Class 11, Chapter- 2: Structure of Atom

1. Dalton atomic theory:

2. Atomic models:

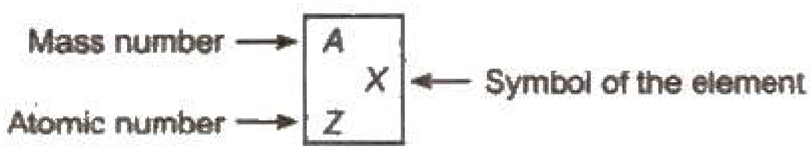
a) Thomson's atomic model:

b) Rutherford atomic model:

3. Subatomic particle of atoms:

PARTICLE	ELECTRON	PROTON	NEUTRON
Discovery	Sir. J. J. Thomson (1869)	Goldstein (1886)	Chadwick (1932)
Nature of charge	Negative	Positive	Neutral
Amount of charge	1.6×10^{-19}Coloumb	1.6×10^{-19}Coloumb	0
Mass	9.11×10^{-31}kg	1.672614×10^{-27}kg	1.67492×10^{-27}kg

4. Atomic number (Z) and Mass number (A):



5. Isotopes, Isobar and Isotone:

- **Isotopes** Species with same atomic number but different mass number are called isotopes, e.g. $_1H^1$, $_1H^2$.
- **Isobars** Species with same mass number but different atomic number are called isobars. e.g., $_{18}Ar^{40}$, $_{19}K^{40}$.
- **Isotones** Species having same number of neutrons are called isotopes, e.g., $_1H^3$ and $_2He^4$
- **Velocity** of light = frequency x wavelength ; $c = \nu\lambda$
  (also ν = 1/Period or duration in Second)

6. Electromagnetic waves:

7. Electromagnetic spectrum:

8. Planck quantum theory (Planck Equation):

$E = nh\upsilon$ (Where n=1, 2,3,4,..& h= Planck's constant =6.626 x 10^{-34} Js)

9. Photo electric effect: KE = E - Wo $\therefore 1/2\ mv^2 = h(\upsilon - \upsilon_o)$

Where; υ = frequency of incident radiation
Threshold frequency (υ_o) = minimum frequency require for removal of the electron.
Work function (Wo) = required minimum energy for removal of electron.

10. Bohr's atomic model:

11. Application of Bohr's atomic model:

a) Radius of the *n*th orbit:

$r_n = n^2h^2 / 4\pi^2me^2Z$

$r_n = .0529\ nm\ [n^2\ / Z]\ nm$

b) Velocity of electron in *n*th orbit:

$V_n = 2\pi e^2Z / nh$

$V_n = 2.165 * 10^6\ [Z / n]$ meter per second

c) Energy of an electron in *n*th orbit:

$E_n = (-)\ 2\pi^2mk^2Z^2e^4 / n^2h^2$

$E_n = (-)\ 2.18 * 10^{-18}\ [Z^2 / n^2]$ Jule

(1 Jule = 10^7 ergs; 1ev = 1.602 x 10^{-19}Jule)

$E_n = (-)\ 13.6\ [Z^2 / n^2]$ ev

When electron is jump from higher energy level to lower energy level its Emitted / release energy (show by (–) minus sign). It can be calculated as-

$$\Delta E = E_2 - E_1 = (+)\ 13.6\ Z^2\ [1/n_1^2 - 1/n_2^2]\ ev$$

And when jump from lower to higher energy level its absorbed energy. It can be calculated as-

$$\Delta E = E_1 - E_2 = (+)\ 13.6\ Z^2\ [1/n_2^2 - 1/n_1^2]\ ev$$

12. Atomic spectrum (Emission and Absorption spectrum):

13. Hydrogen Spectrum:

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

Calculation number of spectral line: = $(n_2 - n_1)$ **[**$(n_2 - n_1)$ **+ 1] / 2**

For calculate wavelength of hydrogen spectrum line we can use following formula-

$$\bar{\nu} = 109,677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \ cm^{-1}$$

R = Rydberg's constant = $2\pi^2 mk^2 Z^2 e^4 / ch^3$ = 109677 cm-1

14. Limitations of Bohr's theory:

15. Bohr – Sommerfield model:

16. Dual Behaviour of Matter (de Broglie equation):

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

Where *p* is the momentum of particle = *mv*.

Confirmation of de- Broglie equation: The above relation can be confirmed by using Einstein energy equation and Planck equation as

$E = mc^2$ (Einstein energy equation) ------- (1)

$E = h\upsilon$ (Planck energy equation) ------- (2)

From eq. (1) and (2)

$mc^2 = h\upsilon$

$mc^2 = hc/\lambda$

$\lambda = h / mc$

$\lambda = h / p$

17. Relation b/w kinetic energy and wave length:

$\lambda = h / 2\ KE^*\ m$

18. Heisenberg's Uncertainty Principle:

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

Δx = uncertainty in position,

Δp = uncertainty in momentum

19. Schrodinger wave function:

20. Quantum numbers:

- **Principal quantum number (n):** It tells us about shell.

N	1	2	3	4
Shell no.:	K	L	M	N
Total number of orbitals in a shell = n^2	1	4	9	16
Maximum number of electrons = $2n^2$	2	8	18	32

- **Azimuthal quantum number (l):** Azimuthal quantum number. 'l' is also known as orbital angular momentum or subsidiary quantum number. l. It identifies sub-shell, determines the shape of orbitals, energy of orbitals. The number of orbitals in a subshell = $2l + 1$. For a given value of *n*, it can have *n* values ranging from 0 to n-1.

Subshell notation	s	p	d	f	g
Value of '*l*'	0	1	2	3	4
Number of orbitals	1	3	5	7	9

- **Magnetic Quantum Number (m):** It tells about the number of orbitals and orientation of each subshell. Value of m = – l to + 1 including zero like -2, -1, 0, +1, +2 (for d orbital).

- **Spin Quantum Number** (s): It indicates the direction of spinning of electron, i.e., clockwise or anti- clockwise.

- **Difference between Orbit and Orbitals:**

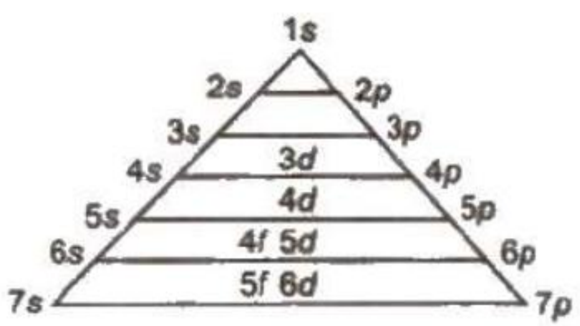
	Orbit	Orbital
1.	An orbit is a well defined circular path around the nucleus in which the electron revolves.	An orbital is the three dimensional space around the nucleus within which the probability of finding an electron is maximum.
2	The maximum number of electrons in any orbit is given by $2n^2$ where n is the number of the orbit.	The maximum number of electrons present in any orbital is two.

21. Shape of orbitals:

22. Nodes, Nodal plane (Angular nodes) and Nodal surface (Radial nodes):
Nodes = (n-1) Here n = Number of shell
Nodal plane (Angular nodes) = l
Nodal surface (Radial nodes) = (n – l – 1),   Here l = Number of orbitals

23. Rules for filling of electron:
a) Aufbau principle:
b) Filling of electron in orbitals- (n + l) rule: The orbitals are filled in order of increasing value of n + l. For the orbitals having the same value of n + 1, the orbital having lower value of n is filled up first. The general order of increasing energies of the orbital is

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



c) Hund's maximum multiplicity principle:

d) Pauli's exclusion principle: