Chapter 17
The Atom

Unit 4: Atom and Nucleus

Introduction

- **White light**: is a polychromatic light consisting of a mixture of all the wavelengths of the visible spectrum.
- **Dispersion of light:** is the separation of a polychromatic light into its different colors.
- **Spectroscope**: is an optical apparatus that uses a dispersing system (like prism or diffraction grating) to separate light in order to observe the different monochromatic lights forming it (Figure 1-a).

Fig (1-a)

- Dispersion of light can be done using a:
  - **prism**: when light falls on a prism, it undergoes refraction, so light is dispersed or separated (Figure 1-b).
  - **diffraction grating**: when light falls on a grating it undergoes diffraction, so light is dispersed (Figure 1-c).

Fig (1-b)

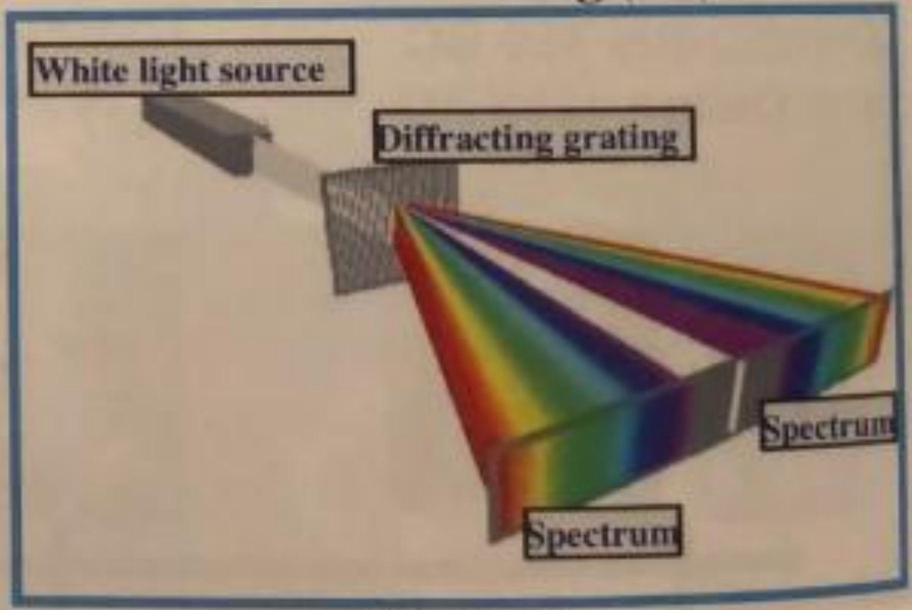


Fig (1-c)

➢ **Emission spectrum of an element** is the set of wavelengths that constitute the radiation emitted by this element when it is excited.

Types of emission spectra of an element:

<u>Continuous emission spectrum</u>: Continuous band of colors without dark lines (Figure 2-a).

<u>Discrete emission spectrum</u>: Set of discrete bright lines on a dark background (Figure 2-b).

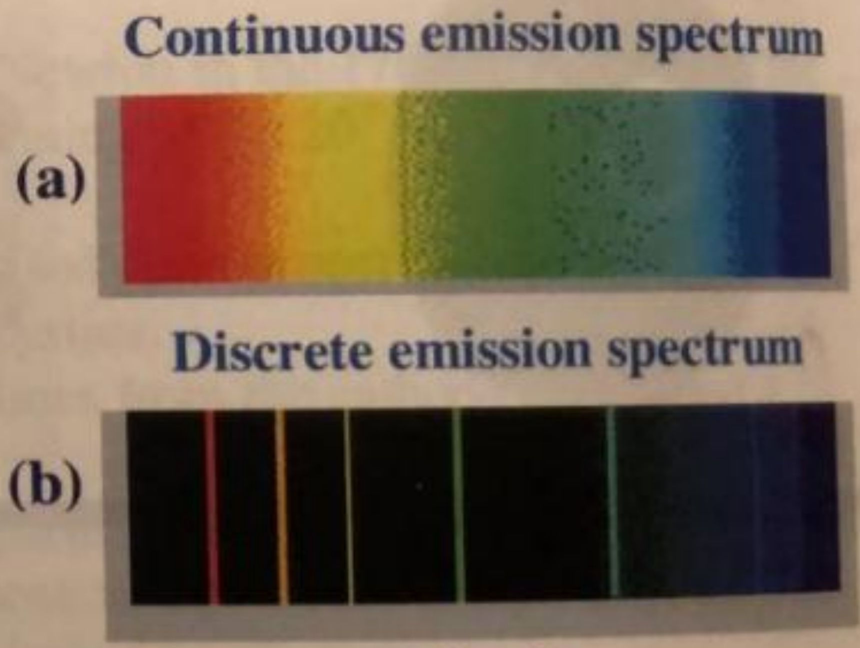
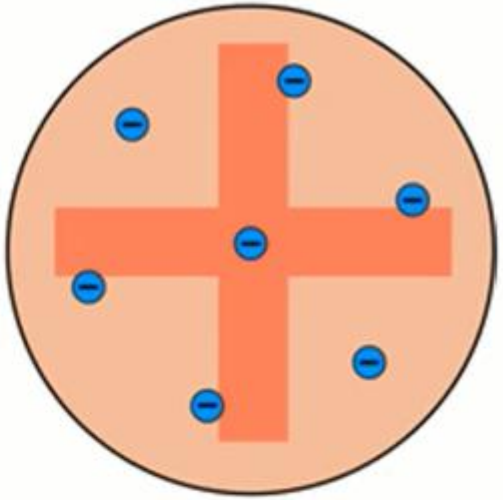
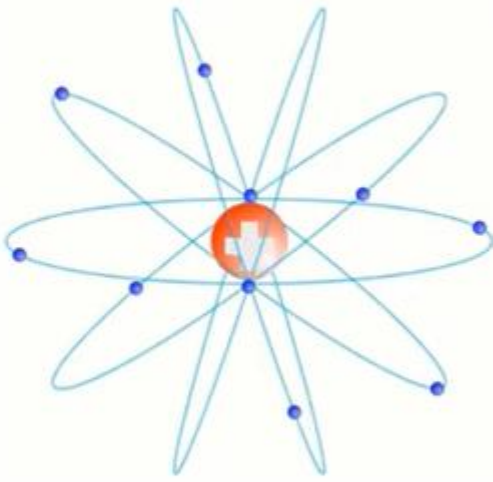
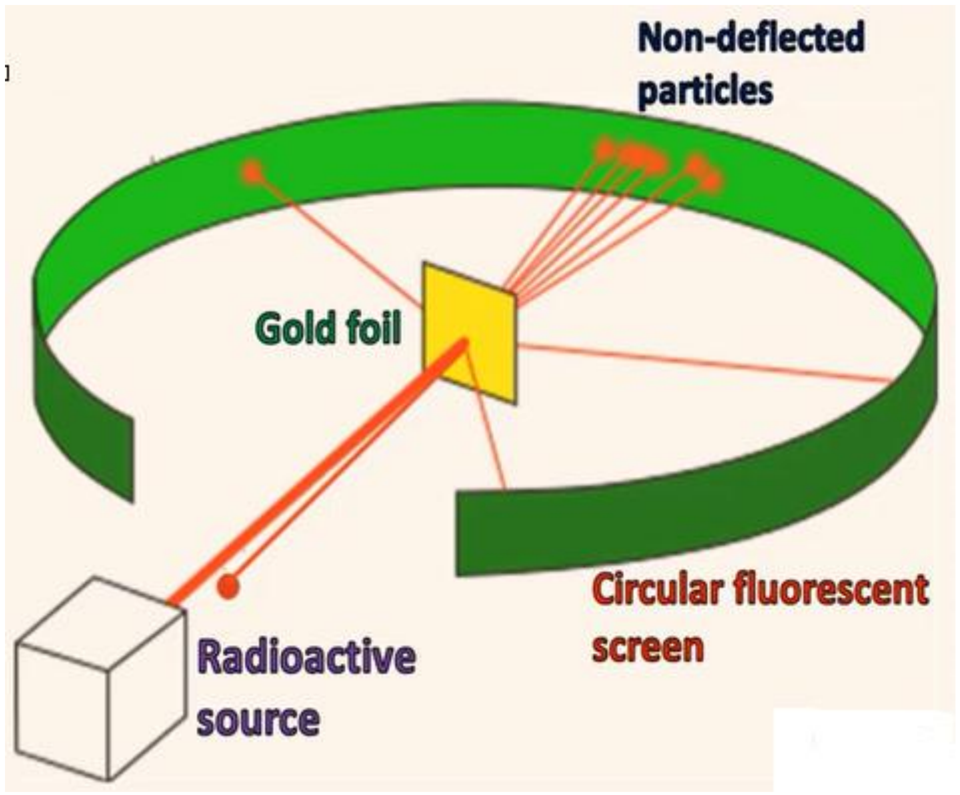


Fig. 2

- Introduce a brief history of the atom
- Explain Thomson's model of the atom (Plum pudding model)
- Explain the Gold foil experiment ("Geiger-Marsden experiment"
- Introduce Rutherford's explanation of the gold foil experiment
- Explain the planetary model (the nuclear model)
- Explain the inadequacy of Rutherford's model of the atom







Early models of the atom

Joseph John Thomson
(1856-1940)

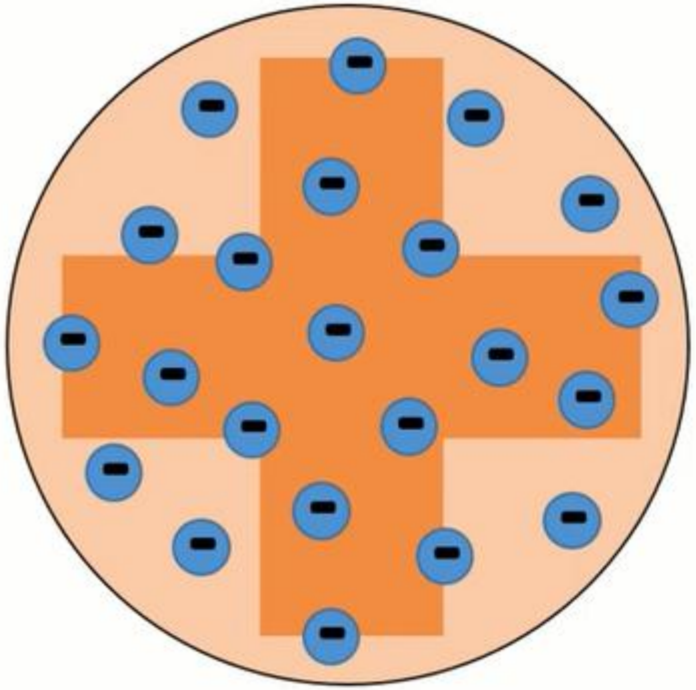
In 1879, Thomson discovered the electron and measured its charge to its mass.

Robert Andrew Millikan
(1868-1953)

In 1909, Millikan performed the oil drop experiment and determine the electric charge and the mass of the electron.

Thomson's model of the atom (Plum pudding model) - 1902

The atom is constituted of a sphere of positive charge, and the electrons are embedded in it like raisins in a spherical bun.



Joseph John Thomson
(1856-1940)

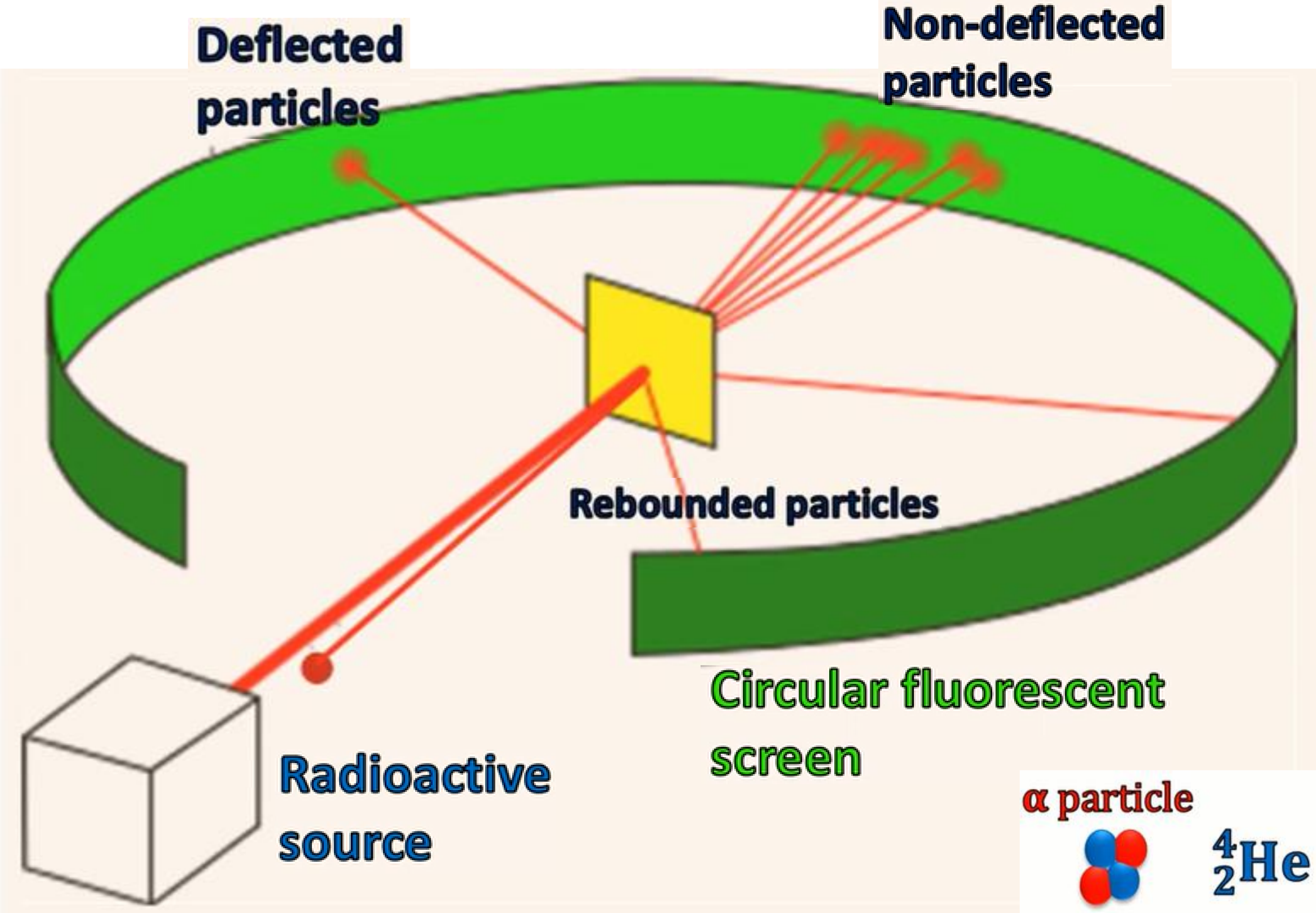
Gold foil experiment ("Geiger-Marsden experiment")

In 1911, Rutherford set evidence of the existence of the atomic nucleus in his famous experiment Gold foil experiment.

In this experiment, a beam of alpha particles which are positively charged helium nuclei emitted from a radioactive source towards a thin gold foil.

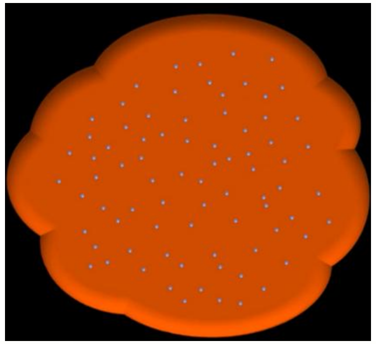
A circular fluorescent screen is used to detect the scattered alpha particles.

Rutherford was astonished with the results, he found that most of the fast moving alpha particles pass straight through the foil, some alpha particles were deflected through large angles, and few alpha particles rebounded.



Expected results according to Thomson's model

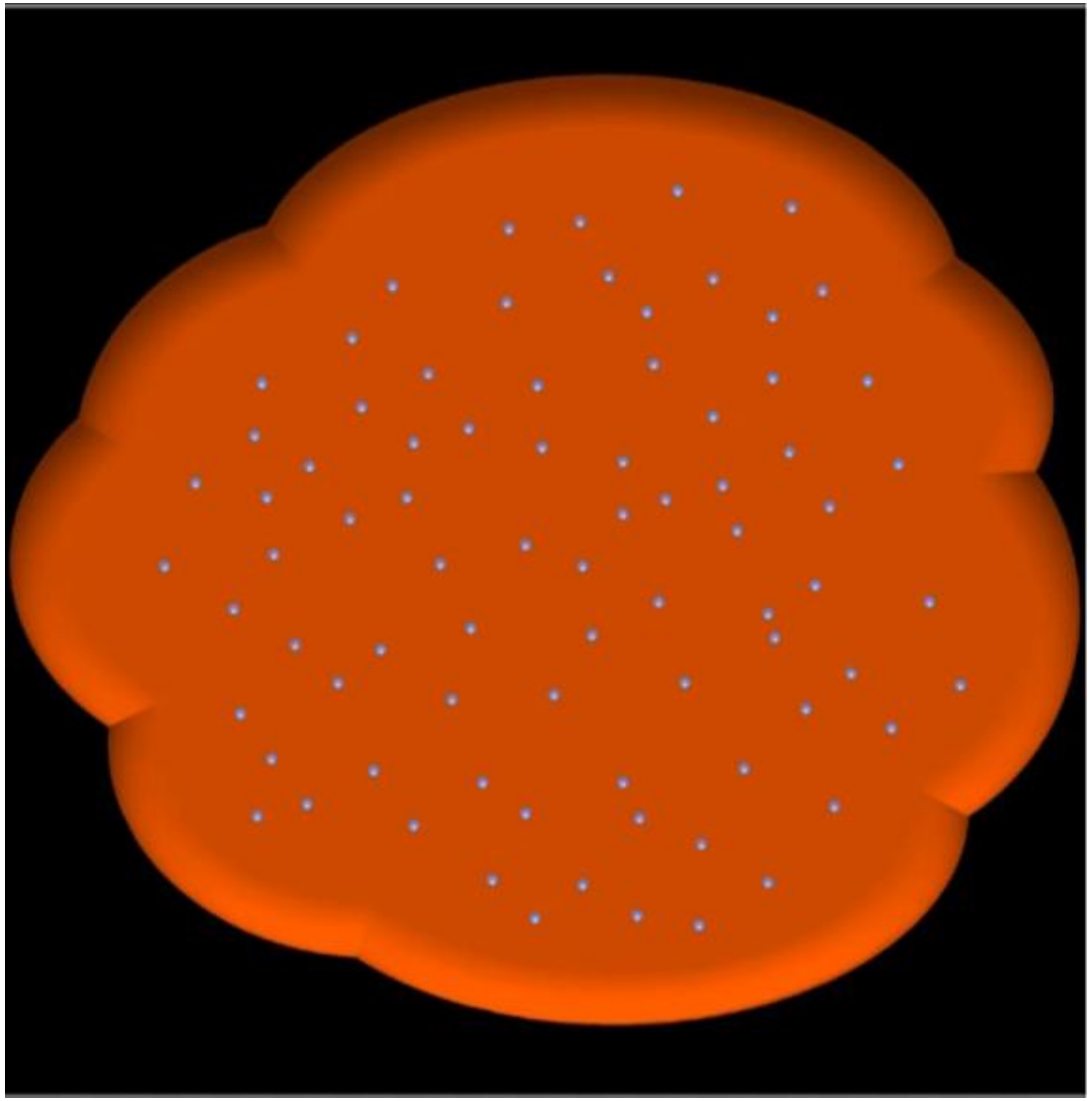
The mass and the positive charge of the atom are distributed over the entire atom.



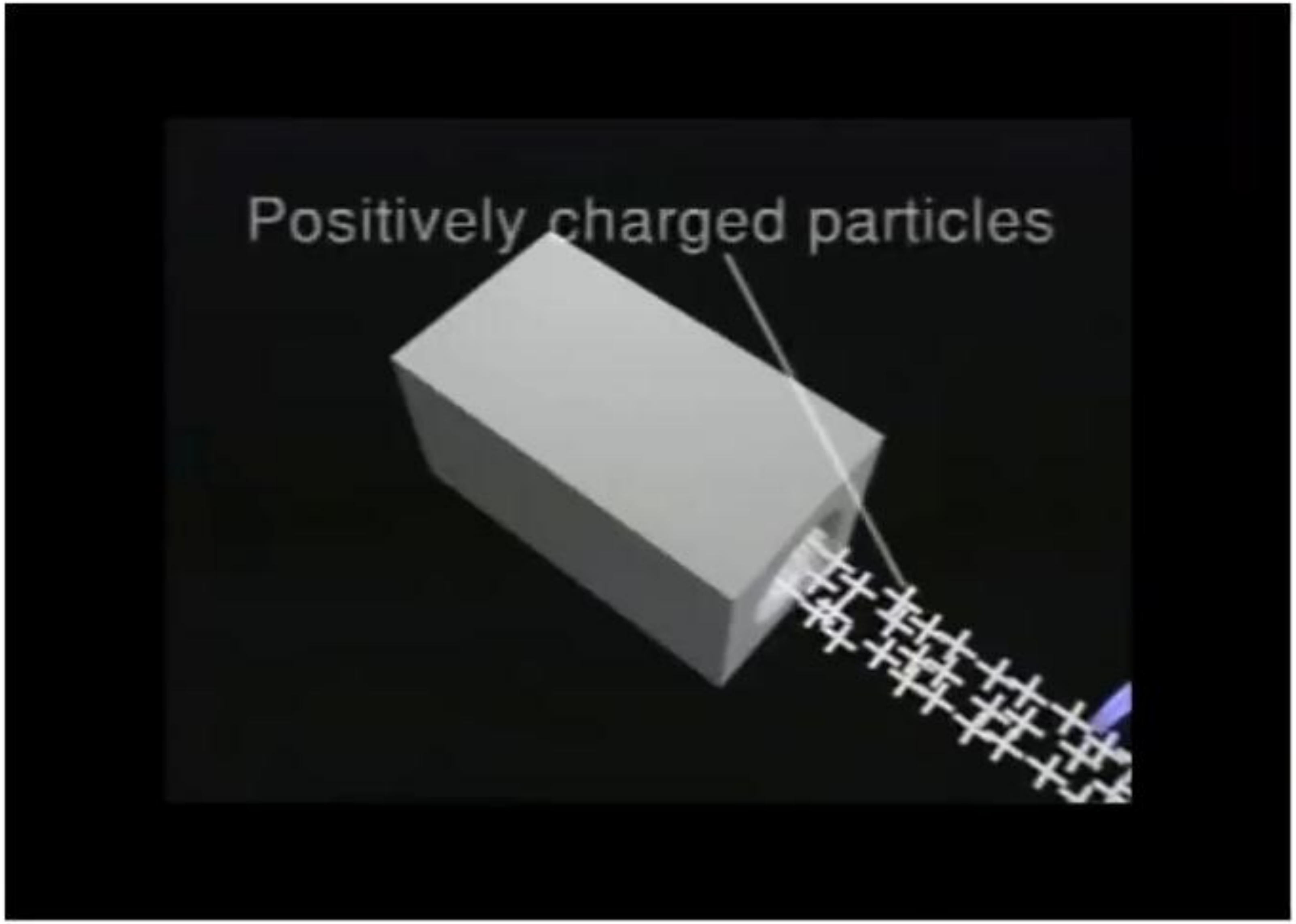
Then, **the concentration of positive charge is not strong enough to cause deflection of the heavy positively charged α particles.**

Therefore, **all α-particles must pass through the atom without deflection.**

But this did not happen!!

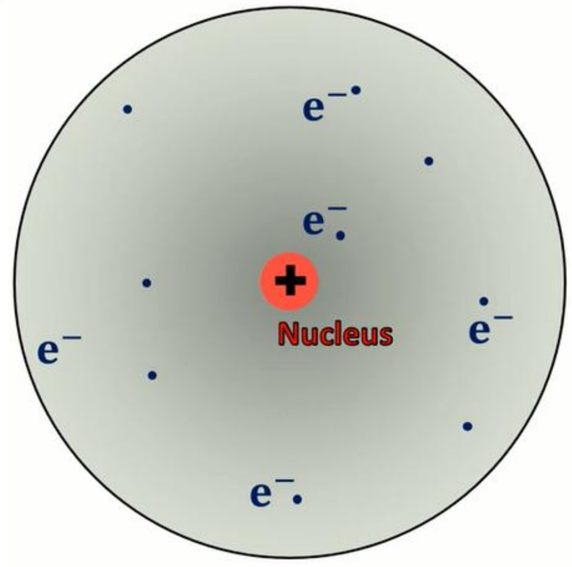


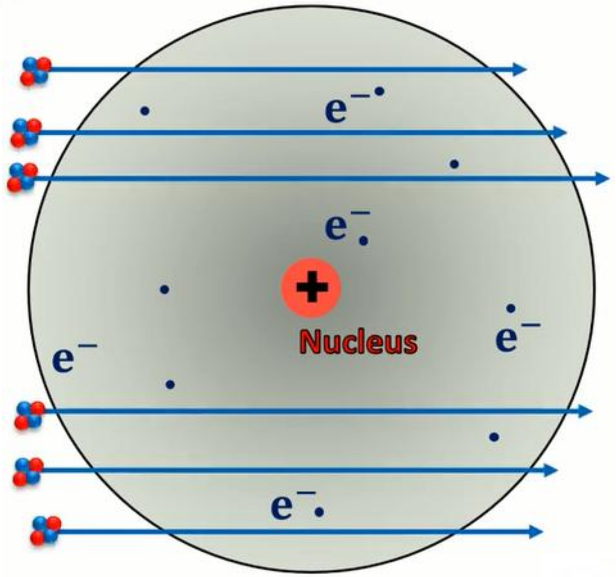
Experimental results



Rutherford's explanation of the gold foil experiment

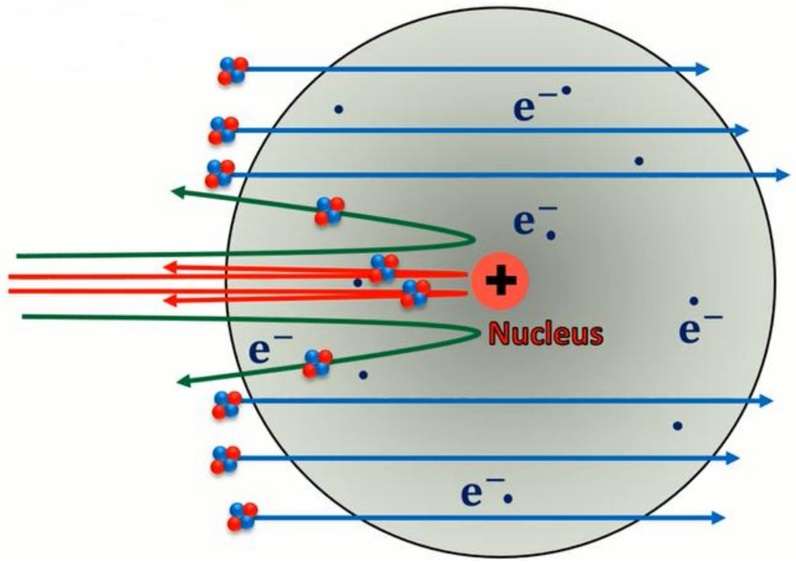
- Most of the α-particles passed straight through the foil, then **most of the atom is empty space.**





- Some particles were deflected through large angles.
  Very few particles rebounded.

  Then, **the atom has a center of concentrated positive charge which forms most of the mass of the atom.**



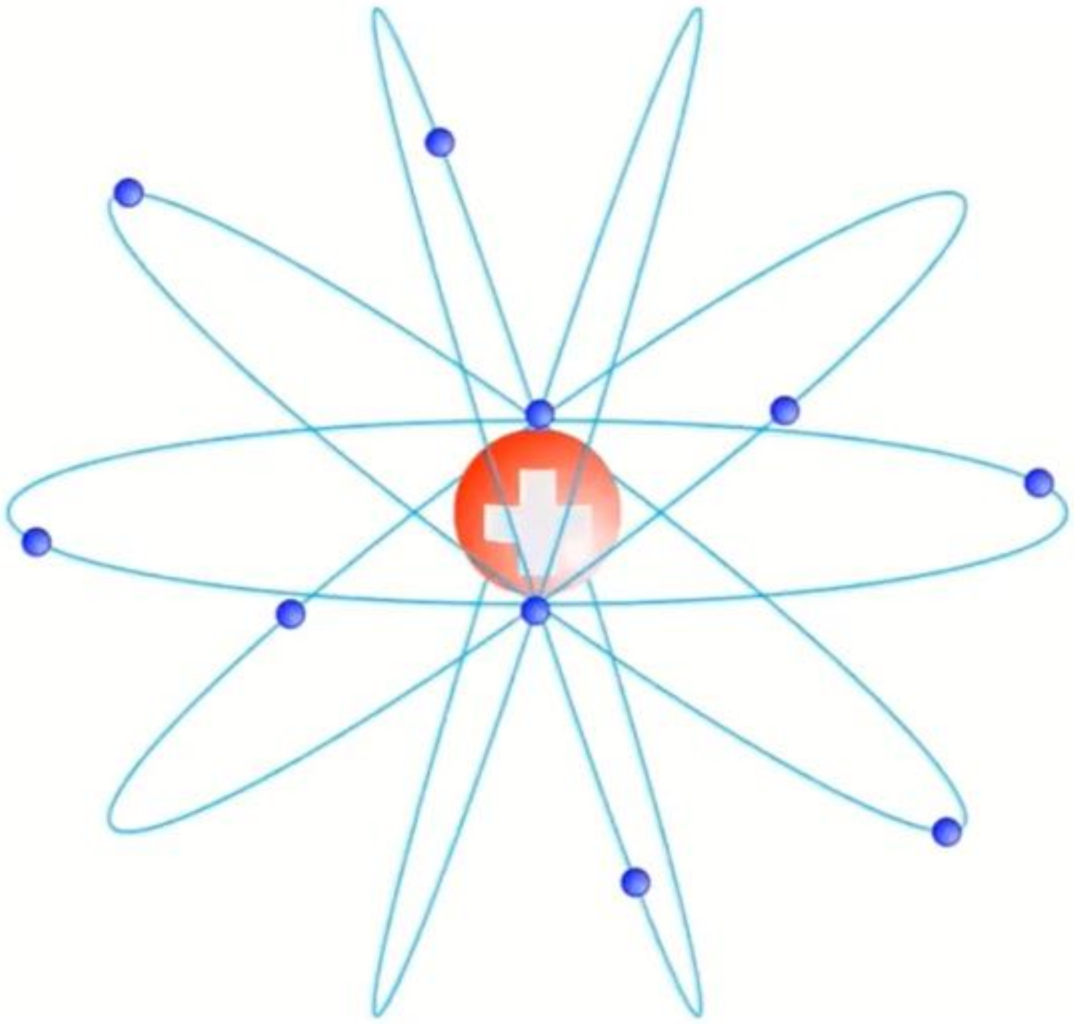
Center: **Atomic nucleus**

The planetary model (The nuclear model)

Based on the results of the gold-foil experiment, Rutherford put his model of the atom as planetary model.

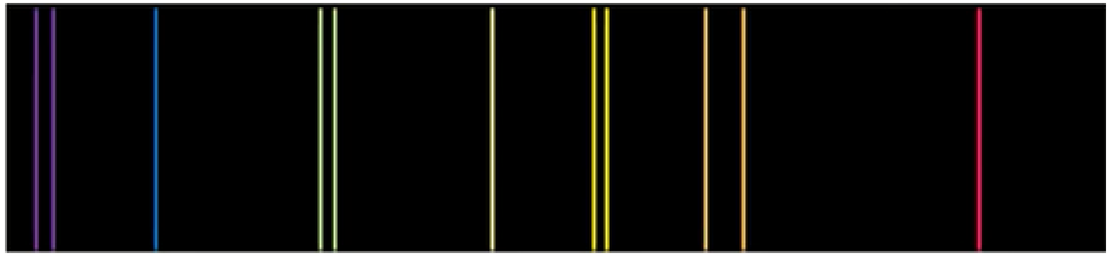
The atom is constituted of a positive nucleus concentrated in a very small space.

The electrons move in orbits around the nucleus in a motion similar to the revolution of a planet around the Sun.



Failure of Rutherford's atomic model

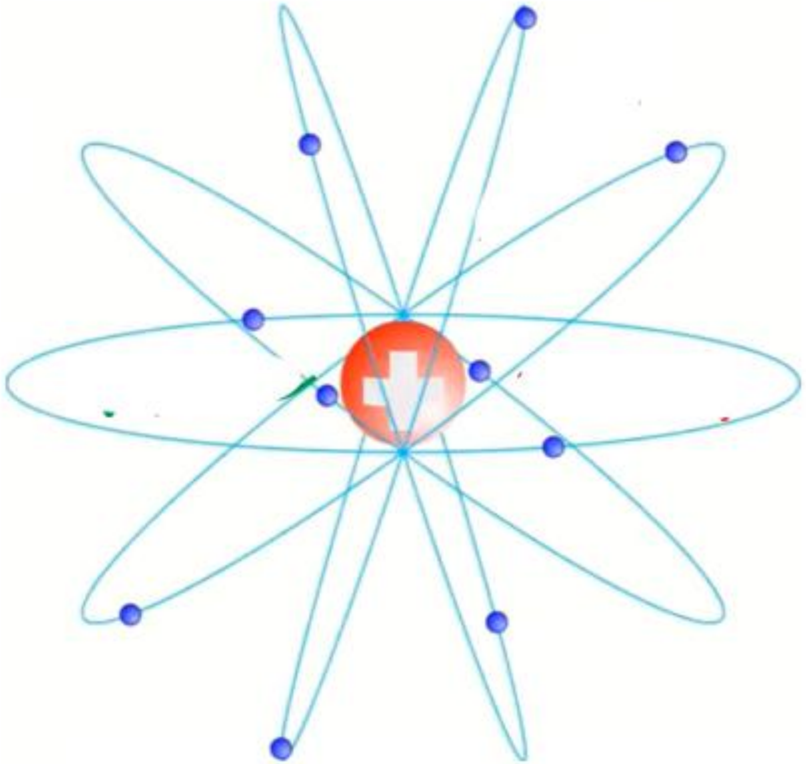
* Rutherford's model failed to explain the discrete line spectrum of the atom.



* According to Rutherford's model, the electron revolves around the nucleus, so it has a centripetal acceleration.

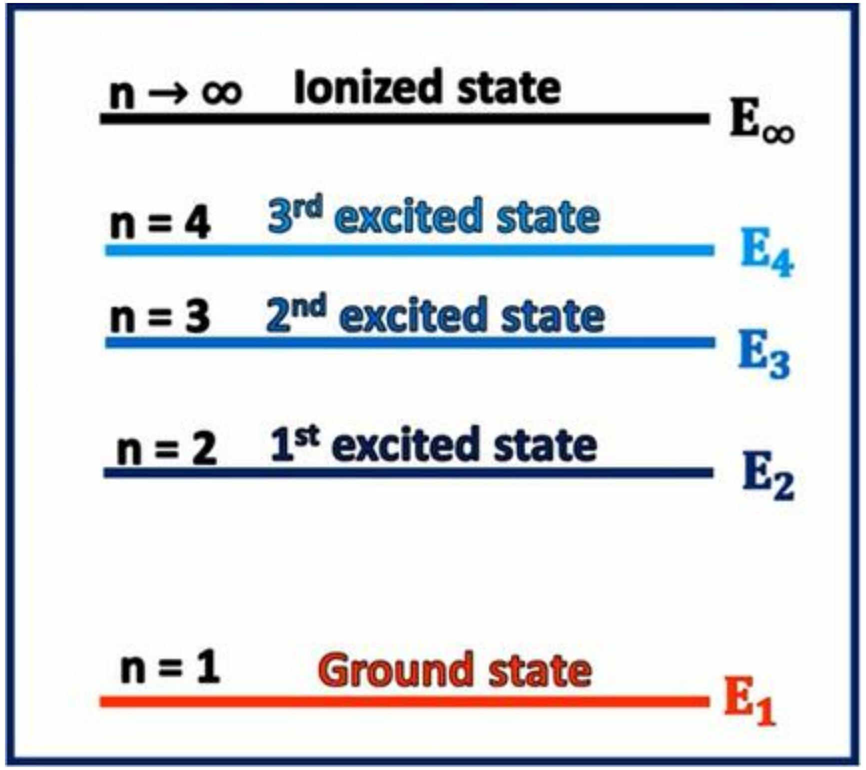
  But, accelerating charges radiate electromagnetic waves.

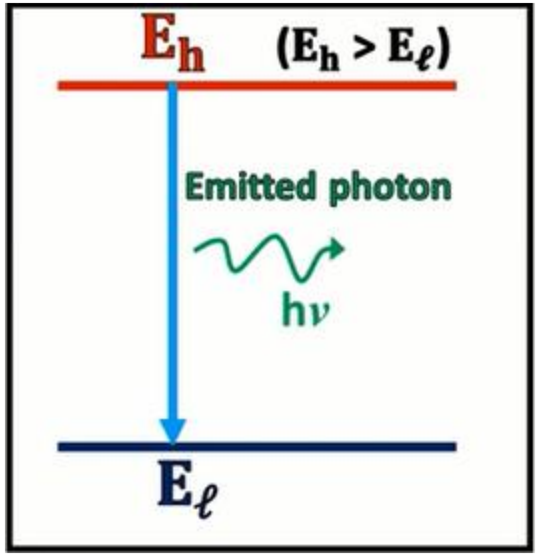
  So, the electron will lose its energy continuously.
  Finally, it will be captured by the nucleus.

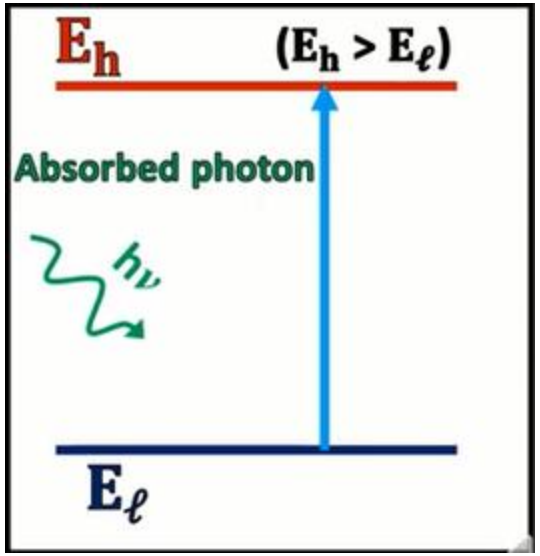


But this is not the case!

- *Explain the motion of the electron around the nucleus*
- *Identify the energy levels of the hydrogen atom, ground state, excited states, and ionized state*
- *Explain the electronic transitions of the atom*
- *Solve an application on Bohr model of the hydrogen atom*







Bohr model of the hydrogen atom (1913)

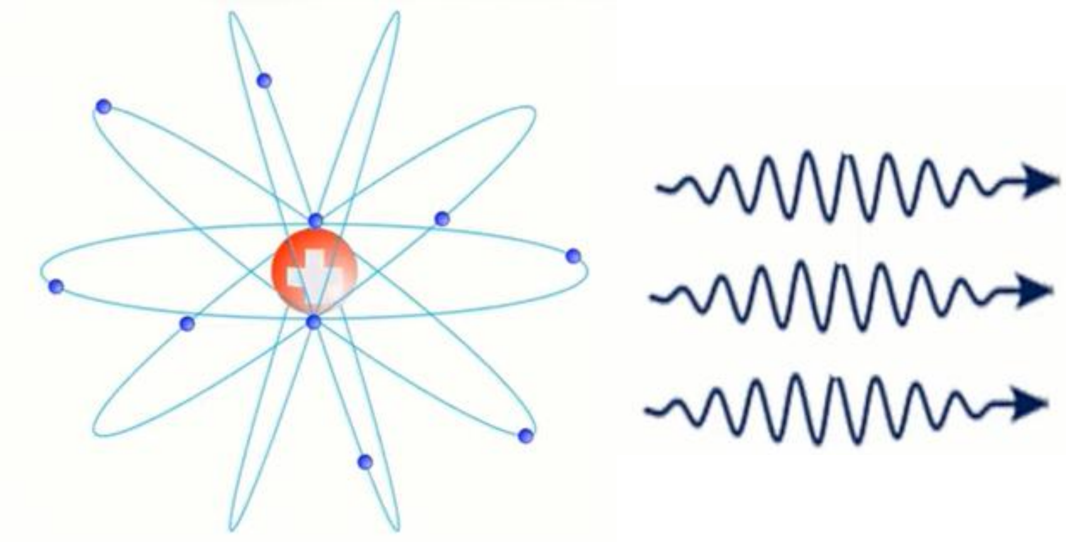
Joseph John Thomson
(1856-1940)

Niels Bohr (1885-1962)

Ernest Rutherford
(1871-1937)

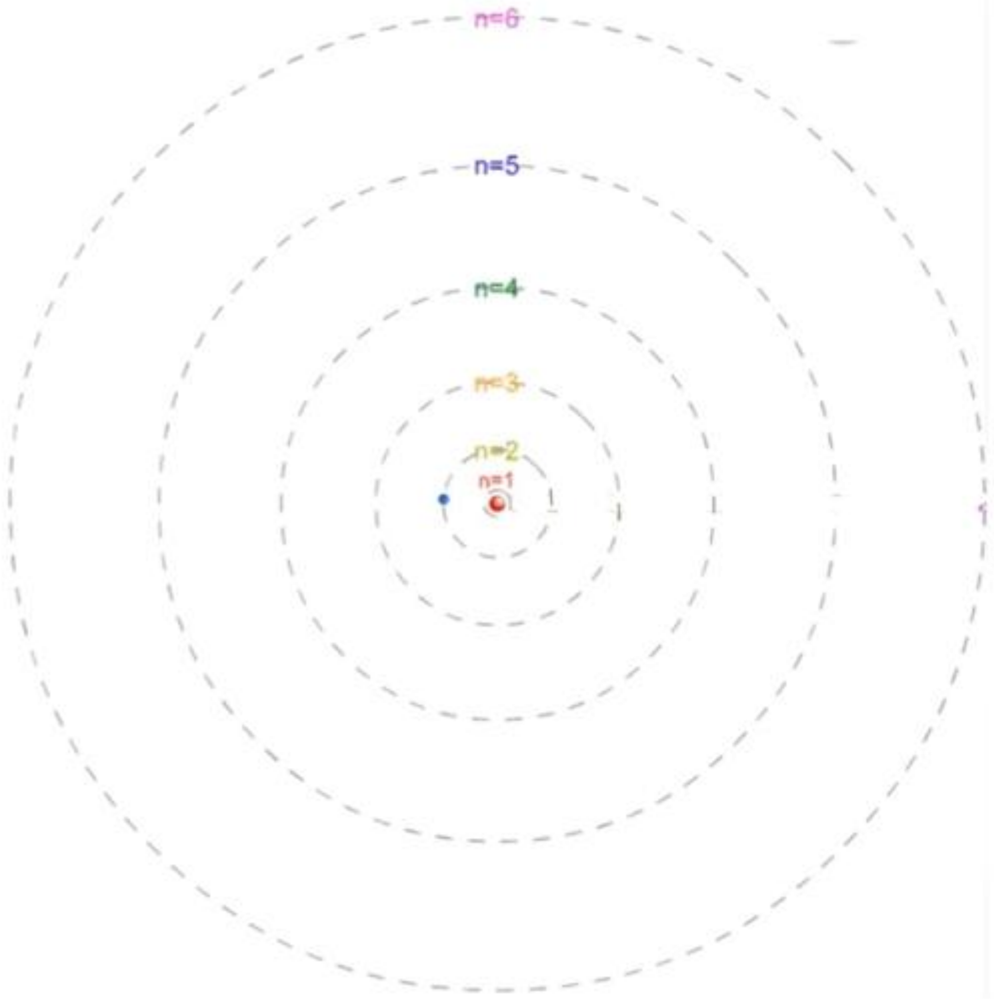
We have studied Thomson and Rutherford models of the atom.

Bohr's model for the hydrogen atom was based on his observation of the discrete emission spectrum of the hydrogen atom, Rutherford's planetary model of the atom, and on Plank's quantum hypothesis.



Bohr model of the hydrogen atom (1913)

- The electron moves in circular orbits around the nucleus.



We have seen that one of the reasons behind the failure of Rutherford atomic model is that according to that model, the accelerating electrons radiate electromagnetic radiations and captured finally by the nucleus.

How did Bohr solve this problem ?

Bohr postulated that:

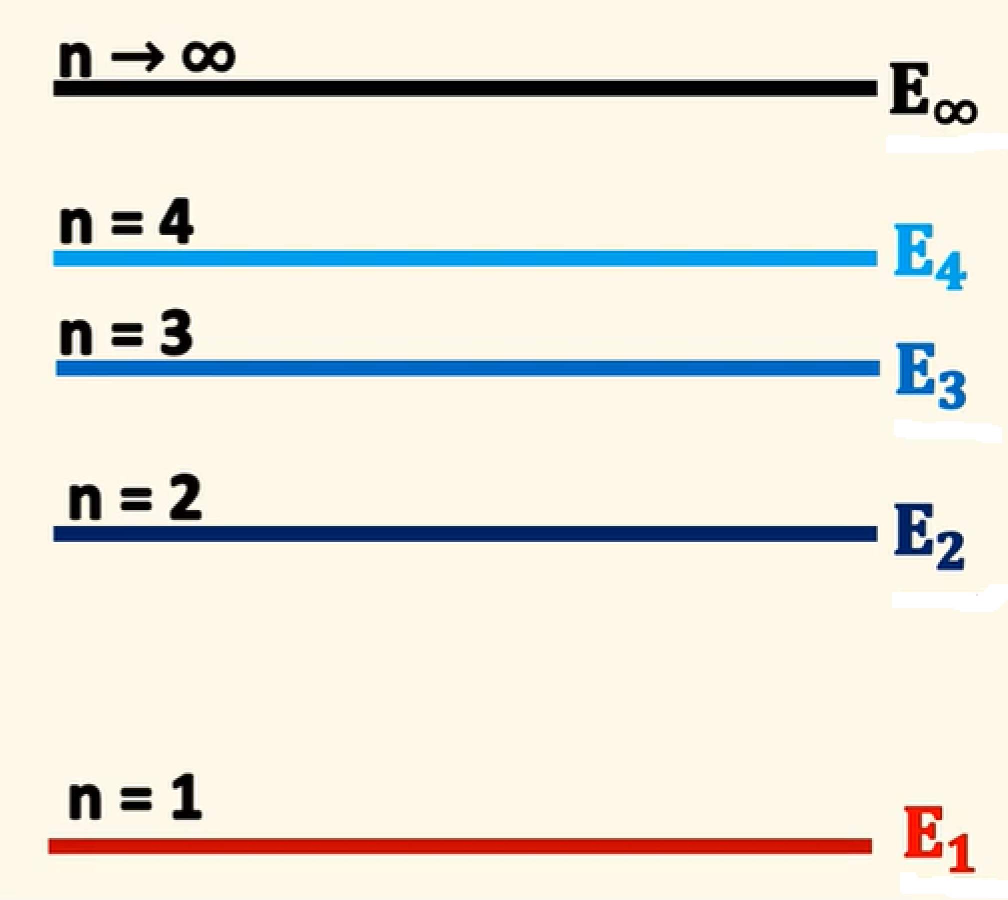
- **An isolated hydrogen atom has its own set of energy levels:**

  E_1 , E_2 , E_3 , ... , E_∞

  When the atom is in one of these levels, it does not emit radiation, even though the electron is accelerating.

  Each level is associated with a quantum number n.

  $n = 1, 2, 3, \ldots$



The ground state

The first energy level E_1 has the smallest energy.

It is called the ground state.

Its quantum number is $n = 1$.

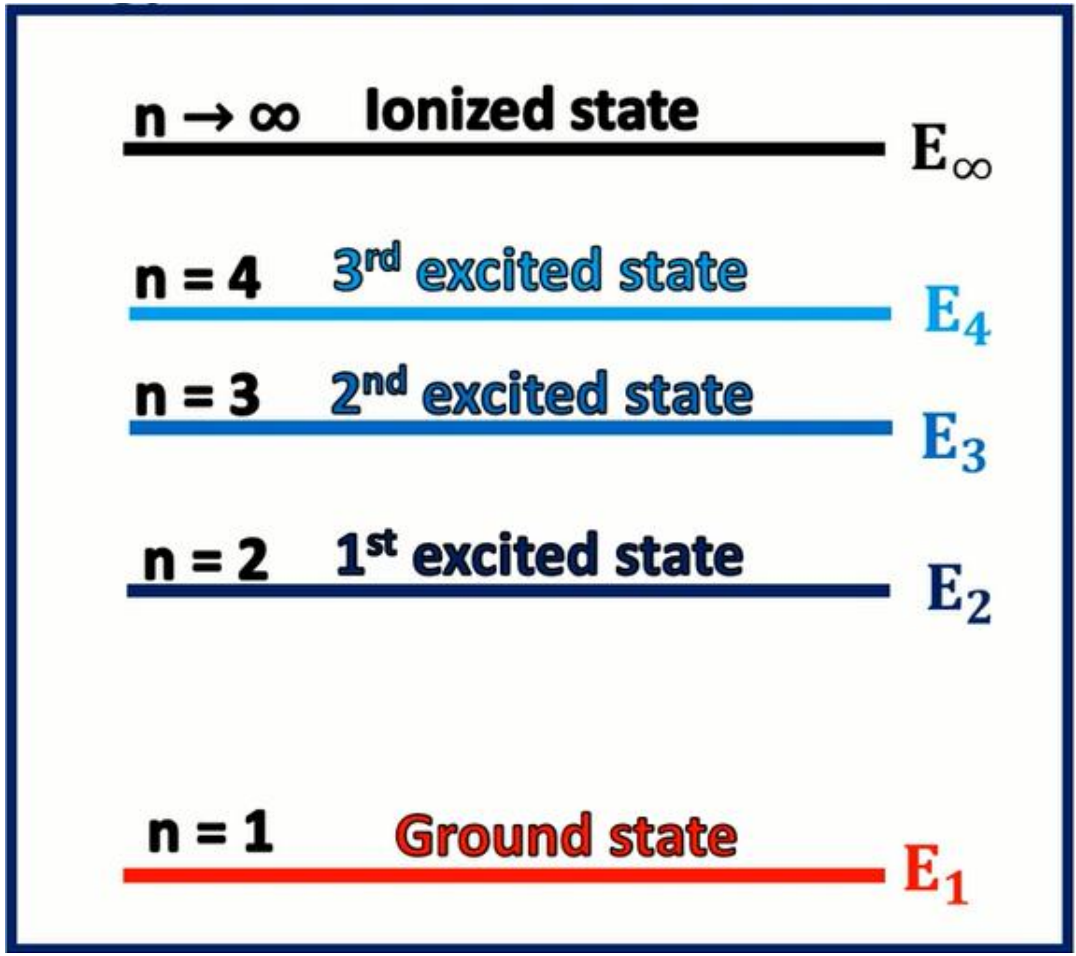
The excited states

The second energy level E_2 is called the first excited state.

The third energy level E_3 is called the second excited state, etc.

The ionized state

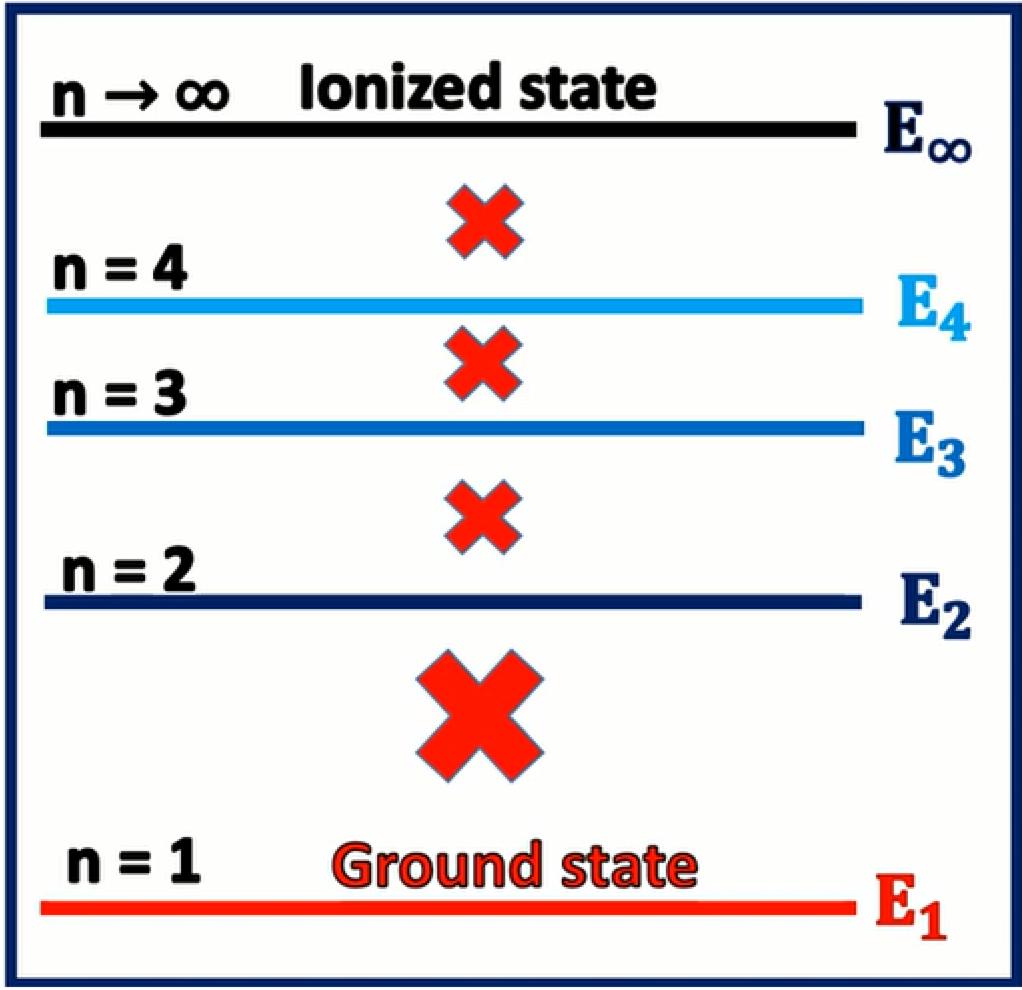
E_∞ is called the ionized state or the reference state.



The energy of the atom can be equal to one of these energy levels.

It is impossible for the atom to have an energy intermediate between two energy levels.

The energy of the atom is quantized.

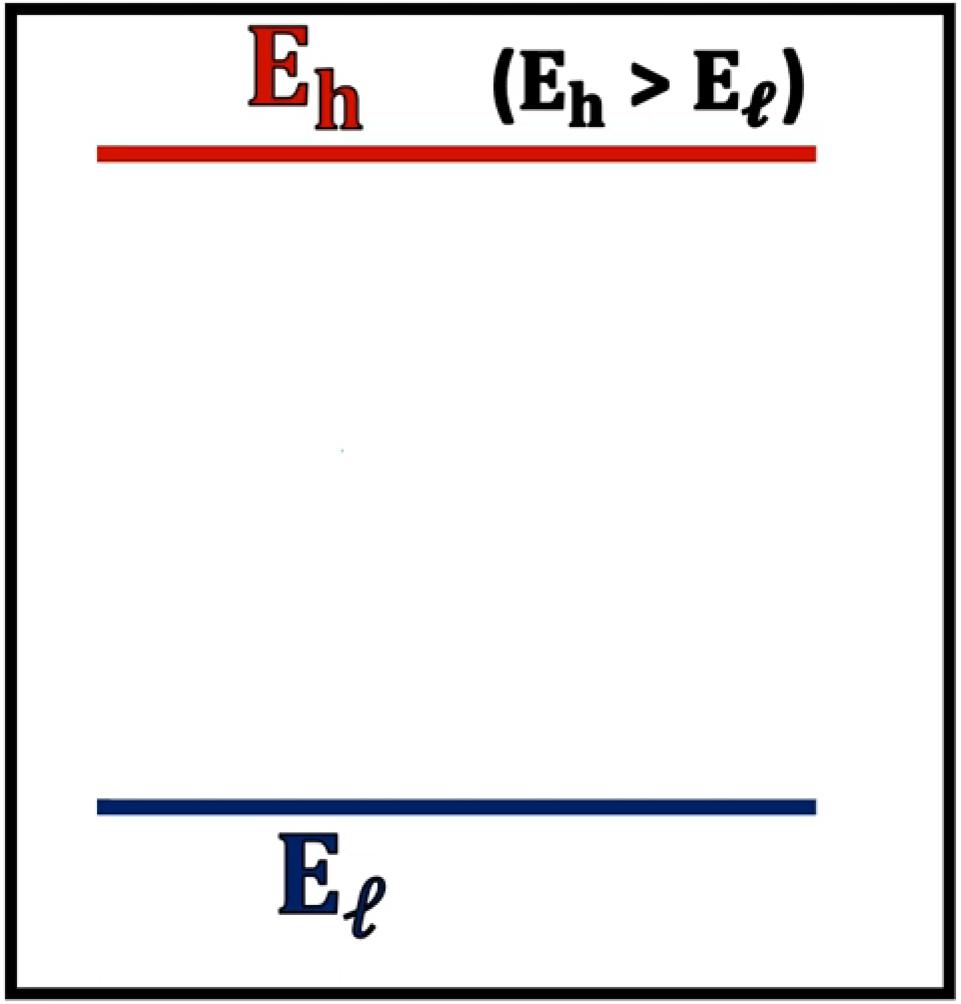


Quantized energy: Energy having only certain discrete values.

- **Emission of a photon**

When the atom makes a transition from a higher energy level, E_h, to a lower energy level, E_ℓ, it emits a photon.

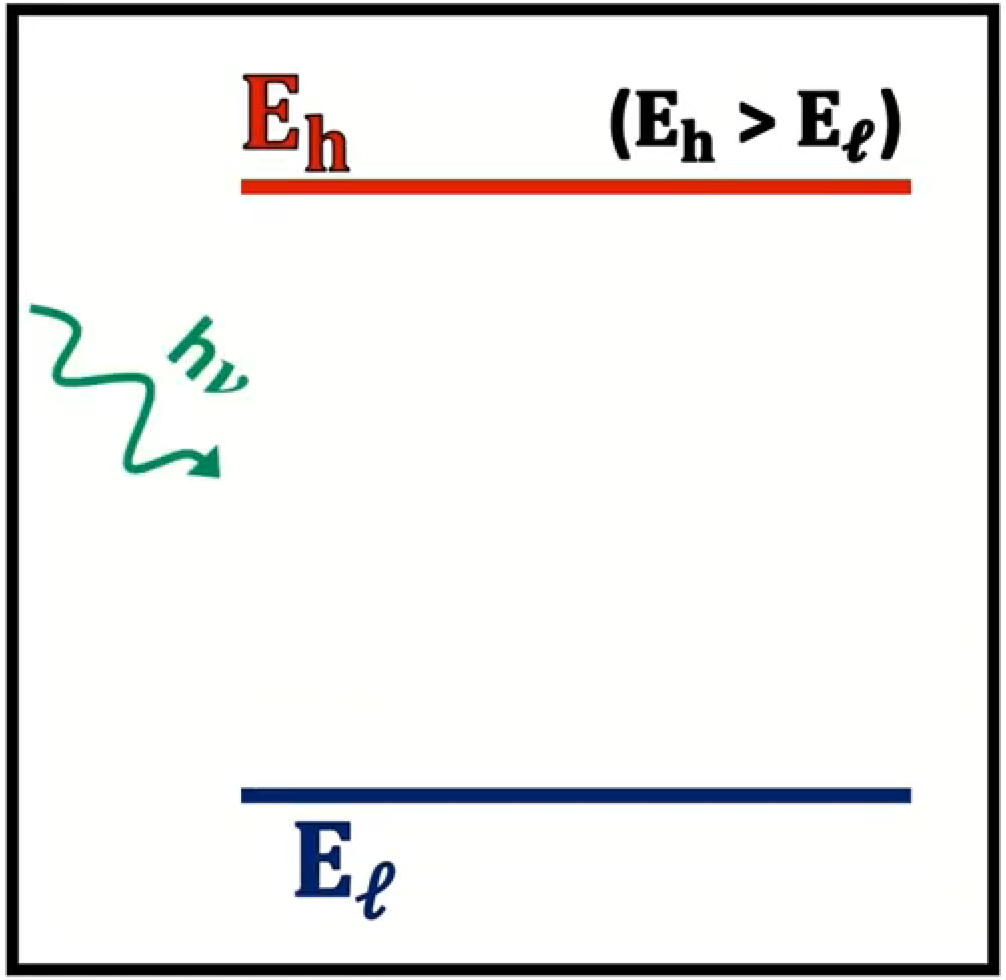
$$\boxed{E_{ph} = h\nu = \frac{hc}{\lambda} = E_h - E_\ell}$$



• Absorption of a photon

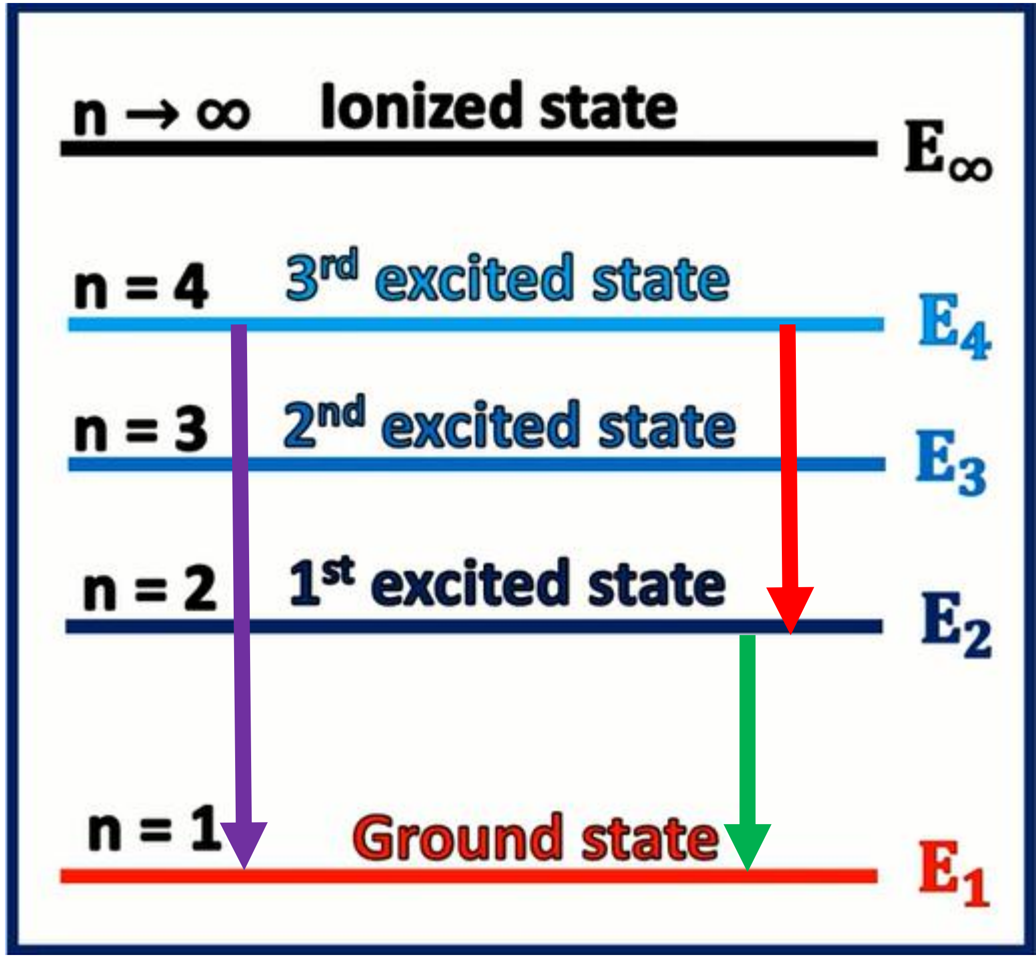
A photon can cause the transition of an atom from a lower energy level, E_ℓ, to a higher energy level, E_h, (excitation) **only** if its energy is **exactly** equal to the energy difference between E_h and E_ℓ.

$$E_{ph} = h\nu = \frac{hc}{\lambda} = E_h - E_\ell$$



When an atom is excited, it remains in its excited state few nanoseconds, and then it returns to the ground state.

- **Return directly by emitting one photon**
- **Return in several steps by emitting one photon in each transition.**

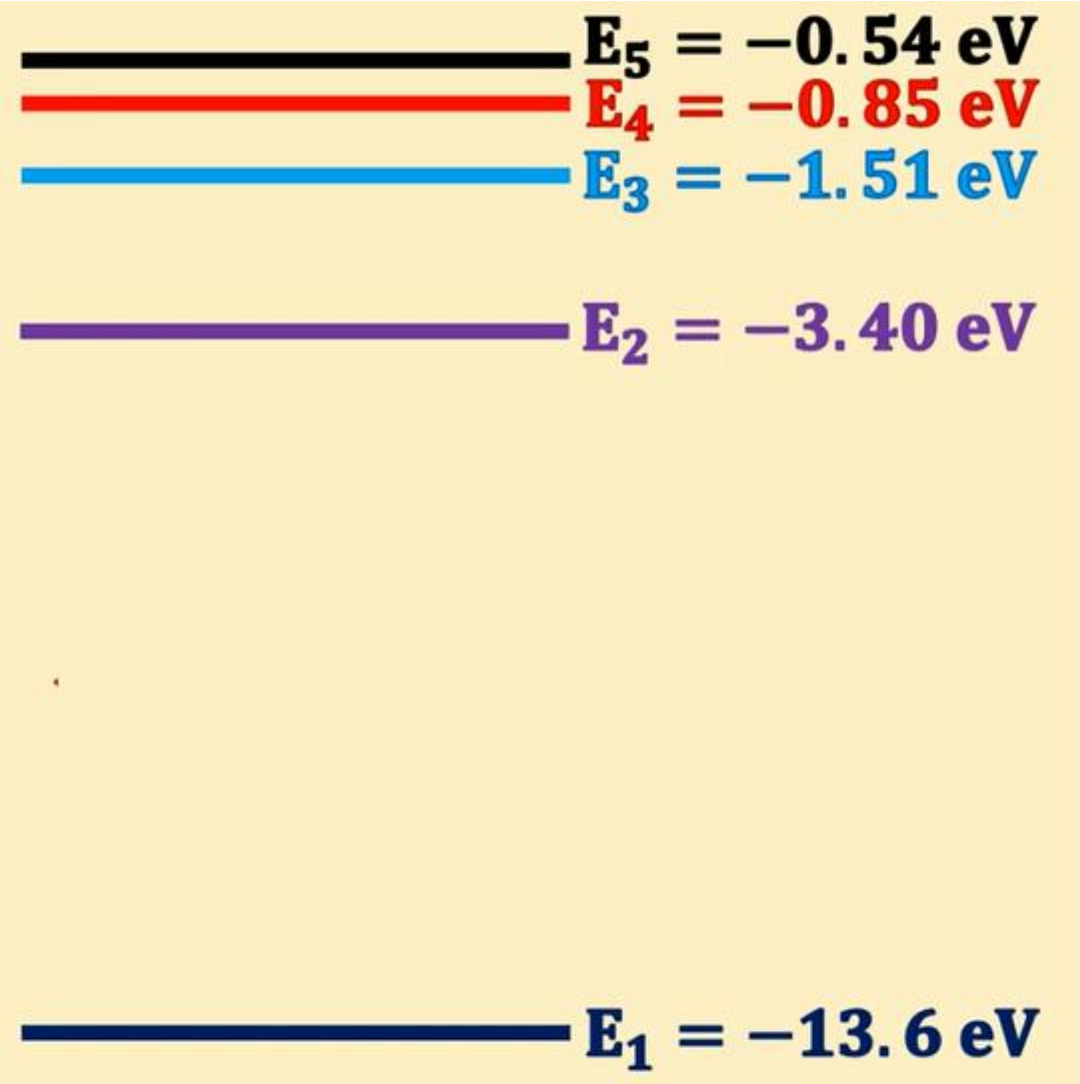


These postulates of Bohr model can be applied for other atoms.

Application

$h = 6.63 \times 10^{-34}$ J.s $\quad$ 1 eV $= 1.60 \times 10^{-19}$ J

1. Calculate the energy of the emitted photon when the atom makes a downward transition (electronic transition) from the energy level E_2 to E_1.



$$E_{ph} = E_2 - E_1 = -3.40 - (-13.6)$$

$$E_{ph} = 10.2 \text{ eV}$$

Application

$h = 6.63 \times 10^{-34}$ J.s

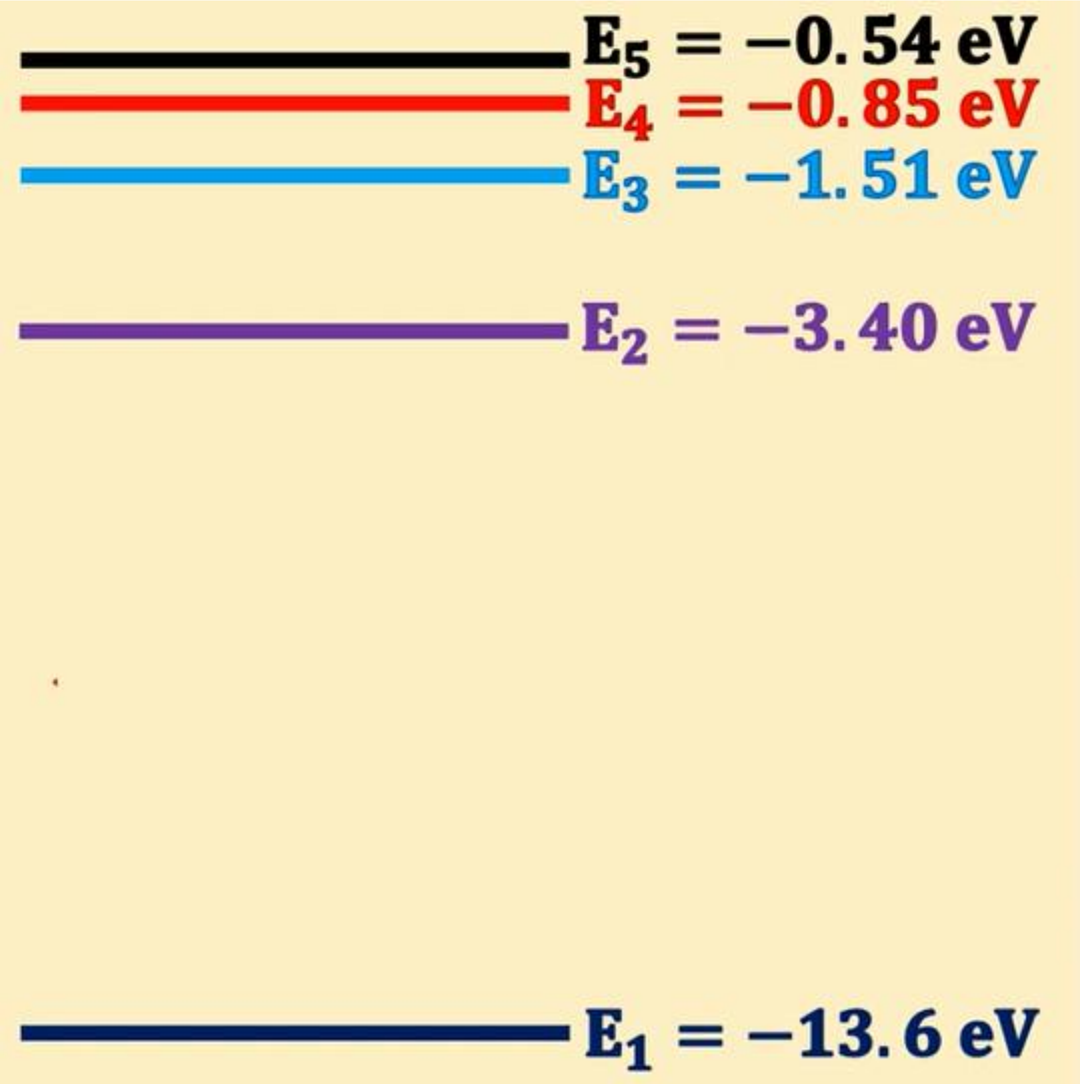
$1\ eV = 1.60 \times 10^{-19}$ J

2. Calculate the frequency of the emitted photon.

$E_{ph} = h\,\nu$, then $\nu = \dfrac{E_{ph}}{h}$

$$\nu = \frac{10.2 \times 1.60 \times 10^{-19}}{6.63 \times 10^{-34}}$$

$$\nu = 2.46 \times 10^{15}\ \text{Hz}$$



Application

$h = 6.63 \times 10^{-34}$ J.s $1\ eV = 1.60 \times 10^{-19}$ J

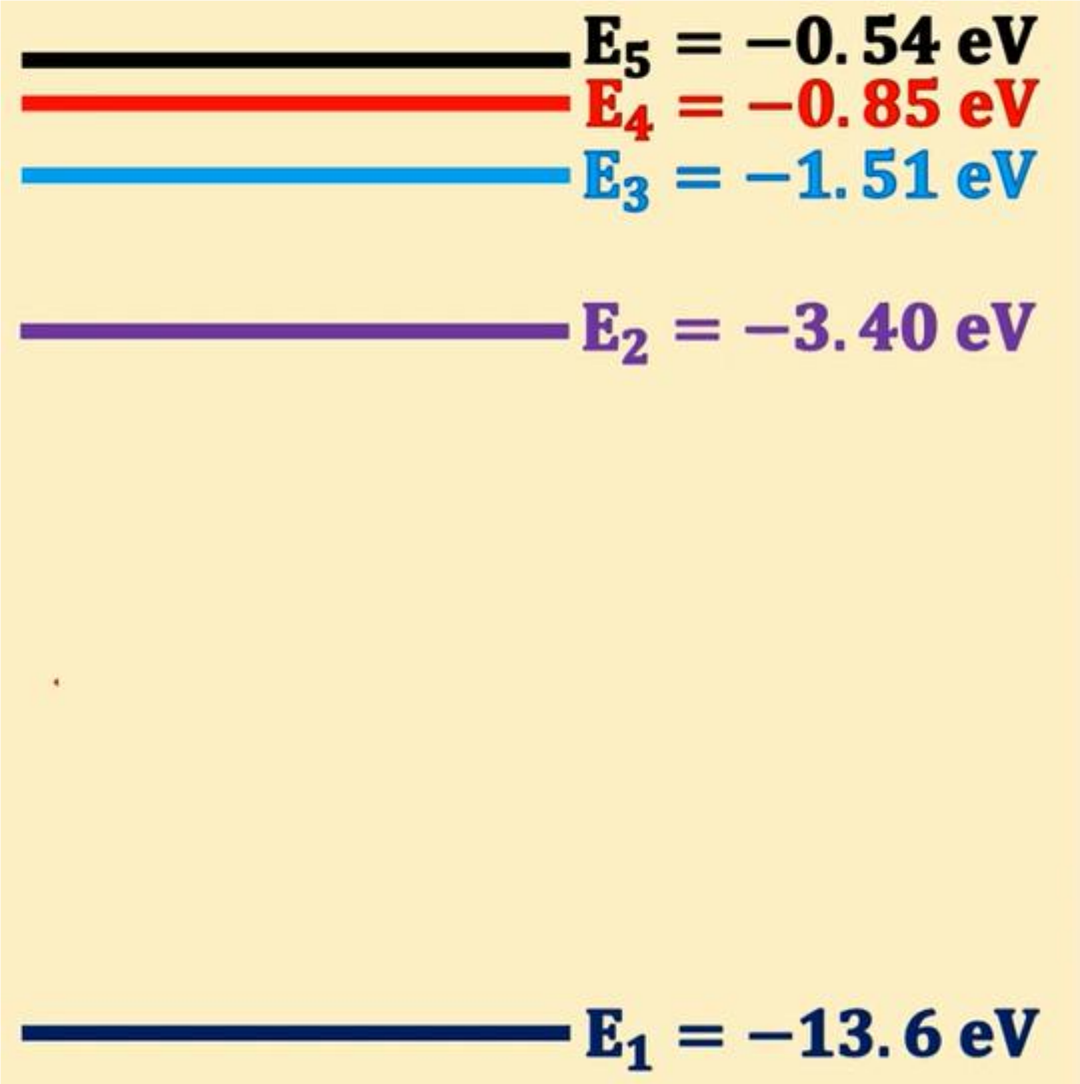
3. The hydrogen atom is in the ground state. The atom is hit by a photon of energy 8.00 eV. Specify whether this photon could be absorbed.

$$E_{ph} = E_h - E_\ell$$

$8.00 = E_h - (-13.6)$, so $E_h = -5.6$ eV

$$E_1 < E_h < E_2$$

Then, **the atom could not absorb this photon.**



Application

$h = 6.63 \times 10^{-34}$ J.s

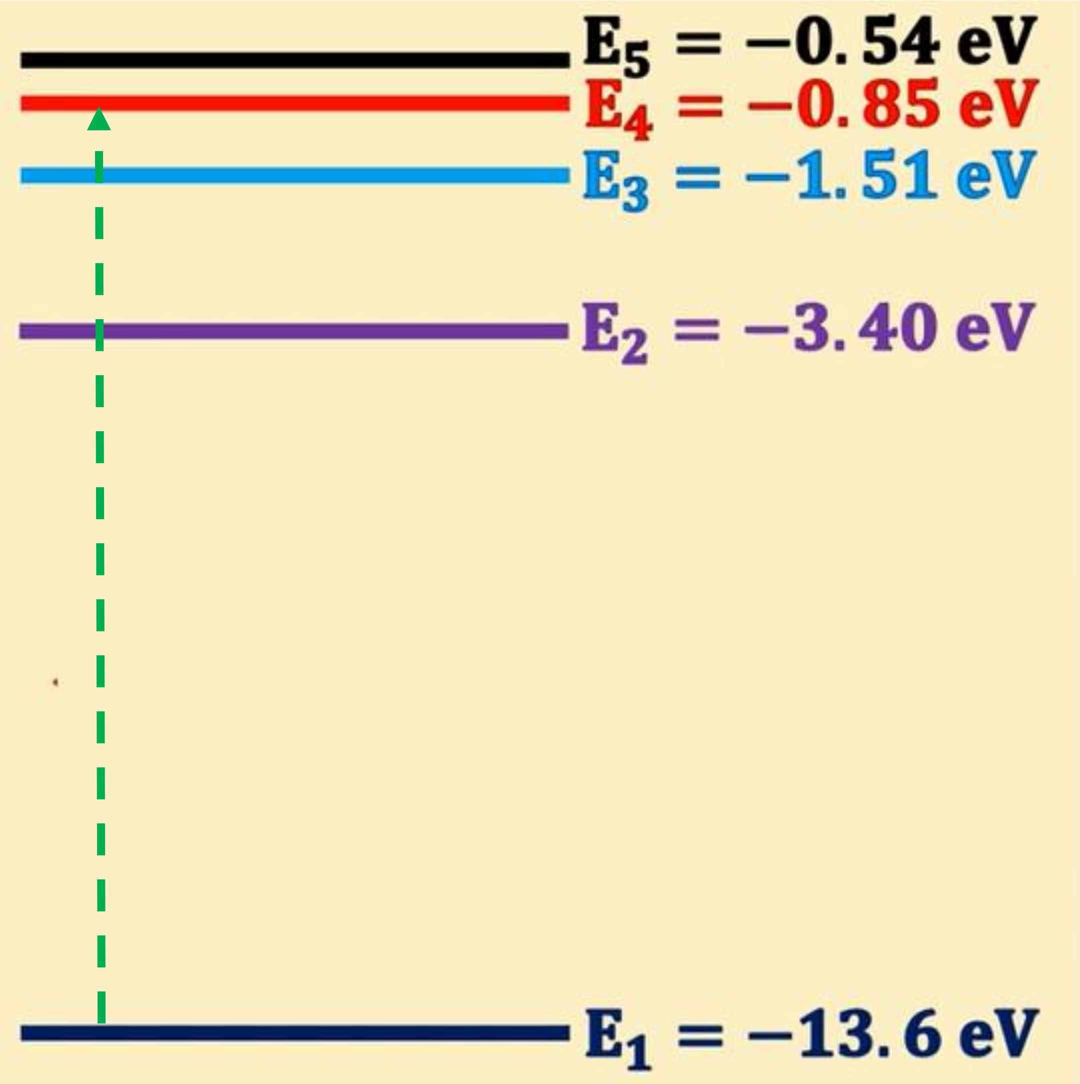
$1 \text{ eV} = 1.60 \times 10^{-19}$ J

4. Calculate the energy of a photon capable of exciting the atom to the 3^{rd} excited state when it is in the ground state.

$$E_{ph} = E_4 - E_1$$

$$E_{ph} = -0.85 - (-13.6)$$

$$E_{ph} = 12.75 \text{ eV}$$



Application

$h = 6.63 \times 10^{-34}$ J.s

$1\ eV = 1.60 \times 10^{-19}$ J

5. The atom is in the third excited state. Indicate all the possible downward transitions of the atom.

- $E_4 \longrightarrow E_3$
- $E_3 \longrightarrow E_2$
- $E_2 \longrightarrow E_1$
- $E_4 \longrightarrow E_2$
- $E_4 \longrightarrow E_1$