Hello Sir !

CHEMISTRY Academy

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Class XI, Unit 2

Topic 1: Bohr's theory: 2h Coaching

1. Bohr's postulates
2. Descriptive questions and answers
3. Bohr's Radius:
   i. Equation
   ii. Relationship of "r"with n and Z
   iii. Questions & answers
4. Electron energy from Bohr's theory
   i. eqaution
   ii. Relationship of "E"with others n and Z
   iii. Questions & answers
5. Energy change for Electron transition in Bohr's model
   i. equation
   ii. Questions & answers
6. Light emission for Electron transition in Bohr's model
   Questions & answers
7. Ionization Energy from Bohr's theory
   Questions & answers: simple

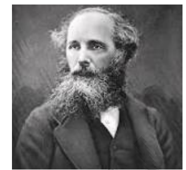
Bohr’s atomic model of hydrogen atom (1913), nobel prize (1922)


J J Balmer


Luice Planck


Niels Bohr


Rutherford


Maxwell

Background

1. Niels Bohr proposed atomic model (1913) of hydrogen atom based on Planck’s quatum theory (1890), Which is $E=h\nu$.

2. He agreed with Rutherford atomic model (1911):

    i. First idea about **nucleus** and which is positively charged (gold foil experiment)
    ii. First idea of existence of **orbit** in an atom
    iii. **Electron moves around positively charged nucleus** in a orbit of atom.

3. Niels Bohr’s theory didn’t violate Maxwell’s electromagnetic hypothesis

4. Balmer noticed spectral lines from hydrogen atom (1985), but failed to explain. Bohr’s theory expalined that fact.

Bohr’ postulates of hydrogen atom:

1. **Movement of electron:** electron moves around positively charged nucleus in a orbit of hydrogen atom

2. Energy:

i. Energy of orbit: Electron doesn’t radiate energy while moves in orbit around nucleus of hydrogen atom

ii. Energy of electron:

a. electron radiates energy while it moves “towards” nucleus
b. electron absorbs energy while it moves “away from” nucleus

3. Two forces

Electrostatic attraction between positively charged nucleus and negatively charged elctron is equal to the centripetal force.

Bohr's postulates of hydrogen atom explanation

1. **Movement of electron:** electron moves around positively charged nucleus in a orbit of hydrogen atom (**as per Rutherford model**)
2. **Energy:**

i. **Energy of orbit:** Electron doesn't radiate energy while moves around nucleus of hydrogen atom

*Because its exists in quantized energy levels within the atom (**as per fundamental quantum mechanics of Planck's theory**)*

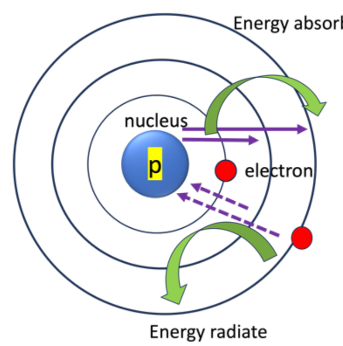
Angular mometum of electron = Quantized energy

$$mvr = \frac{nh}{2\pi} \quad [n=1, 2, 3.....]$$

Note: that's why no chances of **energy loss/radiation** of electron: electron movement never slowing down (not violating **Maxwell's** electro magenitc radiation hypothesis)

ii. Energy of electron It explains **Balmer** spectral lines (1885) of hydrogen atom

a. electron radiates energy while it moves "towards" nucleus.
   (n^{th} orbit energy higher,)



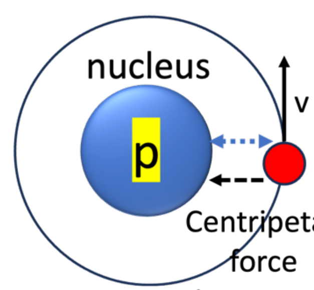
b. electron absorbs energy while it moves "away from" nucleus because more attraction force on closest orbit's electron from nucleus, so energy required to cancel attraction force.

3. Two forces

Electrostatic attraction = centripetal force.

(between positively charged nucleus and negatively charged elctron)

(Which tends to throw the electron out of orbit)



Centripetal force = Electrostatic force

Questions from Bohr's theory

1. Electron moves, but not slowing down?

Ans. Electron doesn't radiate energy while moves around nucleus of hydrogen atom
Because it exists in quantized energy levels within the atom

2. First Concept of ***light emission*** from an atom

From Bohr's theory: Electron radiates energy as light while it moves "towards" nucleus, because energy is related to wavelength according to Planck's quantum equation as below.

$$\boxed{E = h\nu = \frac{hc}{\lambda}}$$

3. Why electrons don't fall into the nucleus
Ans. because
   1. electrons exist in specific energy levels or orbits around the nucleus,
   2. electrons behave as waves, so it impossible for them to be localized in a single point.

4. In Bohr's theory, angular mometum of electron in quantized. Which property of electron signifies?

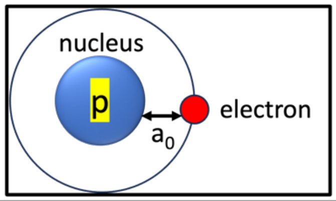
Ans.Particles property

Bohr's Equation for Hydrogen atom

1. **Bohr's radius (a_0)**
(hydrogen atom radius)
2. **Electron's energy**

Bohr's radius

Bohr's radius is the most probable distance between the nucleus and the electron in a hydrogen atom in its ground state (n = 1).
It is denoted as a_0 and has a value of approximately 5.29 x 10^{-11} meters



Bohr's radius euation

Bohr's 3rd postulate

Centripetal force = Electrostatic force

$$\frac{mv^2}{r} = \frac{Z\,e^2}{4\pi\,\varepsilon_0 r^2} \quad \text{..........Eq.1}$$

Bohr' 2nd postulate

Angular mometum of electron = quantized energy of orbit

Planck equation, photon=quanta $mvr = \frac{nh}{2\pi}$

$$\text{So, } v^2 = \frac{n^2h^2}{4\pi^2 m^2 r^2} \quad \text{..........Eq.2}$$

From Eq. 1 & 2

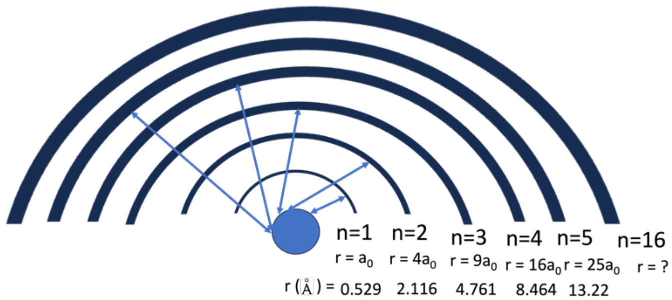
$$\frac{Z\,e^2}{4\pi\, m\varepsilon_0 r} = \frac{n^2h^2}{4\pi^2 m^2 r^2} \longrightarrow r = \frac{\varepsilon_0 n^2h^2}{\pi\ m\, e^2\, z} = \boxed{0.529\frac{n^2}{z}} = 0.529\ A^{O} = a_0 \text{ (For H atom)}$$

m= mass of electron = 9.1×10^{-31} kg
v= velocity of electron
n=orbit number = principal quantum number
h=Planck's constant = 4.136×10^{-15} eV/s
= 6.626×10^{-34} J s
= 6.626×10^{-34} Kg m^2 s^{-1}

ε_0= permittivity of free space
e: charge of the electron
z= atomic number

Relationship: Radius (r) with principal quantum number (n):
r with ***n* square relation**

With increasing principal quantum number the attraction force over the electron from the nucleus is reduced. **Electron being far away from nucleus** resulting increase of atomic radius.



Radius (r) with atomic number (z):

i. With increasing atomic number the attraction force over the electron from the nucleus is increased. **Electrons closer to nucleus** resulting decrease of atomic radius.

Comprehensive Questions and answers from Bohr's radius

1. What is Bohr's radius? Derive its mathematical expression.

Answer:

Bohr's radius is the radius of the **lowest energy orbit (ground state)** of the hydrogen atom. It represents the most probable distance between the nucleus and the electron in the ground state of hydrogen.

To derive it, we use Bohr's postulates:

Bohr's 3rd postulate

Centripetal force = Electrostatic force

$$\frac{mv^2}{r} = \frac{Z e^2}{4\pi \varepsilon_0 r^2} \quad \text{.......... Eq.1}$$

Bohr' 2nd postulate

Angular mometum of electron = quantized energy of orbit

Planck equation, photon=quanta $\quad mvr = \frac{nh}{2\pi}$

$$\text{So, } v^2 = \frac{n^2h^2}{4\pi^2m^2 r^2} \quad \text{.......... Eq.2}$$

From Eq. 1 & 2

$$\boxed{\frac{Z e^2}{4\pi m \varepsilon_0 r} = \frac{n^2h^2}{4\pi^2m^2 r^2}} \longrightarrow r = \frac{\varepsilon_0 n^2h^2}{\pi\ m e^2 z} = \boxed{0.529 \frac{n^2}{z}} = 0.529 \text{ A}^{O} = a_0 \text{ (For H atom)}$$

2. What is the physical significance of Bohr's radius?

Answer:

Bohr's radius value indicates the existence of electron around the nucleus within this area.

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3. How does the radius of an orbit change with the principal quantum number n?
Answer:
In Bohr's model, the radius of the electron orbit increases with the square of the principal quantum number *n*:
$r_n = a_0 n^2$
This means:

- The first orbit (n=1) has radius a_0
- The second orbit (n=2) has radius $4a_0$
- The third orbit (n=3) has radius $9a_0$ and so on.

So, the radius increases **quadratically** with energy levels, indicating that higher energy electrons are further from the nucleus.

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4. What factors does Bohr's radius depend on?
Answer:
Bohr's radius is given by:
$a_0 = 4\pi\varepsilon_0\hbar 2/me^2$
It depends on:

- ε_0= permittivity of free space
- $\hbar$= reduced Planck's constant
- m: mass of the electron
- e: charge of the electron

It does **not** depend on atomic number Z for hydrogen, but for hydrogen-like ions (like He^+, Li^{2+}), the radius formula becomes:

$$r_n = a_0 n^2/Z$$

Energy

Total Energy = K.E. + P.E $= \frac{1}{2}mv^2 + \left(-\frac{Z e^2}{4\pi \varepsilon_0 r}\right) = \frac{1}{2}\frac{Z e^2}{4\pi \varepsilon_0 r} - \frac{Z e^2}{4\pi \varepsilon_0 r}$

$$= -\frac{Z e^2}{8\pi \varepsilon_0 r} = -\frac{e^4 m Z^2}{8 \varepsilon_0^2 h^2 n^2} = -R\frac{Z^2}{n^2} = \boxed{-13.6 \frac{Z^2}{n^2} \text{ eV/atom}}$$

$$= -21.8 \times 10^{-9} \frac{Z^2}{n^2} \text{ J /atom}$$

$$= -1.097 \times 10^{-7} \frac{Z^2}{n^2} \text{ m}^{-1}\text{/atom}$$

Planck constant, $h = 4.136 \times 10^{-15}$ eV. s

$= 6.626 \times 10^{-34}$ J s

$= 6.626 \times 10^{-34}$ Kg m^2 s^{-1}

Total Energy (E) = Kintetic energy (KE) + Potential Energy (PE)

K.E.= -E,   PE= 2E

Relationship: E with Z

$$E = \boxed{-13.6 \frac{Z^2}{n^2} \text{ eV/atom}}$$

Explanation

1. A higher atomic number means a greater positive charge in the nucleus, which exerts a stronger attractive force on the electrons to stabilize electrons resulting in reduction their energy.

Energy Difference

Energy Difference

n_L= lower energy state, n_H= higher energy state $n_L \sim n_H$

$$E = \boxed{-13.6 \frac{Z^2}{n^2} \text{ eV/atom}}$$

$$\Delta E = -\frac{Z^2 e^4 m}{8 \varepsilon_0^2 n_H^2 h^2} - \left(-\frac{Z^2 e^4 m}{8 \varepsilon_0^2 n_L^2 h^2}\right) = \frac{Z^2 e^4 m}{8 \varepsilon_0^2 h^2}\left(\frac{1}{n_L^2} - \frac{1}{n_H^2}\right) = 13.6\left(\frac{1}{n_L^2} - \frac{1}{n_H^2}\right)$$

Electron transition/jumps

Calculate the energy for electron transition from n=2 to n=1 for hydrogen atom

$$E_1 - E_2 = -13.6 \frac{1}{n_{=1}^2} - \left(-13.6 \frac{1}{n_{=2}^2}\right) = 13.6\left(\frac{1}{n_{=2}^2} - \frac{1}{n_{=1}^2}\right) = 13.6\left(\frac{1}{4} - \frac{1}{1}\right) = -10.2 \text{ eV/atom}$$

Ans. Absolute value: 10.2 eV

Calculate the energy for electron transition from n=1 to n=2 for hydrogen atom

$$E_2 - E_1 = -13.6 \frac{1}{n_{=2}^2} - \left(-13.6 \frac{1}{n_{=1}^2}\right) = 13.6\left(\frac{1}{n_{=1}^2} - \frac{1}{n_{=2}^2}\right) = 13.6\left(\frac{1}{1} - \frac{1}{4}\right) = 10.2 \text{ eV/atom}$$

Ans. Absolute value: 10.2 eV

Objective Questions and answers on energy and energy difference from Bohr's theroy

1. The energy difference ΔE between two levels in the hydrogen atom is responsible for:
A) Radioactivity
B) Emission or absorption of photons
C) Nuclear fission
D) Change in atomic number
Answer: B

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2. Which transition in hydrogen atom emits the highest energy photon?
A) $n=3 \rightarrow n=2$
B) $n=4 \rightarrow n=2$
C) $n=2 \rightarrow n=1$
D) $n=5 \rightarrow n=4$
Answer: C
(Because ΔE is maximum for transitions to the ground state)

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3. The energy of an orbit in Bohr's model is:
A) Directly proportional to n
B) Inversely proportional to n
C) Inversely proportional to n^2
D) Proportional to $n^{1/2}$
Answer: C

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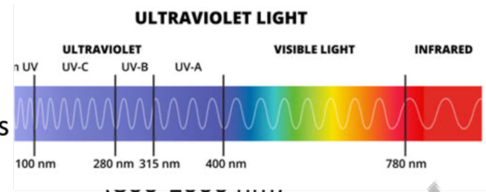
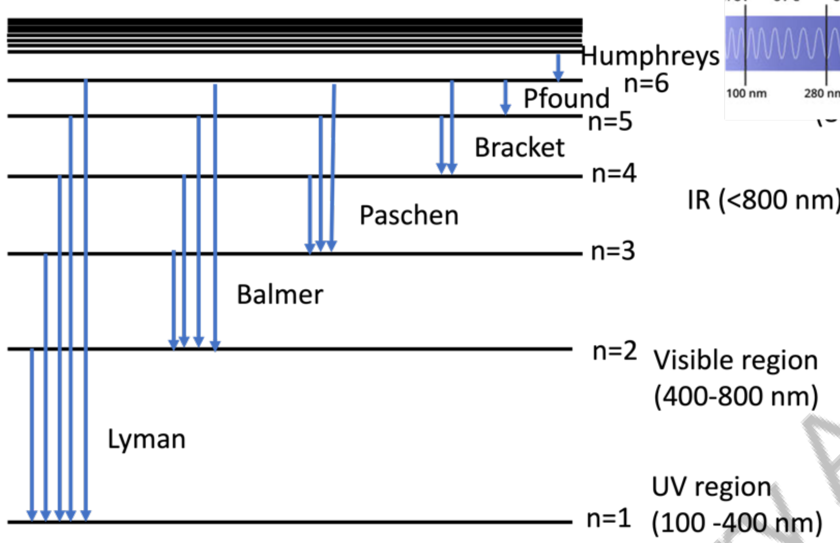
4. Which spectral series results from transitions to n=1 level in hydrogen?
A) Balmer
B) Lyman
C) Paschen
D) Brackett
Answer: B
(Lyman series involves transitions to the ground state)

Light emission or light absorption during electron transition from Bohr's theory

Spectral lines (light emission) from Bohr's theory





No. of photons emitted by a sample of H atom : $\Delta n(\Delta n+1)/2$

Descriptive Questions and answers on light absorption/emission from Bohr's theroy

1. Why does hydrogen emit radiation at specific wavelengths rather than a continuous spectrum?

Answer:
In Bohr's model, electrons in a hydrogen atom can only occupy **discrete energy levels** (quantized orbits). When an electron transitions between these levels, it must **absorb or emit energy equal to the difference** between them.
$\Delta E = E_{final} - E_{initial} = h\nu = hc/\lambda$
Since the energy levels are discrete, only **specific values of ΔE** (and hence specific wavelengths) are possible. This results in **line spectra**, not continuous spectra.

2. A hydrogen atom absorbs a photon of energy 12.09 eV. To what energy level has the electron jumped?

Answer:
The ground state energy of hydrogen is E1=−13.6 eV
$E_{final} = E1 + 12.09 = -13.6 + 12.09 = -1.51$ eV
Now, use Bohr's formula:
$E_n = -13.6/n^2$

$\Rightarrow -1.51 = -13.6/n^2$
$\Rightarrow n2 = 13.6/1.51 \approx 9 \Rightarrow n=3$

3. Why are the spectral lines of hydrogen-like ions such as He^+ or Li^{2+} not visible to the naked eye?

Answer:
Spectral lines of hydrogen-like ions occur at much shorter wavelengths than those of hydrogen because the energy level differences scale with Z^2. For example, Li^{2+} has Z=3, so transitions are 9× more energetic than in hydrogen. These high-energy transitions typically fall in the ultraviolet or X-ray regions, which are invisible to the naked eye.

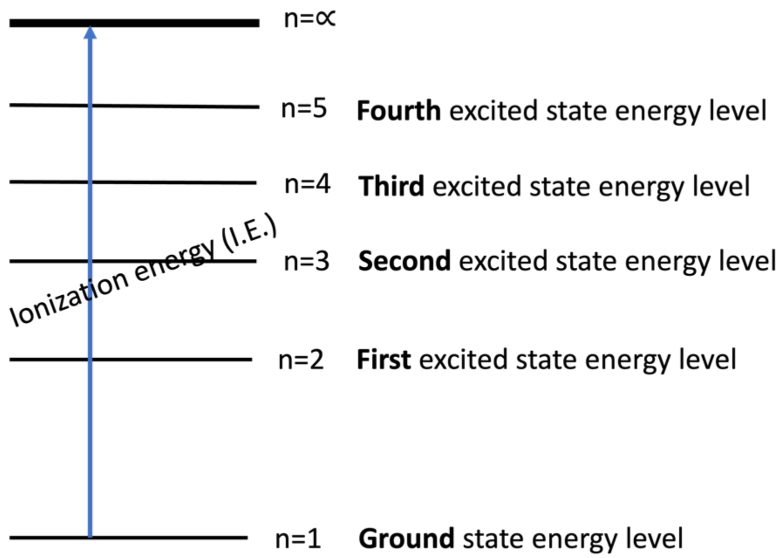
4. A transition from n=6 to n=1 in hydrogen emits a photon. Calculate the energy and wavelength.

Answer:
$\Delta E = 13.6(1/1^2 - 1/6^2) = 13.6(1/1 - 1/36) = 13.6 \times 35/36 \approx 13.22$ eV
$\lambda = hc/E = 1240$ eV nm $/13.22$ eV ≈ 93.8nm
Answer: 13.22 eV, 93.8 nm (UV range)

5. Why are transitions to n=1 referred to as the Lyman series, and why are they only visible in the UV region?

Answer:
The Lyman series consists of transitions that end at n=1. These transitions involve **large energy differences**, resulting in **shorter wavelengths** (ultraviolet region). The energy gap between higher levels and n=1 is significant, making emitted photons very energetic. Hence the UV nature of Lyman series lines.

Ionization Energy



$$E_{\propto} - E_1 = 13.6\left(\frac{1}{1^2} - \frac{1}{\propto^2}\right)$$

=13.6 eV/atom

How much energy required (ionization) to inonize 1 mole of hydrogen ?

Ans. 1312 kJ/mole

Descriptive Questions and answers on Ionization energy from Bohr's theroy

Q: What is ionization energy according to Bohr's theory? Derive the expression for the ionization energy of hydrogen and hydrogen-like atoms. Explain how it varies with atomic number.

Definition of Ionization Energy:
Ionization energy (IE) is the **minimum energy required to remove an electron** from the ground state of an atom (i.e., from the lowest energy level n=1) to an infinite distance (n=∞), where the electron is completely free from the atom.

Bohr's Theory Expression for Energy Levels:
According to Bohr's model, the energy of an electron in the n^{th} orbit of a hydrogen-like atom is:

$$E_n = -13.6 . Z^2/n^2 \text{ eV}$$

Where:

- Z is the atomic number (1 for H, 2 for He^+, etc.)
- n is the principal quantum number
- 13.6 eV is the ionization energy of hydrogen in its ground state.

For He^+:
$IE = 13.6 \times (2)^2$, IE = 54.4eV