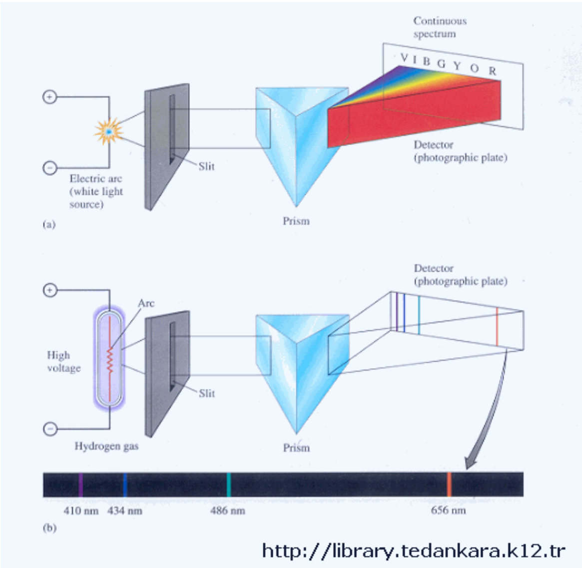
Topic:	THE BOHR MODEL OF THE ATOM	
General topic	**Objective ID**	**Objective**
FK_B1_03 **Bohr's atomic theory and atomic emission spectra**	**FK_B1_03_02**	**Given** the energetic values of the electronic orbits in the case of the hydrogen atom **determine:** • all the possible electronic transitions • the order of those electronic transitions according to their relative energies (without calculations) • the amount of energy released in a given electronic transition • the frecuency of the light emitted (using Planck's equation) • the region of the light emitted (infrared, visible, ultraviolet) **using** Planck's constant and the relationships between period, frequency and wavelength

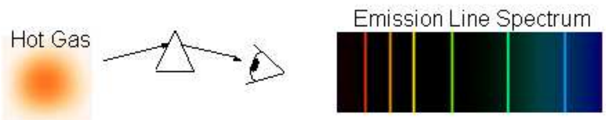
Introduction

The light emitted when an element is vaporized and then thermally or electrically excited, does not create a continous spectrum; instead, it creates a **line spectrum**, a series of fine lines of individual colors separated by black spaces.



The emission spectra of elemental atoms provide information about their structures and properties. Emission occurs from atoms that have been given a lot of energy by a spark or a flame.

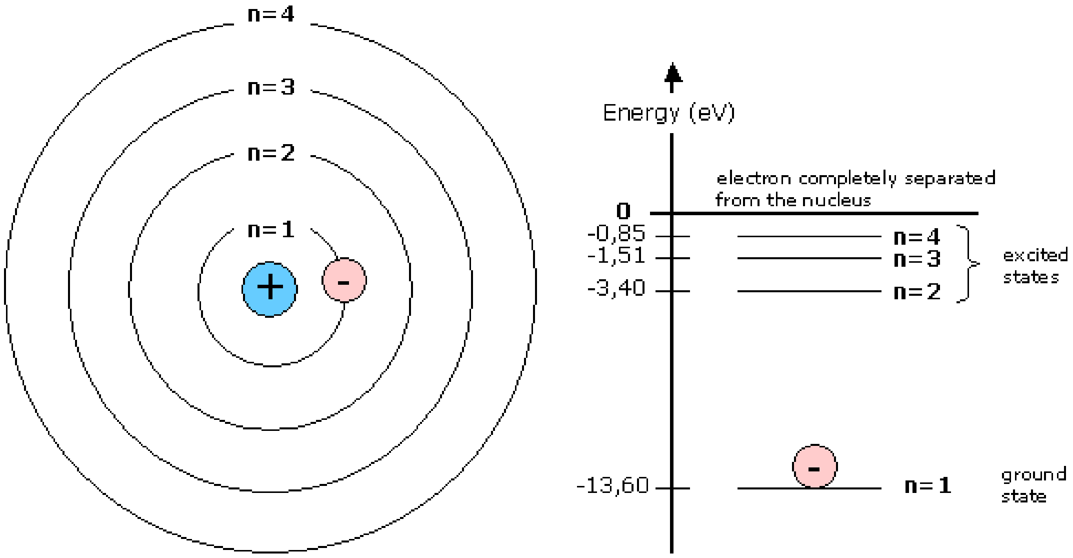
Provided that every element has its own specific atomic spectrum, it means that atomic spectra is connected to the structure of atoms.



The Bohr Model of the Hydrogen Atom

Bohr proposed a model for the H (hydrogen) atom that explained the line spectra of the atom. The model proposed by Bohr is based on the following postulates:

- **The H atom has only certain circular orbits. Each orbit has a specific energy value (level).**

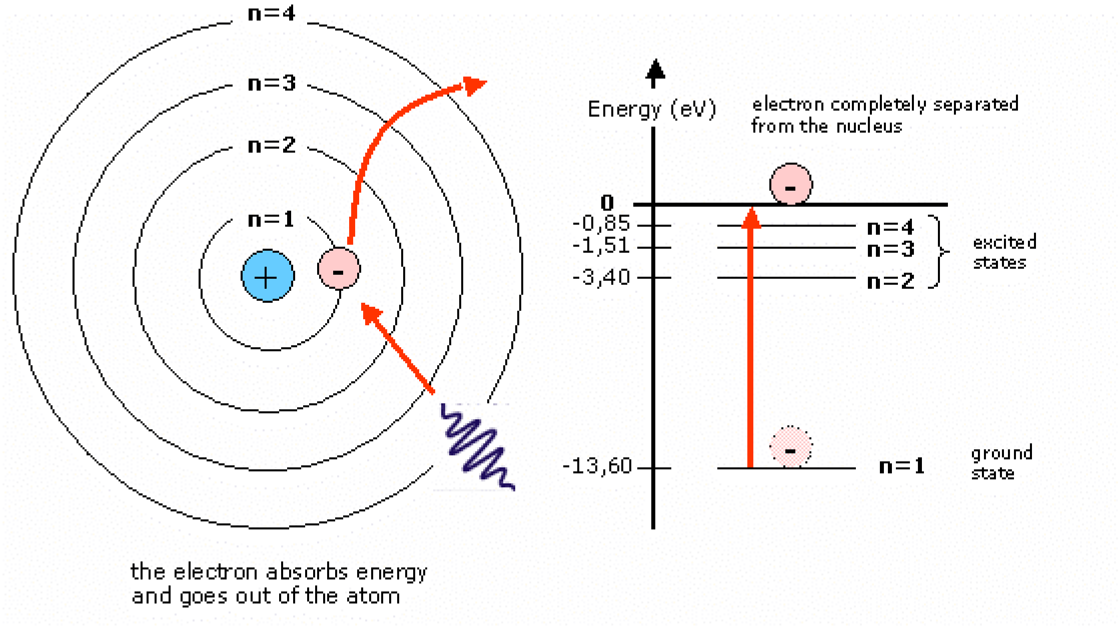


The orbits are identified by integers: n=1 (the closest to the nucleus), n=2, n=3, n=4, n=5...

Without absorbing energy from outside, the electron of any hydrogen atom will be at ground state (n=1). If the electron absorbs energy, it can skip to an excited state (n=2, n=3..., depending on the amount of energy absorbed)

The energy levels are determined as follows:

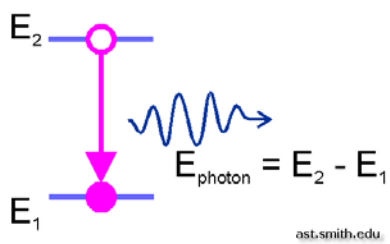
a) the reference is energy=0, when the electron is completely removed from the nucleus
b) the energy of the orbits is negative (the electron needs energy to skip out of the atom)
c) the energy of the n=1 orbit is the most negative (the case when the electron needs most energy to escape the atom; the nucleus exert the most powerful attraction force)



- **The atom does not radiate energy while the electron is placed in one of the circular orbits.**

- **When the electron moves to another orbit, it absorbs (when it goes to an outer orbital) or emits (when it goes to an inner orbital) energy as a photon (light).**

The energy of the emitted photon is the difference between initial (E_2) and final (E_1) energy levels of the electron.



Therefore, we can calculate the frequency of the emitted light as follows:

$$E_{photon} = E_2 - E_1 = h.f_{emitted\ light}$$

$$f_{emitted\ light} = \frac{E_2 - E_1}{h}$$

and check if that frequency exists in the line spectrum of the hydrogen atom.

The previous equation can be seen as an application of the conservation of the energy:

$$E_{initial} = E_{final} \rightarrow E_{2\ (initial\ of\ the\ electron)} = E_{1\ (final\ of\ the\ electron)} + E_{\ photon}$$

$$E_2 = E_1 + h.f_{emitted\ light} \rightarrow h.f_{emitted\ light} = E_2 - E_1$$

Worked exercise

Determine the frequency and wavelength of the light emitted by an electron (of an hydrogen atom) when it skips from n=2 to n=1.

Determination of the energy of the photon

The energy released as a photon must be equal to the difference in energy between both states (n=2, initial and n=1, final). Therefore:

$$E_{photon} = E_{(n=2)} - E_{(n=1)} = -3,40\ eV - (-13,60\ eV) = 10,20\ eV$$

$$E_{photon} = 10,20\ eV \frac{1,6.10^{-19}\ J}{1\ eV} = 1,632.10^{-18}\ J$$

Determination of the frequency

Using the Planck's equation, we can find the frequency:

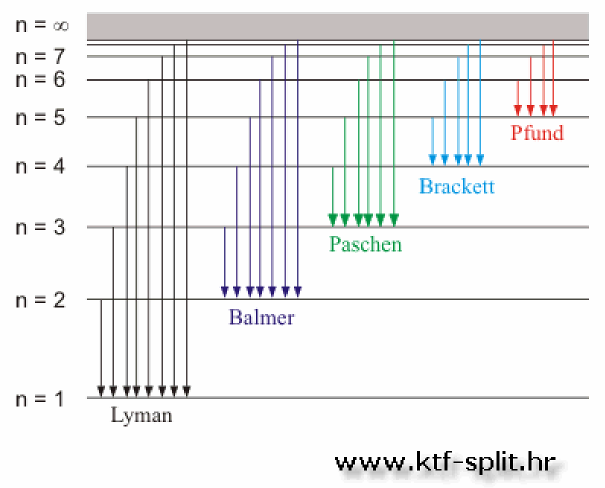
$$E = h.f \rightarrow 1,632.10^{-18}\ J = 6,63.10^{-34}\ J.s\ f \rightarrow$$

$$f = \frac{1,632.10^{-18}\ J}{6,63.10^{-34}\ J.s} = 2,46.10^{15}\ Hz$$

Determination of the wavelength

$$c = \lambda.f \rightarrow \lambda = \frac{c}{f} = \frac{3.10^{8}\ m/s}{2,46.10^{15}\ Hz} = 1,22.10^{-7}\ m = 122\ nm$$

Line spectrum of the hydrogen: ultraviolet (Lyman) series, visible (Balmer) series and infrared (Paschen) series.



When a sample of gaseous H atoms (remember that there are millions and millions of atoms in a sample) is excited, different atoms absorb different quantities of energy.

Even though each atom has only one electron, so many atoms are present that all the energy levels (orbits) are populated by electrons.

When the electrons drop from the outer orbits to the n=3 orbit the emitted photons create the **infrared series** of spectral lines. An infrared light has a wavelength longer than 750 nm.

The **visible series** (**Balmer**) arises when electrons drop to the n=2 orbit. A visible light has a wavelength between 400 nm (violet) and 750 nm (red).

The **ultraviolet series** (**Lyman**) arises when electrons drop to the n=1 orbit. An ultraviolet light has a wavelength shorter than 400 nm.

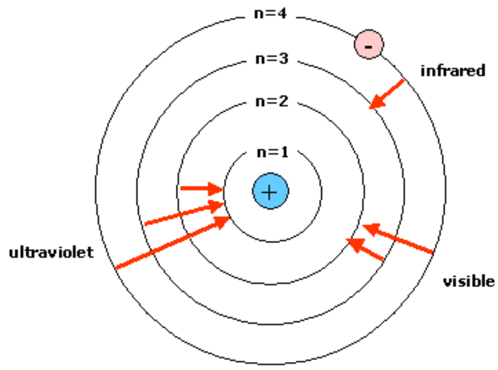
Worked example

Specify all the possible electronic transitions from n=4 downwards (towards n=1) and place them in order.

ultraviolet	visible	infrared

All the possible transitions are:

- ultraviolet region (Lyman):
    - o n=4 → n=1
    - o n=3 → n=1
    - o n=2 → n=1
- visible region (Balmer):
    - o n=4 → n=2
    - o n=3 → n=2
- infrared region:
    - o n=4 → n=3



Taking into account that a bigger jump means a more energetic photon (more towards the left side), we can conclude that the spectral lines will appear as shown below: