CHAPTER 12

Atoms

Rutherford's Atomic Model, Bohr's Model & Energy Level Diagram

Introduction

- Atoms in simple terms are defined as the smallest unit of matter.
- Atoms are electrically neutral because they contain same number of electrons and protons.

Plum-Pudding Model

- In 1898, J. J. Thomson proposed the first model of atom.
- He stated, there is a uniform distribution of the positive charge of the atom throughout the volume of the atom and like seeds in a watermelon, the negatively charged electrons are embedded in it. This model was picturesquely called plum pudding model of the atom.

Alpha-Particle Scattering

- Rutherford used a "Gold foil experiment"
- Rutherford only identified one of type of radiation given off by radioactive elements like polonium, uranium and named them as alpha particles.
- The alpha particles are fast moving and positively charged Helium nuclei with two protons and two neutrons.

Rutherford observed the deflection of alpha particles after passing through metal sheet and proposed his atomic model

- After passing through the metal sheet, the alpha particles strike on fluorescent screen which was coated with zinc sulphide and produced a visible flash of light
- He concluded that an atom consists of a minute positively charged body at its center called as nucleus. The nucleus, though small, contains all the protons and neutrons.

Alpha-Particle Trajectory

- The trajectory traced by an α particle depends on the impact parameter, b of collision.
- The particle near to the nucleus suffers large scattering.
- Only a small fraction of the number of incident particles rebound back indicating that the number of α-particles undergoing head on collision is small.

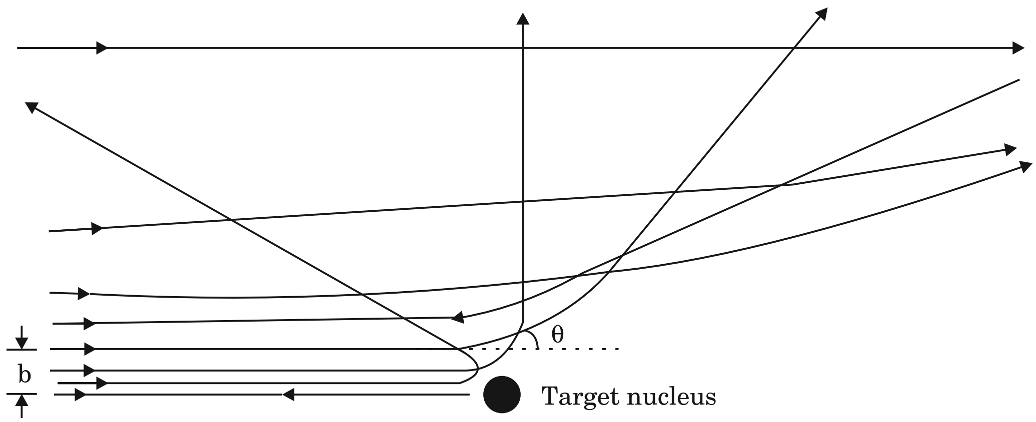

Fig.: Alpha-Particle Trajectory

Rutherford's nuclear model of Atom

- According to Rutherford's model, the entire positive charge and most of the mass of the atom is concentrated in a small volume called the nucleus with electrons revolving around the nucleus just as planets revolve around the sun.
- Rutherford scattering is a powerful way to determine an upper limit to the size of the nucleus.
- **Drawbacks of Rutherford's model:** There were two major drawbacks in Rutherford nuclear model in explaining the structure of atom:

  It cannot explain the characteristic line spectra of atoms of different elements.

  It contradicts the stability of matter because it speculates that atoms are unstable because the accelerated electrons revolving around the nucleus must spiral into the nucleus.

Electron Orbits

- The electrostatic force of attraction, F_e between the revolving electrons and the nucleus provides the requisite centripetal force (F_c) to keep them in their orbits. Hence, for a dynamically states orbit in a hydrogen atom $F_e = F_c$
- The total energy of the electron is negative. It is given by $E = -\dfrac{e^2}{8\pi\varepsilon_0 r}$.

Atomic Spectra

- Each element has a characteristic spectrum of radiation, which it emits.
- Study of emission line spectra of a material can therefore serve as a type of "fingerprint" for identification of the gas.
- The atomic hydrogen emits a line spectrum consisting of various series as:

  Lyman series: $v = Rc\left(\dfrac{1}{1^2} - \dfrac{1}{n^2}\right)$: $n = 2, 3, 4, \ldots$

  Balmer series: $v = Rc\left(\dfrac{1}{2^2} - \dfrac{1}{n^2}\right)$: $n = 3, 4, 5, \ldots$

  Paschen series: $v = Rc\left(\dfrac{1}{3^2} - \dfrac{1}{n^2}\right)$: $n = 4, 5, 6, \ldots$

  Brackett series: $v = Rc\left(\dfrac{1}{4^2} - \dfrac{1}{n^2}\right)$: $n = 5, 6, 7, \ldots$

  Pfund series: $v = Rc\left(\dfrac{1}{5^2} - \dfrac{1}{n^2}\right)$: $n = 6, 7, 8, \ldots$

Bohr Model of the Hydrogen Atom

Bohr combined classical and early quantum concepts, explained the spectrum of hydrogen atom based on quantum ideas and gave his theory in the form of three postulates. These are:

- Bohr's first postulate was that an electron in an atom could revolve in certain stable orbits without the emission of radiant energy, contrary to the predictions of electromagnetic theory. According to this postulate, each atom has certain definite stable states in which it can exist, and each possible state has definite total energy. These are called the stationary states of the atom.
- Bohr's second postulate defines these stable orbits. This postulate states that the electron revolves around the nucleus only in those orbits for which the angular momentum is some integral multiple of $h/2\pi$ where h is the Planck's constant ($= 6.6 \times 10^{-34}$ Js). Thus the angular momentum (L) of the orbiting electron is quantised. That is $L = nh/2\pi$.

- Bohr's third postulate incorporated into atomic theory the early quantum concepts that had been developed by Planck and Einstein. It states that an electron might make a transition from one of its specified non-radiating orbits to another of lower energy. When it does so, a photon is emitted having energy equal to the energy difference between the initial and final states. The frequency of the emitted photon is then given by

  $h\nu = E_i - E_f$, where E_i and E_f are the energies of the initial and final states and $E_i > E_f$.
- Bohr radius is represented by the symbol a_0, is given by $a_0 = \dfrac{h^2\varepsilon_0}{\pi m e^2}$.
- The total energy of the electron in the stationary states of the hydrogen atom is given by

  $$E_n = -\frac{13.6}{n^2} eV$$

De Broglie's Explanation of Bohr's Second Postulate of Quantisation

- De Broglie hypothesis provided an explanation for Bohr's second postulate for the quantisation of angular momentum of the orbiting electron. The quantised electron orbits and energy states are due to the wave nature of the electron and only resonant standing waves can persist.
- De Broglie's hypothesis is that electrons have a wavelength $\lambda = \dfrac{h}{mv}$.

Limitations of Bohr's model: Bohr's model however has many limitations.

- It is applicable only to hydrogenic (single electron) atoms.
- It cannot be extended to even two electron atoms such as helium.
- While the Bohr's model correctly predicts the frequencies of the light emitted by hydrogenic atoms, the model is unable to explain the relative intensities of the frequencies in the spectrum.

Exercise

1. If in nature there may not be an element for which the principal quantum number n > 4, then the total possible number of elements will be

(*a*) 60 (*b*) 32
(*c*) 4 (*d*) 64

2. In the Bohr's hydrogen atom model, the radius of the stationary orbit is directly proportional to (n = principal quantum number)

(*a*) n^{-1} (*b*) n
(*c*) n^{-2} (*d*) n^2

3. In the following atoms and molecules for the transaction from n = 2 to n = 1, the spectral line of minimum wavelength will be produced by

(*a*) Hydrogen atom (*b*) Deuterium atom
(*c*) Uni-ionized helium (*d*) di-ionized lithium

4. The Lyman series of hydrogen spectrum lies in the region

(*a*) Infrared (*b*) Visible
(*c*) Ultraviolet (*d*) Of X-rays

5. The size of an atom is of the order of

(*a*) 10^{-8} m (*b*) 10^{-10} m
(*c*) 10^{-12} m (*d*) 10^{-14} m

6. Which one of the series of hydrogen spectrum is in the visible region

(*a*) Lyman series (*b*) Balmer series
(*c*) Paschen series (*d*) Bracket series

7. The energy required to remove an electron in a hydrogen atom from n = 10 state is

(*a*) 13.6 eV (*b*) 1.36 eV
(*c*) 0.136 eV (*d*) 0.0136 eV

8. The kinetic energy of the electron in an orbit of radius r in hydrogen atom is (e = electronic charge)

(*a*) $\dfrac{e^2}{r^2}$ (*b*) $\dfrac{e^2}{2r}$
(*c*) $\dfrac{e^2}{r}$ (*d*) $\dfrac{e^2}{2r^2}$

9. The ratio of the energies of the hydrogen atom in its first to second excited state is

(*a*) $\frac{1}{4}$ (*b*) $\frac{4}{9}$

(*c*) $\frac{9}{4}$ (*d*) 4

10. According to Bohr's theory the radius of electron in an orbit described by principal quantum number n and atomic number Z is proportional to

(*a*) Z^2n^2 (*b*) $\frac{Z^2}{n^2}$

(*c*) $\frac{Z^2}{n}$ (*d*) $\frac{n^2}{Z}$

11. The radius of electrons second stationary orbit in Bohr's atom is R. The radius of the third orbit will be

(*a*) 3R (*b*) 2.25 R

(*c*) 9R (*d*) $\frac{R}{3}$

12. In any Bohr orbit of the hydrogen atom, the ratio of kinetic energy to potential energy of the electron is

(*a*) ½ (*b*) 2

(*c*) –½ (*d*) –2

13. The spectral series of the hydrogen spectrum that lies in the ultraviolet region is the

(*a*) Balmer series (*b*) Pfund series

(*c*) Paschen series (*d*) Lyman series

14. In Bohr model of the hydrogen atom, the lowest orbit corresponds to

(*a*) Infinite energy

(*b*) The maximum energy

(*c*) The minimum energy

(*d*) Zero energy

15. If an electron jumps from 1st orbital to 3rd orbital then it will

(*a*) Absorb energy (*d*) Release energy

(*c*) No gain of energy (*d*) None of these

16. To explain his theory, Bohr used

(*a*) Conservation of linear momentum

(*b*) Conservation of angular momentum

(*c*) Conservation of quantum frequency

(*d*) Conservation of energy

17. Number of spectral lines in hydrogen atom is

(*a*) 3 (*b*) 6

(*c*) 15 (*d*) Infinite

18. Minimum excitation potential of Bohr's first orbit in hydrogen aotm is

(*a*) 13.6 V (*b*) 3.4 V

(*c*) 10.2 V (*d*) 3.6 V

19. According to the Rutherford's atomic model, the electrons inside the atom are

(*a*) Stationary (*b*) Not stationary

(*c*) Centralized (*d*) None of these

20. According to classical theory, the circular path of an electron in Rutherford atom is

(*a*) Spiral (*b*) Circular

(*c*) Parabolic (*d*) Straight line

21. Rutherford's α-particle experiment showed that the atoms have

(*a*) Proton (*b*) Nucleus

(*c*) Neutron (*d*) Electrons

22. In hydrogen atom, when electron jumps from second to first orbit, then energy emitted is

(*a*) –13.6 eV (*b*) –27.2 eV

(*c*) –6.8 eV (*d*) None of these

23. For ionising an excited hydrogen atom, the energy required (in eV) will be

(*a*) A little less than 13.6

(*b*) 13.6

(*c*) More than 13.6

(*d*) 3.4 or less

24. The kinetic energy of electron in the first Bohr orbit of the hydrogen atom is

(*a*) –6.5 eV (*b*) –27.2 eV

(*c*) 13.6 eV (*d*) –13.6 eV

Answer Keys

1. (*a*)	**2.** (*d*)	**3.** (*d*)	**4.** (*c*)	**5.** (*b*)	**6.** (*b*)	**7.** (*c*)	**8.** (*b*)	**9.** (*c*)	**10.** (*d*)
11. (*b*)	**12.** (*c*)	**13.** (*d*)	**14.** (*c*)	**15.** (*a*)	**16.** (*b*)	**17.** (*d*)	**18.** (*c*)	**19.** (*b*)	**20.** (*a*)
21. (*b*)	**22.** (*d*)	**23.** (*d*)	**24.** (*c*)						

Solutions

1. For n = 1, maximum number of states = $2n^2 = 2$ and for n = 2, 3, 4, maximum number of states would be 8, 18, 32 respectively. Hence, number of possible elements
   = 2 + 8 + 18 + 32 = 60
2. Bohr radius
   $$r = \frac{\varepsilon_0 n^2 h^2}{\pi Z m e^2}$$
   $\therefore r \propto n^2$
3. $$\frac{1}{\lambda} = RZ^2\left(\frac{1}{1^2} - \frac{1}{2^2}\right)$$
   For di-ionised lithium the value of Z is maximum.
4. Lyman series lies in the UV region.
5. The size of the atom is of the order of 1 Å = 10^{-10} m
6. Balmer series lies in the visible region.
7. Energy required
   $$= \frac{13.6}{n^2} = \frac{13.6}{10^2} = 0.136\text{eV}$$
8. Potential energy of electron in n^{th} orbit of radius r in H-atom
   $$U = -\frac{e^2}{r}(\text{in CGS})$$
   $$\because \text{ K.E.} = \frac{1}{2}|\text{P.E.}| \Rightarrow K = \frac{e^2}{2r}$$
9. First excited state i.e., second orbit (n = 2)
   Second excited state i.e., third orbit (n = 3)
   $$\because \text{ E.} = -\frac{13.6}{n^2} \Rightarrow \frac{E_2}{E_3} = \left(\frac{3}{2}\right)^2 = \frac{9}{4}$$
10. $$r = \frac{\varepsilon_0 n^2 h^2}{\pi Z m e^2}$$
    $$\therefore r \propto \frac{n^2}{Z}$$
11. $$r \propto n^2 \Rightarrow \frac{r_{(n=2)}}{r_{(n=3)}} = \frac{4}{9}$$
    $$\therefore r_{(n=3)} = \frac{9}{4}R = 2.25R$$
12. $$\text{K.E} = \frac{kZe^2}{2r} \text{ and P.E.} = -\frac{kZe^2}{r}$$
    $$\therefore \frac{\text{K.E}}{\text{P.E.}} = -\frac{1}{2}$$
13. Lyman series lies in the UV region.
14. In hydrogen atom, the lowest orbit (n = 1) correspond to minimum energy (–13.6 eV).
15. When an electron jumps from the orbit of lower energy (n = 1) to the orbit of higher energy (n = 3), energy is absorbed.
16. Bohr postulated that the angular momentum of the electron is conserved.
17. Infinitely large transitions are possible (in principle) for the hydrogen atom.
18. $$\text{Excitation potential} = \frac{\text{Excitation energy}}{e}$$
    Minimum excitation energy corresponds to excitation from n = 1 to n = 2
    ∴ Minimum excitation energy in hydrogen atom
    = –3.4 – (–13.6) = + 10.2eV
    Hence, minimum excitation potential = 10.2 V
19. According to the Rutherford's atomic model, the electron, inside the atoms are not stationary.
20. According to classical theory, the circular path of an electron in Rutherford atom is spiral.
21. Rutherford's α-particle experiment showed that the atoms have nucleus.
22. $$E_{n_1 \to n_2} = -13.6\left[\frac{1}{n_2^{\ 2}} - \frac{1}{n_1^{\ 2}}\right]; n_1 = 2 \ \& \ n_2 = 1$$
    $$\Rightarrow E_{II} \to E_I = -13.6 \times \frac{3}{4} = -10.2\text{eV}$$
23. Energy required ionising an excited hydrogen atom = ionisation energy –excitation energy (in first excited state)
    = 13.6 – 10.2 = 3.4 eV
24. K.E. = –(Total energy)
    = – (–13.6 eV) = + 13.6 eV