STRUCTURE OF ATOM

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Structure of atom

Subatomic particles

Name	Discovery	Charge	Mass/kg
Electron (e)	Cathode rays	-1.6×10^{-19} C	9.1×10^{-31}
Proton (p)	Anode rays	$+1.6 \times 10^{-19}$ C	1.67×10^{-27}
Neutron (n)	α particles bombarded on Be thin sheet	0	1.67×10^{-27}

Atomic no. (Z)	=	Number of protons
Mass no. (A)	=	Number of protons and neutrons
Isotopes	=	Same atomic number but different mass number *e.g.* $^{12}_{6}C$ $^{14}_{6}C$
Isobars	=	Atoms with same mass number but different atomic number *e.g.* $^{14}_{7}C$ $^{14}_{6}C$

- **Photoelectric effect:** metals were exposed to beam of light.

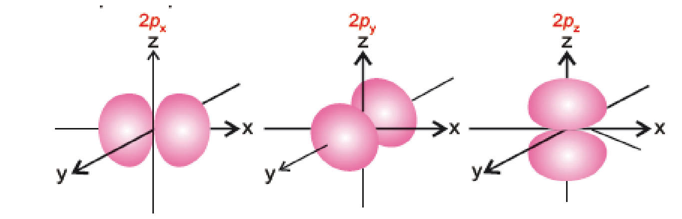
Observation of photoelectric effect

- No time lag between ejection of electrons from metal surface and striking of light beam.
- Number of ejected electrons proportional to the intensity or brightness of light.
- Minimum frequency required to eject electron is known as threshold frequency (ν_0).
- **Einstein photoelectric equation**

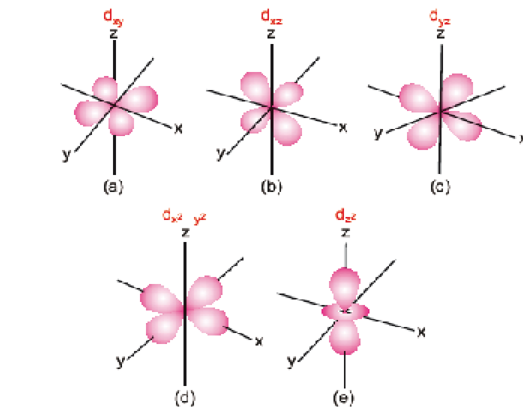
$$h\nu = h\nu_0 + \frac{1}{2}m_e v^2$$

Shapes of atomic orbitals

- Shapes of p-orbital



- Shapes of d-orbitals



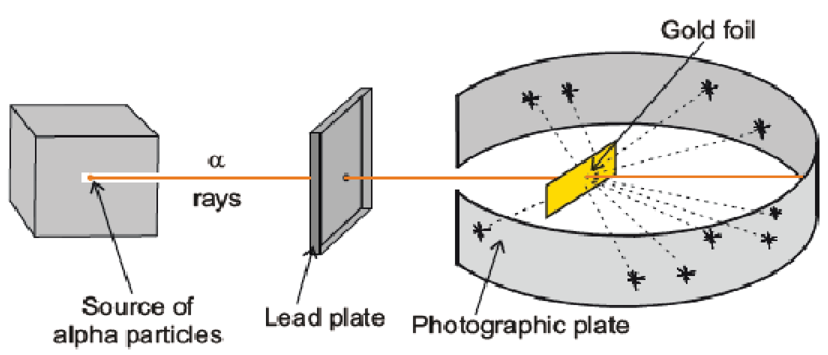
Atomic models

- **Thomson model (Plum pudding model)**

  The atom is of spherical shape in which positive charge is uniformly distributed and electrons are embedded in it

- **Rutherford's Nuclear Model**

  α particle scattering experiment



Observation of experiment
(i) Most of α ray passed through gold foil undeflected
(ii) A small fraction of the α-particles was deflected by small angles
(iii) A very few α -particles bounced back

- On the basis of Rutherford experiment most of the space in an atom is empty, a centre of atom is occupied by the nucleus in which positive charge is concentrated in a very small volume. The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular path called orbits while electrostatic forces of attraction held nucleus and electrons together.
- **Draw back of Rutherford model**

  It cannot explain the stability of atom.

Bohr's model for hydrogen atom

- Key points of Bohr's theory
- Electron in the hydrogen atom can move in circular path of fixed radius and energy known as orbits.
- The energy of orbit does not change with time.
- Electron moves from a lower stationary state to higher state when required amount of energy is absorbed by the electron.
- Electron move from higher energy state to lower energy state leaving the extra energy in the form of electromagnetic waves.
- Angular momentum of electron is quantized.

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

- Frequency of radiation absorbed or emitted

$$\nu = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$$

- $r_n = \frac{52.9(n^2)}{Z}$ pm; radius of nth orbit
- $E_n = -2.18 \times 10^{-18}\left(\frac{Z^2}{n^2}\right)$J; energy of electron in nth orbit

Electromagnetic waves

- Unlike sound wave, electromagnetic waves do not require medium and can move in vacuum.
- Electromagnetic waves are characterised by the properties, frequency (ν) and wave length (λ) and travel with speed of light *i.e.*, $c = 3 \times 10^8$ m/s $c = \nu\lambda$
- **Wave number** $(\bar{\nu}) = \frac{1}{\lambda}$

$\bar{\nu} \xrightarrow{\text{dec}}$, $\lambda \xrightarrow{\text{inc.}}$

λ	X rays	UV	Visible	IR	Micro wave	Radio wave
Rays						

- Experiment supporting wave nature of electromagnetic radiation are interference & diffraction

Planck's Quantum Law: Atoms and molecules could emit or absorb energy only in discrete quantities and not in continous manner known as quantum

E = hν
E = energy of quantum

Dual behaviour of matter

- **de-Broglie relationship** between wavelength (λ) and momentum (P) of the material particle.

$$\lambda = \frac{h}{P} = \frac{h}{mv}$$

- **Heisenberg's Uncertainty Principle**

  It states that it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron

$$\Delta x \cdot \Delta p_x \geq \frac{h}{4\pi}$$

- Heisenberg uncertainty principle is not valid for macroscopic objects.
- **Failure of Bohr model:** It ignores dual behaviour of matter but also contradicts Heisenberg uncertainty principle.

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Atomic spectra

- The spectrum of radiation emitted by a substance that has absorbed energy is called an **emission spectrum**.
- An absorption spectrum is like photographic negative of an emission spectrum.

Line Spectrum of Hydrogen

$$\bar{\nu} = 109{,}677\left(\frac{1}{n_1^2} - \frac{1}{n_f^2}\right)\text{cm}^{-1}$$

where $\bar{\nu}$ is the wave number of spectral line in hydrogen spectrum.

Series	n_1	n_2	Spectral Region
Lyman	1	2,3...	Ultraviolet
Balmer	2	3,4...	Visible
Paschen	3	4,5...	Infrared
Brackett	4	5,6...	Infrared
Pfund	5	6,7...	Infrared

Quantum mechanical model of atom

- **Orbitals and quantum number**

(1) Principal quantum number 'n' determines the size and energy of orbital.
- l Number of allowed orbital in a shell = n^2

(2) Azimuthal quantum number 'l' defines the three-dimeinsional shape of orbital
- For a given n, possible value of l = 0, 1, 2 ... (n – 1)

Value of l	0	1	2	3	4	5
Notation of subshell	s	p	d	f	g	h

(3) **Magnetic orbital quantum number 'm'** gives linformation about the spatial orientation of the orbital with respect to standard set of coordinate axis.
- For any subshell, 2l + 1 values of m are possible

Subshell	s	p	d	f	g	h
Notation of orbitals	1	3	5	7	9	11

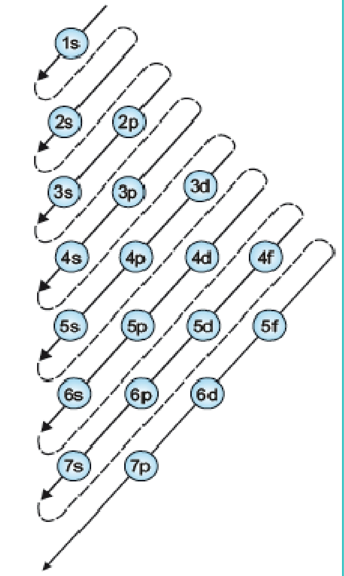
(4) Two orientations of electrons are distinguished by the **spin quantum numbers (m_s)** which can take value of

$+\frac{1}{2}$ and $-\frac{1}{2}$.

Filling of orbitals in atom

- **Aufbau Principle:** In the ground state of the atoms, the orbitals are filled in order of their increasing energies.
- The maximum number of electrons in 2a shell = $2n^2$
- **Hund's Rule of Maximum Multiplicity**

  Pairing of electrons in degenerate orbitals take place only after each degenerate orbitals is singly filled.



- **Pauli Exclusion Principle**

  No two electrons in an atom can have the same set of four quantum numbers.

- **Total nodes** = n – 1, angular nodes = l, radial nodes= n – 1 – l

Energies of Orbitals

- 1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f <(for hydrogen)
- **(n + l) rule** – the lower the value of (n + l) for an orbital, the lower is its energy. It two orbitals have the same value of (n + l), the orbital with lower value of n will have the lower energy for multielectron atom.
- Energies of the orbital in the same subshell decreases with increase in the atomic number (Z_{eff})

  e.g. $E_{2s}(H) > E_{2s}(Li) > E_{2s}(Na) > E_{2s}(K)$