ATOMS

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Atoms

Atomic structure

For quite some time, matter was thought to be continuous but not discrete. Later Dalton, Avagadro, Thomson etc. made considerable contributions to understand the structure of matter. However, the first reasonably successful model was proposed by Rutherford based on his experiments on scattering of α-particles.

Rutherford's α-particle scattering experiment

A narrow collimated beam of α-particles from radioactive bismuth ($_{83}Bi^{214}$) source was directed against a thin gold foil (thickness $\approx 2 \times 10^{-7}$ m). The angular distribution of the scattered α-particles was measured using a detector. The detector consisted of a zinc sulphide screen and a microscope. Each time an α-particle hits the zinc sulphide screen, a flash of light is seen.

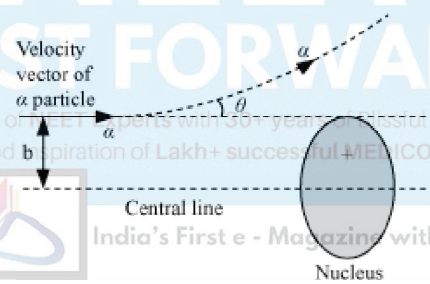
Observations

- Most of the α-particles went undeviated. A very small percentage of α-particles (roughly 0.14%) deflected through an angle of more than 1°.
- An insignificantly small number of α-particles are deflected by almost 180°, as if they bounced back. (Approximately one in 8000 α-particles deflected through more than 90°)

Explanation for observations

- An α-particle is so massive compared to the mass of an electron (roughly 7350 times) that most of the α-particles could pass through the foil undeflected.
- Only if we assume a concentration of complete positive charge in a very small space inside a gold atom, then the coulomb force of repulsion could be large enough to cause a bounce bank of an incident α-particle.
- Also, the passage of a large number of α-particles undeflected, is possible if almost the entire mass of the atom is confined to a very small region of space.
- This tiny central core of the atom which contains +ve charge and almost complete mass of the atom (99.95%) was named by Rutherford as *atomic nucleus*.

Impact parameter



The **impact parameter is the perpendicular distance of the initial velocity vector of the α-particle from the centre of the nucleus when it is far away from the atom.**

According to the theory of Rutherford's α-scattering experiment, impact parameter (b) of an α-particle of kinetic energy E_k, scattered at an angle θ is given by

$$b = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{E_k}\cot\left(\frac{\theta}{2}\right) = \frac{1}{4\pi\varepsilon_0}\frac{2Ze^2}{mv^2}\cot\left(\frac{\theta}{2}\right) \qquad \left(\text{since } E_k = \frac{1}{2}mv^2\right)$$

Distance of closest approach (size of the nucleus)

At this distance, the kinetic energy of the α-particle is transformed into electrostatic potential energy. Hence,

$$K = U$$

$$\frac{1}{2}mv^2 = \frac{1}{4\pi\varepsilon_0}.\frac{(2e)(Ze)}{r_0}$$

$$r_0 = \frac{1}{4\pi\varepsilon_0}.\frac{4Ze^2}{mv^2}$$

The distance of closest approach is of the order of 10^{-14} metre.

Rutherford concluded that

1. much of the space in an atom is empty.
2. the entire positive charge of the atom is concentrated at a very small region at its centre (the region is called nucleus).

3. since, the atom as a whole is electrically neutral, the negatively charged particles (electrons) revolve round the nucleus in circular orbits (the electrons are continuously accelerated towards the centre).

   But according to well established ideas of classical electrodynamics, an accelerated charge must radiate energy. Therefore a revolving electron must radiate energy continuously and hence move in a helical orbit and finally collapse into the nucleus. Thus Rutherford's model fails to provide the picture of a **stable atom**. Also, the Rutherford's model does not explain the spectra of an atom.

Bohr's atom model

Neils Bohr modified Rutherford's model, which could explain

(a) the stability of an atom

(b) the spectral series of hydrogen atom.

Bohr's theory is applicable to hydrogen and hydrogen-like atoms only. *Example*: singly ionised helium, doubly ionised lithium etc.

The model is based on the following postulates

1. An electron can revolve round a nucleus in an orbit called a **stationary orbit** – from which no energy is radiated and the angular momentum of the electron is an integral multiple of $(h/2\pi)$

   i.e., $mvr = \frac{nh}{2\pi}$ where $n = 1, 2, 3 \ldots$

   m = mass of electron,

   v = orbital speed

   r = orbital radius

   h = Plank's constant

2. When an electron jumps from a higher energy orbit (E_2) to a lower energy orbit (E_1), the difference in the energy is emitted as a single packet of energy called a quantum or a photon.

   i.e., $E_2 - E_1 = h\nu$ where ν is the frequency of photon.

List of expressions

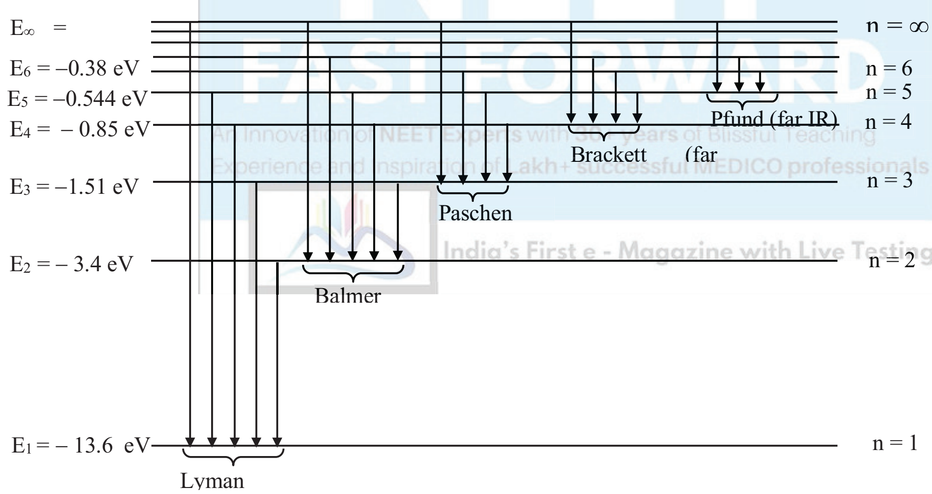
Bohr's quantisation rule	$mvr = \frac{nh}{2\pi}$, $n = 1, 2, 3 \ldots$	
Radius of the n^{th} orbit	$r = \frac{\epsilon_0 n^2 h^2}{\pi m Ze^2}$	$r = \left(\frac{n^2}{Z}\right) 0.53\ A°$
Speed of the electron in n^{th} orbit	$v_n = \frac{Ze^2}{2nh\varepsilon_0}$	$v = \left(\frac{Z}{n}\right) 2.16 \times 10^6\ ms^{-1}$
Kinetic energy of the electron in n^{th} orbit	$K.E = \frac{1}{2}\left(\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r}\right) J$	$KE = \left(\frac{Z^2}{n^2}\right) 13.6\ eV$
Potential energy of the electron	$P.E = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r} J$	$PE = -2\left(\frac{Z^2}{n^2}\right) 13.6\ eV$
Total energy of the electron	$E = \frac{-mZ^2e^4}{8\varepsilon_0^2 n^2 h^2} J$	$TE = -\left(\frac{Z^2}{n^2}\right) 13.6\ eV$
Wavelength of emitted radiation and Wave number	$\therefore \nu = \frac{1}{\lambda} = R\ Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$	
Rydberg constant	$R_H = \frac{me^4}{8\varepsilon_0^2 ch^3} = 1.097 \times 10^7\ m^{-1}$	
Note: $PE = 2\,(TE)$; $KE = \lvert TE \rvert$; $KE = \frac{\lvert PE \rvert}{2}$		

- As n increases, the numerical value of $[1/n^2]$ decreases. Since it is with a negative sign, its negativeness decreases. Therefore, the energy actually increases with the order number.
- The energy of an electron remains constant as long as it remains in a particular orbit.
- If an electron absorbs a photon of frequency ν such that $h\nu = E_2 - E_1$, it transits. Conversely if an electron jumps from a higher energy orbit to a lower energy orbit, it emits a photon of frequency ν such that $E_2 - E_1 = h\nu$.
- Here E_1 represents energy corresponding to a lower energy orbit and E_2 represents energy corresponding to higher energy orbit.
- Since n can take only integral values, r, v and E all can have only certain discrete values i.e., radius, orbital velocity (hence angular velocity) and energy are all quantised.
- n is also called principal quantum number.
- When n = 1, the energy is minimum. This state is called ground state. For n = 2, 3, 4,.... the atom is said to be in excited state.
- A spectral line is the result of an electron transition from higher energy state to lower energy state.

Energy level diagram

To understand the energy states of an atom and the electronic transitions, an energy level diagram is constructed. The diagram shown is for hydrogen.

Horizontal lines are drawn to represent the energy of an electron in different states. Vertical lines are drawn with arrow marks to indicate transitions. Arrows downwards as shown in the figure represent emission, while upward arrows represent absorption.



Spectral series in the hydrogen spectrum

Series	n_1	n_2	Region in the spectrum
Lyman	1	2, 3, 4, ...	Ultraviolet
Balmer	2	3, 4, 5	visible region [prominent lines are H_α (red) H_β, H_γ H_δ (violet)]
Paschen	3	4, 5, 6	Infra red
Bracket	4	5, 6, 7	Infra red
Pfund	5	6, 7, 8	Far infra red