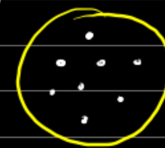
Structure of Atom

Discovery of Electron and Nucleus

* J.J Thompson discovered Electrons by his experiments with cathode ray tubes.

* Thompson's Plum Pudding Model had negatively charged electrons embedded in positively charged soup.



* Rutherford's Gold Foil Experiment ⟶ atom is mostly empty space, with a tiny, dense positively charged Nucleus.

* Based on these results, Rutherford proposed the Nuclear Model of Atom.

Dalton's Atomic Theory

• All matter is made of tiny indivisible particles called ATOMS, which can neither be created, nor destroyed.

Dalton's Atomic Theory (1808)
J. J. Thomson ⟶ electron (1896)
Thomson's atomic model (1898)
Boltzmann (1906)
Rutherford's model (1911)
Bohr's Model (1915)
Schrodinger Model (1925)

Planck's Quantum theory

1900 - 1930
Major changes

1. An alpha particle after passing through a potential difference of 2×10^6 volt falls on a silver foil. The atomic number of silver is 47. Calculate (i) the K.E. (in Joule) of the alpha-particle at the time of falling on the foil. (ii) K.E. (in Joule) of the α – particle at a distance of 5×10^{-14}m from the nucleus, (iii) the shortest distance (in m) from the nucleus of silver to which the α-particle reaches.

1. Potential difference = 2×10^6 V; Charge of alpha particle=2e; Charge of silver = 47e

KE of alpha particle = $qV=2\times1.6\times10^{-19}\times2\times10^{6}=6.4\times10^{-13}$ J

K.E at A = K.E at B + P.E at B

$$6.4\times10^{-13} = \text{K.E at B} + \frac{9\times10^9\times2\times1.6\times10^{-19}\times47\times10^{-19}}{5\times10^{-14}}$$

K.E at B = 2.1×10^{-13} joules

2. Suppose the potential energy between electron and proton at a distance r is given by $-\frac{ke^2}{3r^3}$. Bohr's theory to obtain energy of such a hypothetical atom.

2. Since H atom is a bounded system U cannot be positive

$$U = -\frac{ke^2}{3r^3}; \qquad F = -\frac{dU}{dr} = -\frac{ke^2}{r^4} \text{ ....1}$$

$$\frac{mv^2}{r} = \frac{ke^2}{r^4}; \qquad mv^2 = \frac{ke^2}{r^3}; \qquad KE = \frac{1}{2}mv^2 = \frac{ke^2}{2r^3}$$

$$TE = KE + U = \frac{ke^2}{2r^3} - \frac{ke^2}{3r^3} = -\frac{ke^2}{6r^3}$$

3. Electron present in single electron specie jumps from energy level 3 to 1. Emitted photons when passed through a sample containing excited He^+ ion causes further excitation to some higher energy level (Given $E_n = -13.6\frac{Z^2}{n^2}$). Determine

(i) Atomic number of single electron specie.

(ii) principal quantum number of initial excited level & higher energy level of He^+.

3. Energy of emitted photons cannot be greater than 13.6 eV (otherwise He^+ will ionise) therefore single electron specie must be hydrogen energy emitted = $E_3 - E_1$

$= -1.51 + 13.6 = 12.09$

For He^+ ion this energy corresponds to excitation from 2 to 6.

4. The angular momentum of an electron in a Bohr's orbit of H-atom is 3.1652×10^{-34} kg-m²/sec. Calculate the wave number in terms of Rydberg's constant (R) of the spectral line emitted when an electron falls from this level to the ground state.[Use h = 6.626×10^{-34} Js]

4. $L = \frac{n\times6.625\times10^{-34}}{2\times3.14}$; $n = 3$

$\frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{3^2}\right] = R\left(\frac{8}{9}\right)$

5. A proton captures a free electron whose K.E. is zero & forms a hydrogen atom of lowest energy level (n = 1). If a photon is emitted in this process, what will be the wavelength (in Å) of radiation? In which region of electromagnetic spectrum, will this radiation fall? (Ionisation potential of hydrogen = 13.6 volt, h = 6.6×10^{-34}K/s, C = 3.0×10^{8} m/s)

5. **IE = 13.6 eV ;** $\lambda = \frac{12400}{13.6} = 910A^0$

6. 1.8 g hydrogen atoms are excited to radiations. The study of spectra indicates that 27% of the atoms are in 3rd energy level and 15% of atoms in 2nd energy level and the rest in ground state. If I.P.of H is 21.7×10^{-12} erg. Calculate –

(i) Number of atoms present in III & II energy level.

(ii) Total energy (in kJ) evolved when all the atoms return to ground state.

6. (i) Total number of H atoms =$1.8 \times 6.023 \times 10^{23}$l= 1.084×10^{24}

Number of H atoms in 2nd energy level = $0.15 \times 1.08 \times 10^{24} = 1.62 \times 10^{23}$ atoms,

Number of H atoms in 3nd energy level = $0.27 \times 1.08 \times 10^{24} = 2.92 \times 10^{23}$ atoms,

(ii) Energy evolved for n =2⟶1

$$E_1 = 21.7 \times 10^{-12} \times \left[\frac{1}{1} - \frac{1}{4}\right] = 1.63 \times 10^{-11}$$

Energy evolved for n = 3⟶1

$$E_2 = 21.7 \times 10^{-12} \times \left[\frac{1}{1} - \frac{1}{9}\right] = 1.93 \times 10^{-11}$$

Total energy=$1.63 \times 10^{-11} \times 1.62 \times 10^{23}$l+$1.93 \times 10^{-11} \times 2.92 \times 10^{23}$l = 8.32×10^{12} erg = 8.32×10^{2} KJ

7. The ionisation energy of the hydrogen atom is given to be 13.6 eV. A photon falls on a hydrogen atom which is initially in the ground state and excites it to the (n = 4)state.

(a) show this transition in the energy-level diagram &

(b) calculate the wavelength (in Å) of the photon.

7. $\mathbf{E} = 13.6\times\left(\frac{1}{1}-\frac{1}{16}\right)$ eV = 12.75 eV

$\lambda = \frac{12400}{12.75} = 972.5$ Å

8. The ionization energy of a H-like Bohr atom is 4 Rydbergs

(i) What is the wavelength (in Å) of radiation emitted when the e– jumps from the first excited state to the ground state.

(ii) What is the radius (in cm) of first Bohr orbit for this atom. [1 Rydberg = 2.18×10^{-18} J]

8. $R_H = 2.18\times10^{-18}$

$E_2-E_1 = \frac{hc}{\lambda}$

$E_1 = -4R_H = -4R_H$

$E_2 = \frac{-4R_H}{4} = -RH$

$E_2-E_1 = 3R_H = 3\times2.18\times10^{-18}\text{J} = \frac{hc}{\lambda}$

$\lambda = 303.89\times10^{-10}$m