azimuthal Quantum number (l)

It gives the name of subshells associated with the energy level and number of subshells within an energy level.

The values of l depends upon the value of 'n' and possible values are '0' to (n –1).

It also determines the energy of subshell along with n. s < p < d < f

It gives the total number of orbitals in a subshell & that is equal to (2l + 1) and number of electron in a subshell = 2 (2l + 1)

Orbital

An orbital is defined as that zone in space where electron is most likely to be found.

There are 4 type of orbitals 1. S 2. p 3. D 4. f

Maximum orbitals that a shell can accommodate $= n^2$

l	0	1	2	3
m	0	+1, 0, – 1	+2, 1, 0, –1, –2	+3, +2, +1, 0, –1, –2, –3
Possible Orientation	1	3	5	7
Orbitals	s	p_x, p_y, p_z	d_{z^2}, $d_{x^2-y^2}$, d_{xy}, d_{yz}, d_{xz}	Not in syllabus

Magnetic Quantum number (m)

The value of m depends upon the value of l and it may have integral value $-l$ to $+l$ including zero.

It gives the number of orbitals in a given subshell and orientation of different orbitals in space

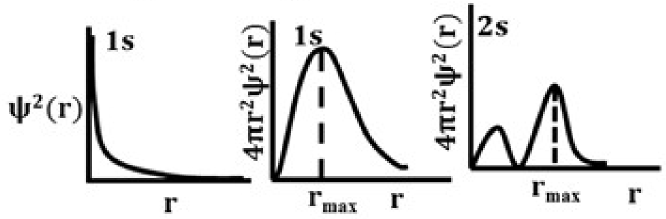
Spin Quantum number (s)

The spin quantum number represents the direction of electron spin (rotation) around its own axis (clockwise or anticlockwise).

Two spins $+\frac{1}{2}$ & $-\frac{1}{2}$
$\uparrow$ $\downarrow$

Radial Probability Distribution Function (RPDF)

Radial probability distribution =



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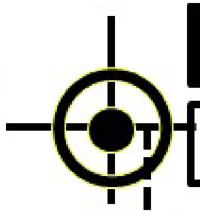
Atomic Structure



Node
The place in three dimensional space in which electron finding probability is zero.

Radial Node
Number of radial nodes = (n – l – 1).

Angular Node
Number of angular nodes = l.

Pauli's exclusion principle
No two electrons in an atom can have same values of all the four quantum numbers.

$\boxed{\uparrow\downarrow}$ K shell (n = 1), s orbital

$l = 0$, m = 0

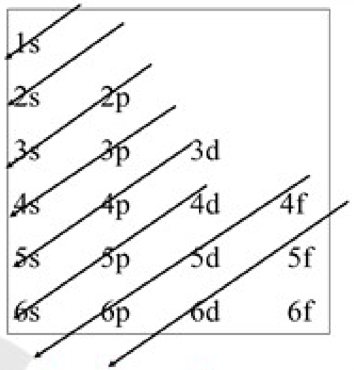
s = +1/2 and -1/2

Aufbau Principle (Means Building up)
The electrons are added progressively to the various orbitals in the order of increasing energies starting with the orbital of the lowest energy

The order of increase of energy of orbitals can be calculated from $(n + l)$ rule.

Lower the value of (n + l) for an orbital, the lower will be its energy.

If two orbitals have the same $(n + l)$ value, then orbital with lower value of n has the lower energy.



Hund's rule of maximum multiplicity
Electrons are distributed among the orbitals of a subshell in such a way as to give the maximum number of unpaired electron with parallel spins.

$\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

+1/2 +1/2 +1/2

Total spin $S = \frac{1}{2} + \frac{1}{2} + \frac{1}{2} = \frac{3}{2}$

Multiplicity $= 2x\frac{3}{2} + 1 = 4$

$\boxed{\uparrow}\boxed{\downarrow}\boxed{\uparrow}$

Total spin $S = \frac{1}{2} - \frac{1}{2} + \frac{1}{2} = \frac{1}{2}$

Multiplicity $= 2x\frac{1}{2} + 1 = 2$

Electronic Configuration
X ⟶ No of Electrons in subshell

$2p^5$ ⟶ No. of electrons in p-subshell

2 ⟶ Second shell

p ⟶ p subshell

Hydrogen n = 1 Total electrons = 1

l = 0 s orbital

$\boxed{\uparrow}$ Orbital diagram method

$1s^1$ Orbital notation method

n= 1, 2

$\boxed{\uparrow\downarrow}$ 1s

$\boxed{\uparrow\downarrow}$ 2s $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ 2p

$1s^2 2s^2 2p^3$

Exceptional Config
Chromium and Copper

Magnetic Moment
$$\mu_{effective} = \sqrt{n(n+2)}$$