3.7 ATOMIC STRUCTURE

https://subarnam.com.np/

Syllabus [6 Hrs.]
· Introduction
· Concept of subatomic particles like electron, proton and neutron concerning their charge, mass and location in an atom (no explanation of the cathode ray experiment is required)
· Rutherford's α - rays scattering experiment and its observations
· Rutherford's atomic model and its drawbacks
· Postulates of Bohr's atomic model
· Atomic number and mass number
· Isotopes, isobars and isotones
· Bohr-Bury scheme
· Aufbau principle
· Electronic Configuration of atoms (atomic number 1-30)
· Hund's rule of maximum multiplicity
· Quantum numbers and their types (principal, azimuthal, magnetic and spin)
· Pauli's exclusion principle

Concept of Subatomic Particles

Atoms are mostly empty space surrounding a very small, dense nucleus that contains protons and neutrons; electrons are found in shells within this empty space.

Particle	Symbol	Mass (g)	Mass (amu)	Charge
Electron	e^-	9.1×10^{-28}	0.00055 (nearly 0)	-1.602×10^{-19} C (−1)
Proton	p^+	1.672×10^{-24}	1.0073 (nearly 1)	$+1.602\times10^{-19}$ C (+1)
Neutron	n^0	1.675×10^{-24}	1.0087 (nearly 1)	0

Key Takeaway: Atoms are composed of three fundamental particles—electrons (negative), protons (positive), and neutrons (neutral). Protons and neutrons form the dense nucleus, while electrons orbit around it.

1. Write short notes on: Sub-atomic particles 2

Rutherford's atomic model

The Experiment (1911):

Ernest Rutherford bombarded a very thin **gold foil** (thickness ~10^{-5} cm) with **α-particles** (helium nuclei, **He^{2+}**) from a radioactive source. Scattered **α-**particles were detected on a circular **ZnS screen** that produced tiny flashes of light (scintillations).

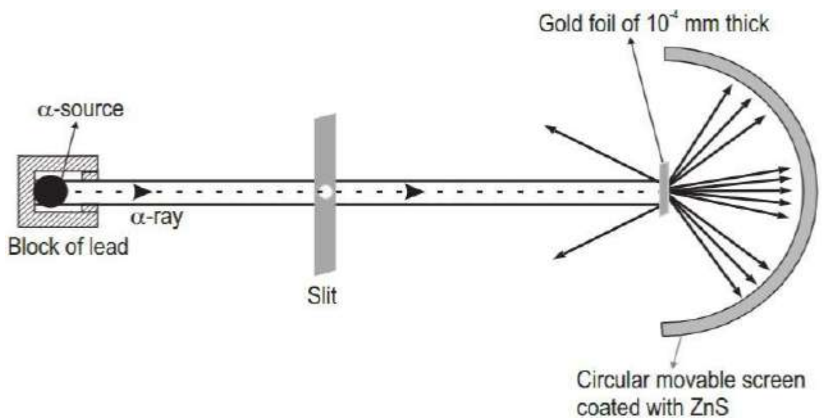


Fig: α-particles scattering experiment of Rutherford

Key Observations:

Observation	What This Tells Us
≥99% of α-particles passed straight through	An atom is mostly empty space
Some deflected at small angles	A small, dense positive region exists inside
Very few (~1 in 10,000) bounced back	The nucleus is extremely small but very massive

Based on the above conclusions, Rutherford put forward the following postulates for his model of the atom, called the Nuclear Model.

1) **Positive nucleus:** A very small, dense, positively charged nucleus sits at the centre and carries almost all the mass.
2) **Mostly empty:** The rest of the atom is largely empty space.
3) **Neutral atom:** The number of electrons outside equals the positive charge of the nucleus.
4) **Revolving electrons:** Electrons revolve around the nucleus, dispersed in the space around it.
5) **Force balance:** Electrostatic attraction between the nucleus and electrons is balanced by the centrifugal force of revolution.

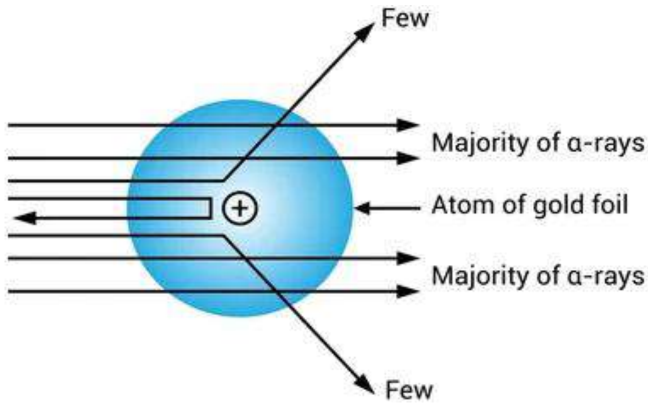
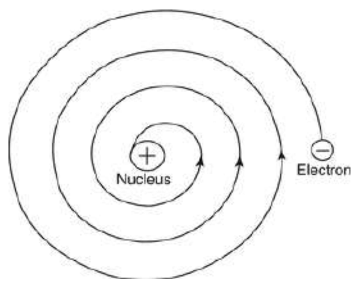


Figure: Explanation of α-particles scattering experiment.

💡 **Remember:**
Most α-particles passed through because atoms are mostly empty space.

Drawbacks of Rutherford's Atomic model

1. ⚠ **The stability of the atom**: According to classical electrodynamics, a moving charged body (electron) loses energy continuously in the field of another charge. If so, the electron should spiral inward and crash into the nucleus — but atoms are stable! Rutherford's model could not explain why atoms don't collapse.



2. ⚠ **Atomic spectra**: Classical physics predicts a continuous spectrum as the electron loses energy gradually. However, hydrogen (and other atoms) produce only **discrete lines** of specific wavelengths. Rutherford's model had no explanation for this

phet.colorado.edu/en/simulations/rutherford-scattering

1. Uranium changes into which element by releasing alpha particles?
   a) Thorium b) Lead c) Radium d) Radon

2. Explain Rutherford's atomic model with the α-ray scattering experiment, with a labelled diagram. [5]
   Explain Rutherford's atomic model with its alpha -ray scattering experiment. [3]
3. Write down the Rutherford atomic model in brief. [2]
   What conclusion did Rutherford make from his α-ray scattering experiment about the structure of the atom? [2+3]
4. Write the postulates of Rutherford's atomic model of the atom. Also, mention its drawbacks. [3+1]
   Explain Rutherford's atomic model and its limitations.
   Explain Rutherford's atomic model and its drawbacks. [4]

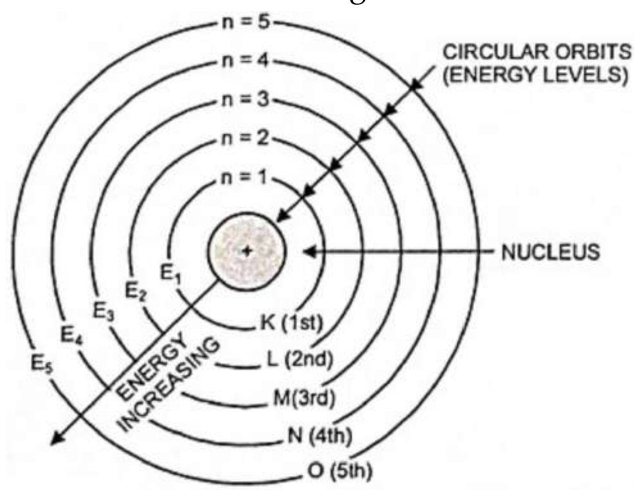
State the drawbacks of Rutherford's atomic model. 2
Write the drawbacks of Rutherford's atomic model.

Bohr's Atomic Model

Niels Bohr (1913), proposed the following postulates based upon Planck's quantum theory to overcome the shortcomings of Rutherford's atomic model.

The main postulates:

1. **Fixed orbits:** Electrons revolve around the nucleus in fixed concentric paths called orbits, shells, or energy levels.
2. **Stationary Orbits:** As long as an electron remains in a particular shell, it neither loses nor gains energy. It means the energy of the electrons is constant; hence, these orbits are also called **stationary orbits**.
   The energies of electrons increase with increasing distance from the nucleus.



3. **Quantized angular momentum**: Only those orbits are possible where the angular momentum of the electrons is an integral multiple of $h/2\pi$. Or the angular momentum of the electrons is *quantised*.

   i.e., angular momentum **(mvr)** = $\frac{nh}{2\pi}$ ………… (1)

   where, m = mass of electron

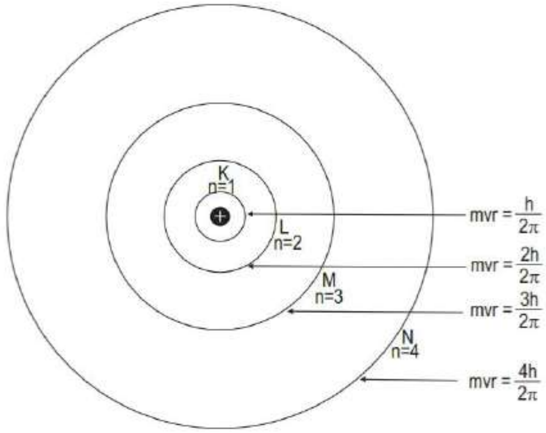
   r = radius of orbit

   v = linear velocity of the electron

   n = an integer representing the orbits

   (n = 1, 2, 3, ……, ∞)

   h = Planck's constant (6.626×10^{-34} Js)



This postulate explains the stability of atoms under ordinary conditions.

4. **Attraction of nucleus & electrons:** The electrostatic force of attraction between the nucleus and an electron is equal to the centrifugal force of the electron. The electrostatic force of attraction provides the necessary **centripetal force** required to keep the electron in circular motion.

   Mathematically, $\frac{1}{4\pi\mathcal{E}o}\frac{e^2}{r^2} = \frac{mv^2}{r}$ ………………(2)

   here, $\mathcal{E}o$ = permittivity of the vacuum

   e = charge of electron (1.602 x 10^{-19} C)

   m = mass of an electron v = velocity of the electron

   r = radius of the orbit (distance from the nucleus)

5. **Emission or absorption of energy:** The electrons emit or absorb energy in the form of photons (or quanta) only when they jump from one energy level to another.

The energy absorbed or radiated equals the difference between the energy of an atom's lower and higher energy levels.

$$\Delta E = E_2 - E_1 = h\upsilon$$

Where ΔE = difference in energy between the higher and the lower energy levels ΔE = energy of radiation emitted or absorbed

hυ = energy of radiation

h = Planck's constant

υ = frequency of radiation (pronounced as 'nu')

E_2 and E_1 = energy of an electron in a higher and lower orbit

The energy is absorbed when an electron jumps from a lower orbit to a higher orbit, and it is radiated when an electron jumps from a higher orbit to a lower orbit.

This postulate explains the origin of atomic spectra.

Fixed orbit → No energy loss

5. Explain the main postulates of Bohr's atomic model. [5]
   Mention the postulates of Bohr's atomic model. [4]
   What are the main postulates of Bohr's atomic model? Explain. [4]
   What are the postulates of Bohr's atomic model? [4]

Write the postulates of Bohr's atomic model. [3]
Explain the postulates of Bohr's atomic model.
What are the basic postulates of Bohr's atomic Model? [3]
Write the postulates of Bohr's atomic model. [2]
Write three postulates of Bohr's atomic model. [3]

6. What is an orbit? 1

Atomic number and mass number

Isotopes, isobars and isotones

1. Write short notes on: Isotopes and isobars 2

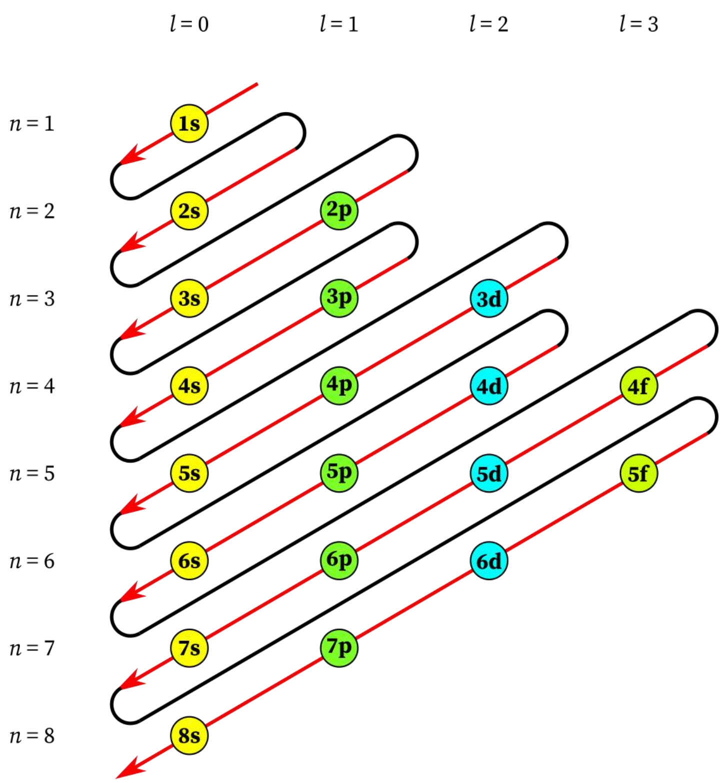
Electronic Configuration

Aufbau is a German word meaning building up or construction. This principle states, "The orbitals are filled with electrons in increasing order of their energy."

The principle may also be explained by the (n+l) rule. According to this rule,

1. The orbitals with the lower value of n+l are filled first.
2. When two orbitals have the same value of n+l, the orbital with a lower value of n is filled first.

The following device diagram can represent the sequence of filling of various subshells according to the Aufbau principle.



So, the electrons are filled in the following order;
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d,………..

Limitations of the Aufbau Principle (Exceptions)

Cr (chromium) (atomic number 24) has an electronic configuration of $1s^2\,2s^22p^6\,3s^23p^63d^5\,4s^1$ instead of $1s^2\,2s^22p^6\,3s^23p^63d^4\,4s^2$ because the half-filled d orbital is relatively more stable due to symmetry.

Similarly, Copper (29) has E.C. $1s^2\,2s^22p^6\,3s^23p^63d^{10}\,4s^1$ instead of $1s^2\,2s^22p^6\,3s^23p^63d^9\,4s^2$, because a fully filled d-orbital is relatively more stable.

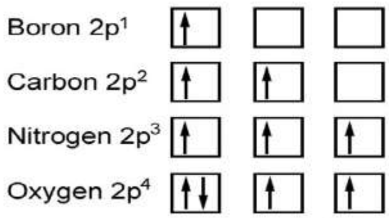
2. Write short notes on: Aufbau principle 2
   State Aufbau principle. 2
3. Write the electronic configuration of chromium (atomic number 24). [2]
   Write the electronic configuration of Cr in terms of s, p, d, and f. [1]
4. Write the electronic configuration of copper. [1]
5. Write the electronic configuration of iron (Fe), chromium (Cr) and zinc (Zn). [3]

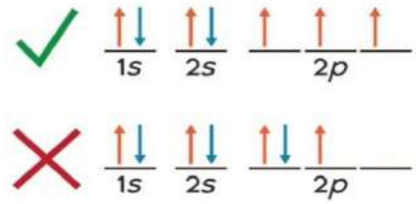
Hund's Rule of Maximum Multiplicity

It states: The pairing of electrons will not occur in any orbital until and unless all the available *degenerate orbitals* have one electron each with *parallel spin.*

This rule is based on the fact that negatively charged electrons repel each other. If two electrons move in the same region of space, e.g., a $2p_x$ orbital, they will repel each other more strongly than if they remain in different regions of space (e.g., $2p_x$ and $2p_y$ orbitals). Thus, energetically, it is more favourable for them to go into different orbitals.

Illustration with examples:





6. Write short notes on: Hund's rule 2.5

Quantum Numbers

Electrons are distributed in atoms in different shells represented by n = 1, 2, 3, 4, etc. These are called principal quantum numbers. Bohr had introduced them. Sommerfield introduced another quantum number called the azimuthal, or Subsidiary, quantum number. The wave mechanical treatment of atoms reveals that these numbers are not arbitrary. Now, it is clear that 4 sets of quantum numbers are needed to characterise an electron completely.

Quantum numbers can be defined as a set of numbers that give the energy, size, shape, and orientation of orbitals of electrons in an atom, as well as the spin of an electron.

1. **Principal Quantum number (n)**

- Introduced by Bohr's atomic model to ***represent the orbits***.
- Has positive integer values n = 1, 2, 3, etc., which represent K, L, and M shells, respectively
- Determines the average ***distance of the electron from the nucleus*** and the average ***energy of an electron***.
- Helps to determine the maximum number of electrons that can be accommodated in a shell by *the $2n^2$ rule.*

n value	Shell	Max electrons ($2n^2$)	Subshells present
1	K	2	1s
2	L	8	2s, 2p
3	M	18	3s, 3p, 3d
4	N	32	4s, 4p, 4d, 4f

2. **Azimuthal quantum number (l)/ Subsidiary quantum number**.

- Mainly determines the *orbital angular momentum of an electron.*
- *Signifies the subshell* to which the electron belongs.
- Helps to find the number of subshells in a given shell.
- Helps to determine the maximum number of electrons that can be accommodated in the given subshell by the *2(2l+1) rule*.

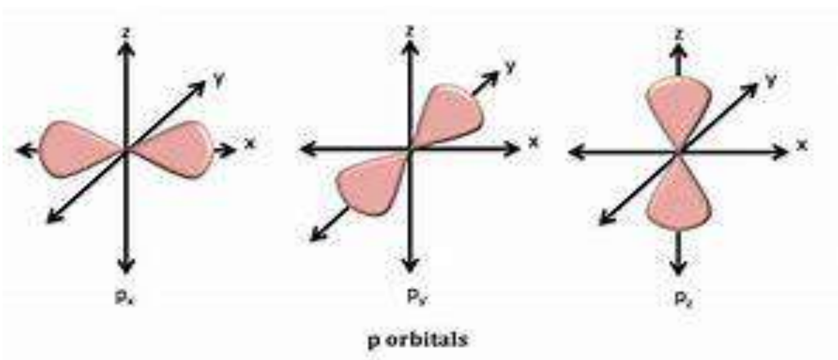
➢ The Azimuthal quantum number may vary from ***0 to (n-1) for a given principal quantum number.***

n	1		2		3		4			
l	0	0	1	0	1	2	0	1	2	3
subshell	s	s	p	s	p	d	s	p	d	f

The azimuthal quantum number represents the relative energies of subshells and their shapes. The energies of the subshells are in the following order: s < p < d < f

3. **Magnetic Quantum Number (m)**

➢ Describes the *behaviour of electrons in a magnetic field*.
➢ Signifies the ***orientation of the orbital*** in space in the presence of an external magnetic field.
➢ Or it *specifies the orbital* in which an electron lies.
➢ This successfully explains the ***Zeeman effect***.



p orbitals

For a given Azimuthal quantum number, l, the magnetic quantum number ranges from –l to 0 to +l.

l value	m values	No. of orbitals	Subshell
0	0	1	s (1 orbital)
1	−1, 0, +1	3	p (p_x, p_y, p_z)
2	−2, −1, 0, +1, +2	5	d (5 orbitals)
3	−3 to +3	7	f (7 orbitals)

4. **Spin quantum numbers (s)**

➢ Determines the spin angular momentum or the direction of the spin of an electron around its axis.
➢ It can have only two values, +1/2 and -1/2, representing the electrons' clockwise and anticlockwise rotation.
➢ Two electrons in an orbital always have opposite spins. These are also represented by ↑ & ↓ or ↿ & ⇂ .

Shell	**Principal quantum number(n)**	**Azimuthal quantum number(l) [0- (n-1)]**	**Magnetic quantum number(m) (-l to +l)**
K shell	n=1	0 (s-subshell)	0 (s-orbital)
L shell	n=2	0 (s-subshell) 1 (p-subshell)	0 (s- orbital) -1 (p_x-orbital) 0 (p_y-orbital) +1 (p_z-orbital)
M shell	n=3	0 (s-subshell) 1(p-subshell 2 (d-subshell)	0 (s orbital) -1 (p_x-orbital) 0 (p_y-orbital) +1 (p_z-orbital) +2 ($d_{x^2-y^2}$ orbital) +1 (d_{xz} orbital) 0 (d_{z^2}-orbital) -1(d_{yz} orbital) -2(d_{xy} orbital)

7. Explain quantum numbers in detail. [4]
   What do you mean by quantum numbers? Explain.
   Write short notes on: Quantum number

8. What is the azimuthal quantum number? [2]

9. Write any two differences between orbit and orbital. 2

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