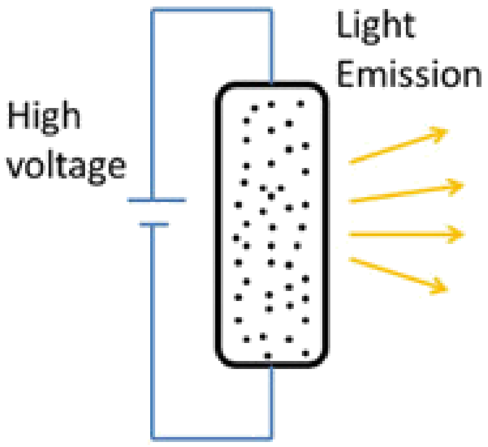
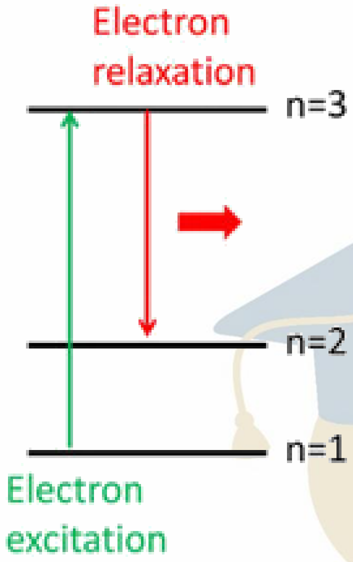
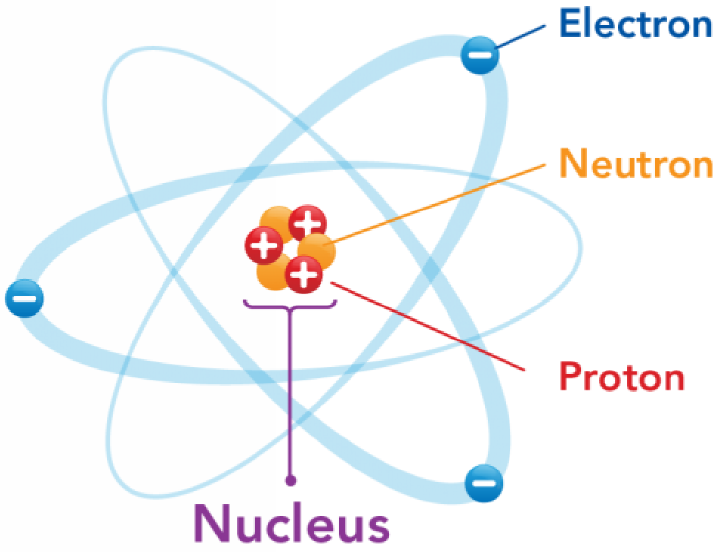






Chapter 17

The atom

Prepared & Presented by: Mr. Mohamad Seif

OBJECTIVES

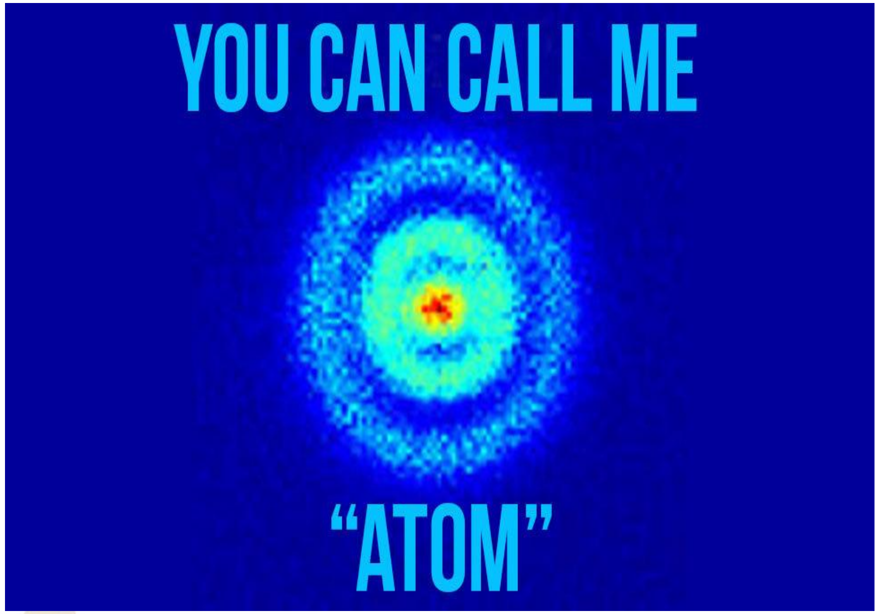
1 Introduction

2 Models of the atom

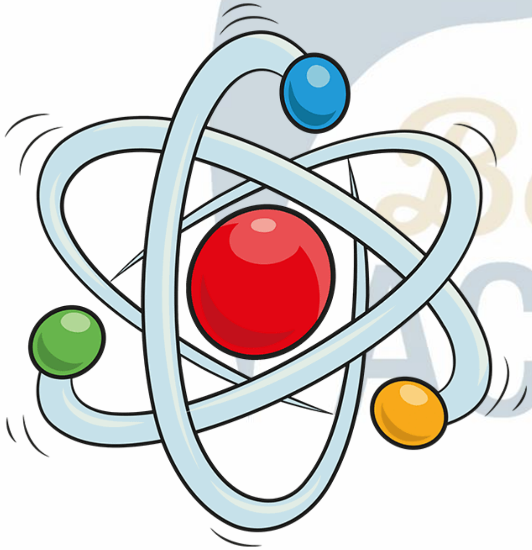
3 Emission and absorption of a photon

Introduction

The atom: **The atom was the smallest indivisible constituent of matter.**



Later discoveries proved that it can be divided into a great number of smaller particles

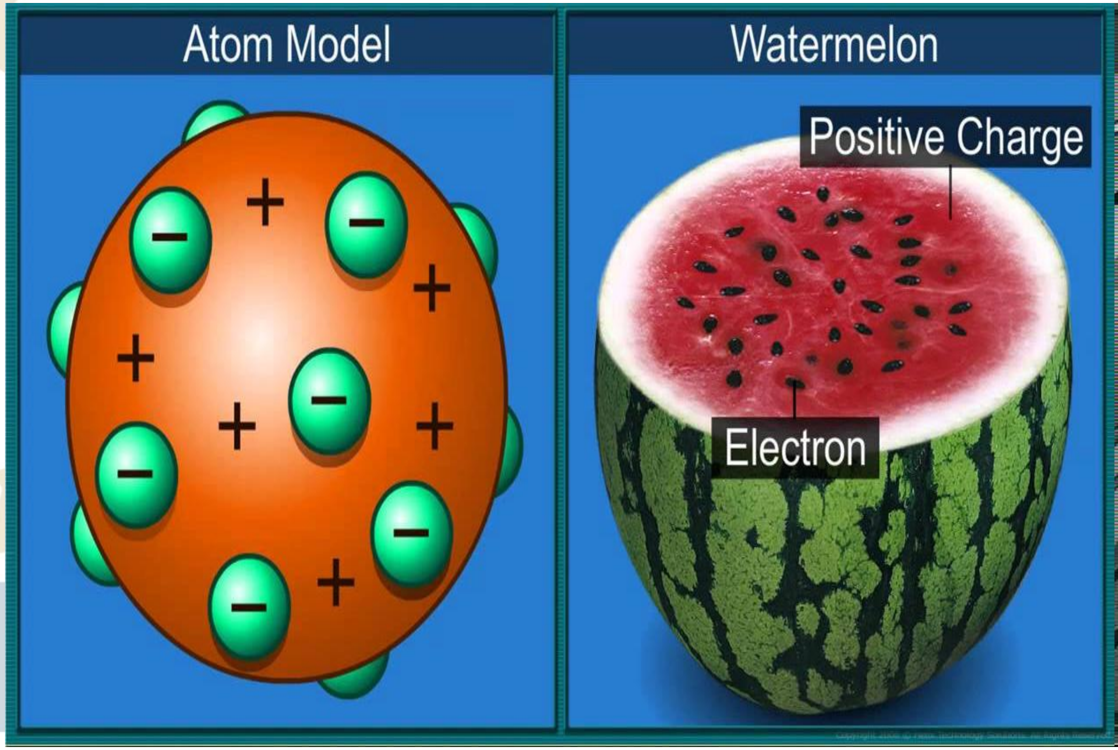


Models of the atom

Thomson model of the atom:

Thomson proposed that an atom is constituted of a sphere of positive charge.

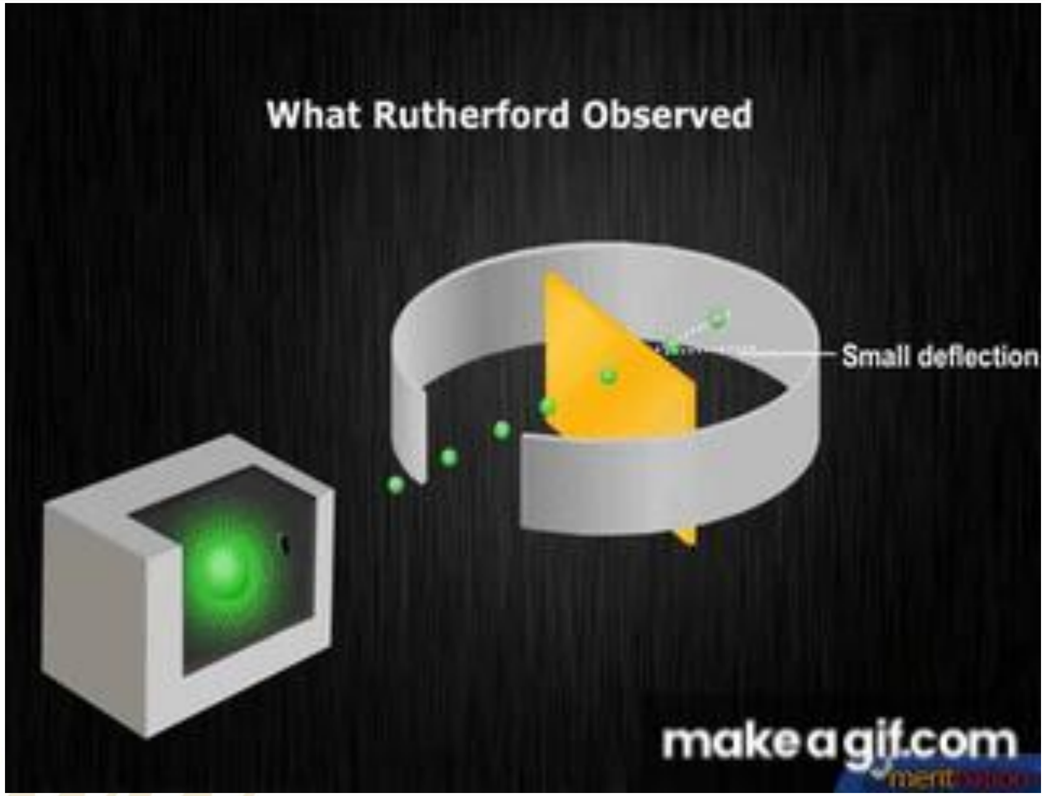
This sphere has almost all the mass of the atom, while electrons are embedded in it.



Models of the atom

Rutherford model of the atom

Rutherford directed a stream of α-particles (^4_2He) towards a very thin gold foil. He found that:



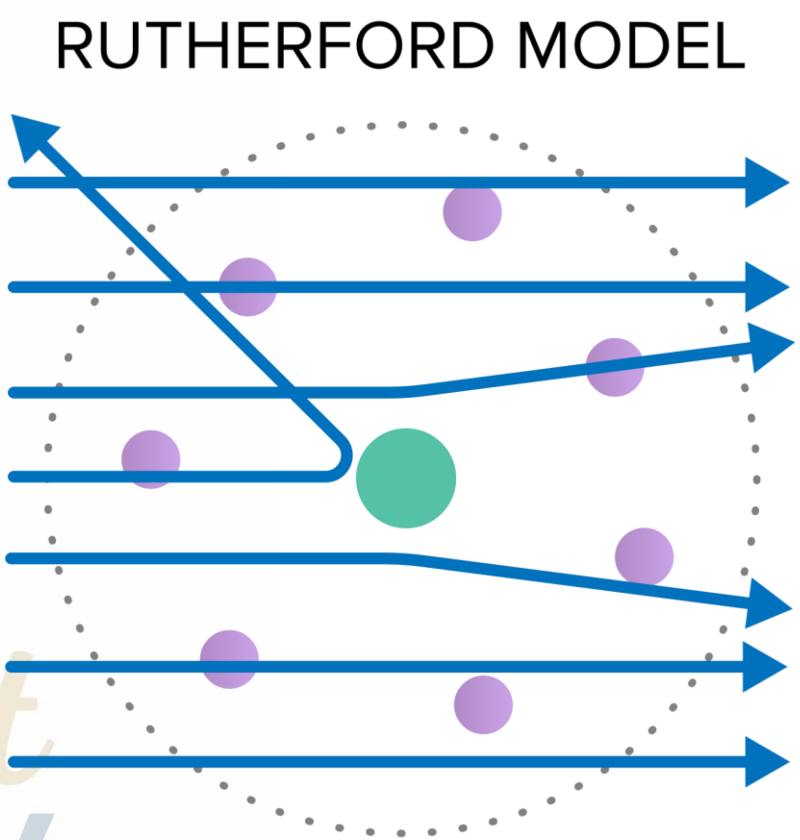
- Most of the α-particles cross the foil without deviation.
- Some of them pass with deviation
- Few of them rebound back.

Models of the atom

He proposed that:

The mass of the atom is concentrated in a small space at the center of the atom (nucleus).

The electrons revolve around it



Models of the atom

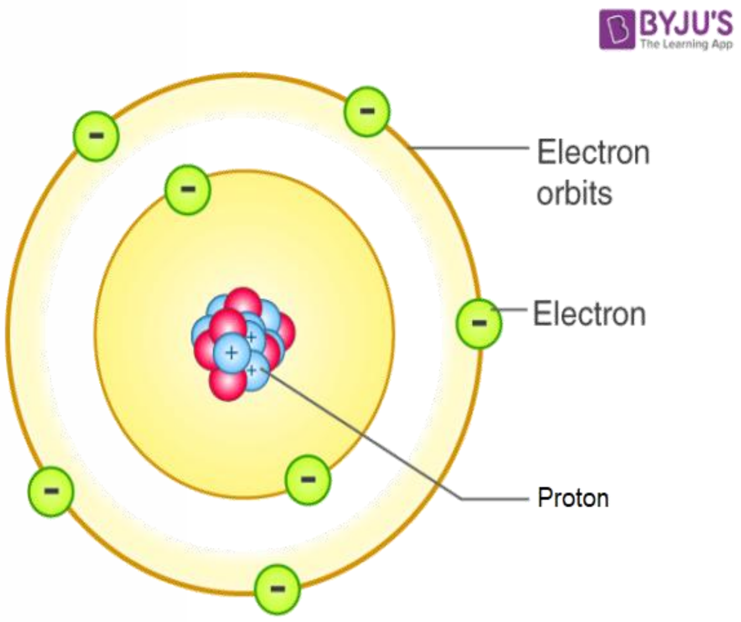
Bohr's model of the atom:

According to Bohr's model, we can deduce:

- The electron moves in circular orbits around the nucleus.

- Electrons can exist only in discrete energy levels (quantized energy).

- Each energy level has a quantum number "n" where it can be n=1, 2, 3…∞



Bohr atomic model of a Nitrogen atom

Models of the atom

Bohr's model of the atom:

1.Ground or fundamental state (E_1):

The first energy level has the smallest energy and is the closest level to the nucleus.

$n = 1$ Ground state E_1

The electrons in this level are in stable state.

Its quantum number is n=1

Models of the atom/ Bohr's model of the atom

2. Excited states:

The electrons are unstable.

The second energy level of energy (E_2) is called 1st excited state, and its quantum number is n=2.

The third energy level of energy (E_3) is called 2nd excited state, and its quantum number is n=3.

The fourth energy level of energy (E_4) is called 3th excited state, and its quantum number is n=4.

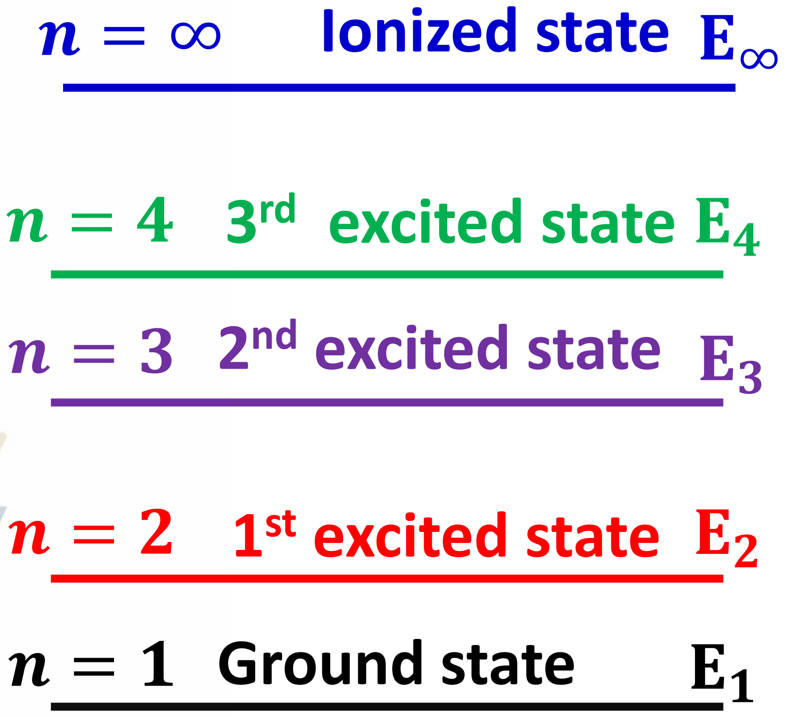
Models of the atom

3. Ionization states (E_∞):

The energy level for which the electron is removed or extracted from the atom. The energy of this level is $E_\infty = 0$

The quantum number of this level is $n = \infty$.

It is impossible for the atom to have energy between these energy levels



The energy of the atom is quantized

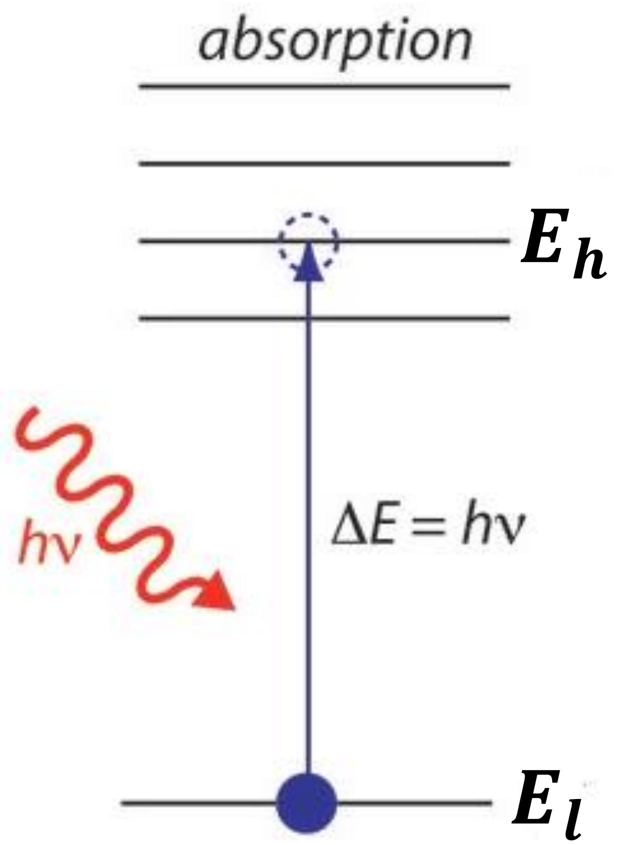
Emission and absorption of a photon

Absorption of photon(excitation):

When an atom in any energy level absorbs a photon, then the atom moves to a higher energy level.

But this occurs only if:

$$E_{ph} = \frac{hc}{\lambda} = E_h - E_l$$



The atom does not remain in its excited state more than 10^{-8} s, it returns back to the ground state.

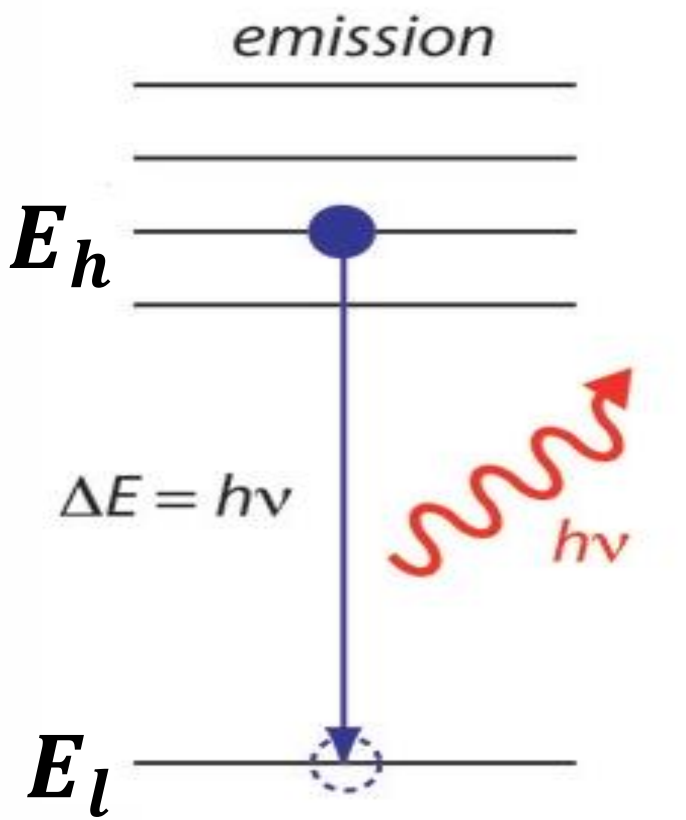
Emission and absorption of a photon

Emission of photon(de-excitation):

When an atom makes a transition from higher energy level (E_{h}) to a lower energy level (E_l);

The **atom emits a photon** of energy:

$$E_{ph} = \frac{hc}{\lambda} = E_{\mathrm{h}} - E_l$$



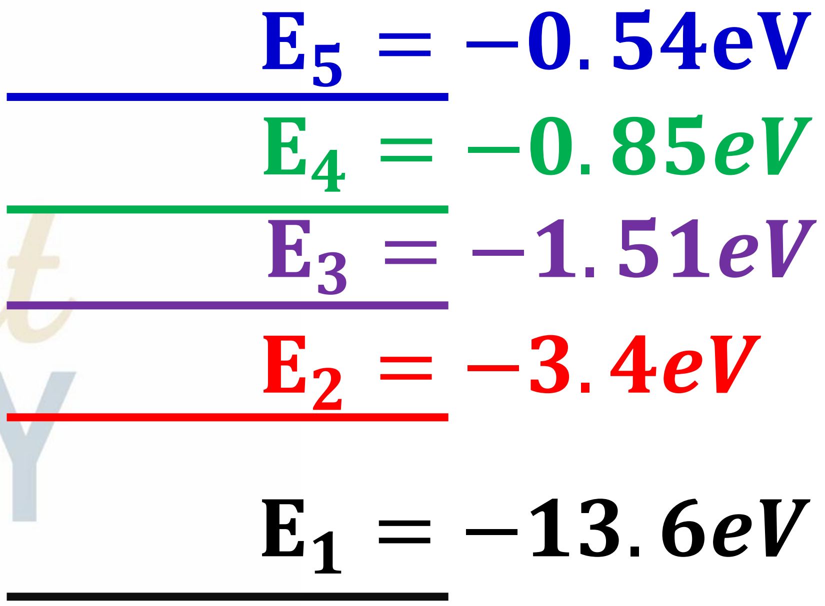
If the wavelength of the emitted photon is $400nm \leq \lambda \leq 800nm$ *; this photon has light that we can see.*

Emission and absorption of a photon

Application 1:

Given the section of energy level diagram of hydrogen atom.

Given: $h = 6.62 \times 10^{-34} J.s$; $1eV = 1.6 \times 10^{-19} J$; $c = 3 \times 10^{8} m/s$; $400nm \leq \lambda_{visible} \leq 800nm$



1)Pick out the energy of the 3rd excited state of the hydrogen atom.

2)Determine the wavelength of the emitted photon when the atom makes a downward transition: E_5 to E_2

Emission and absorption of a photon

3) This photon is visible why?

4)The hydrogen atom is in the ground state.

a)The atom is hit by a photon of energy 11eV. Specify whether this photon is absorbed.

b)Deduce the state of the atom.

$E_5 = -0.54eV$

$E_4 = -0.85eV$

$E_3 = -1.51eV$

$E_2 = -3.4eV$

$E_1 = -13.6eV$

Emission and absorption of a photon

$h = 6.62 \times 10^{-34} J.s$; $1eV = 1.6 \times 10^{-19} J$; $c = 3 \times 10^{8} m/s$;

$400nm \leq \lambda_{visible} \leq 800nm$

1)Pick out the energy of the 3rd excited state of the hydrogen atom.

The energy of the 3rd excited state is

$$E_4 = -0.85eV.$$

$E_5 = -0.54eV$

$E_4 = -0.85eV$

$E_3 = -1.51eV$

$E_2 = -3.4eV$

$E_1 = -13.6eV$

Emission and absorption of a photon

$h = 6.62 \times 10^{-34} J.s$; $1eV = 1.6 \times 10^{-19} J$; $c = 3 \times 10^{8} m/s$;

2) Determine the wavelength of the emitted photon when the atom makes a transition from E_5 to E_2.

$$E_5 = -0.54eV$$
$$E_4 = -0.85eV$$
$$E_3 = -1.51eV$$
$$E_2 = -3.4eV$$
$$E_1 = -13.6eV$$

$$E_{ph} = E_h - E_l = E_5 - E_2$$

$$E_{ph} = -0.54 - (-3.4) = 2.86eV$$

$$E_{ph} = \frac{hc}{\lambda} \Rightarrow \lambda = \frac{6.62 \times 10^{-34} \times 3 \times 10^{8}}{2.86 \times 1.6 \times 10^{-19}}$$

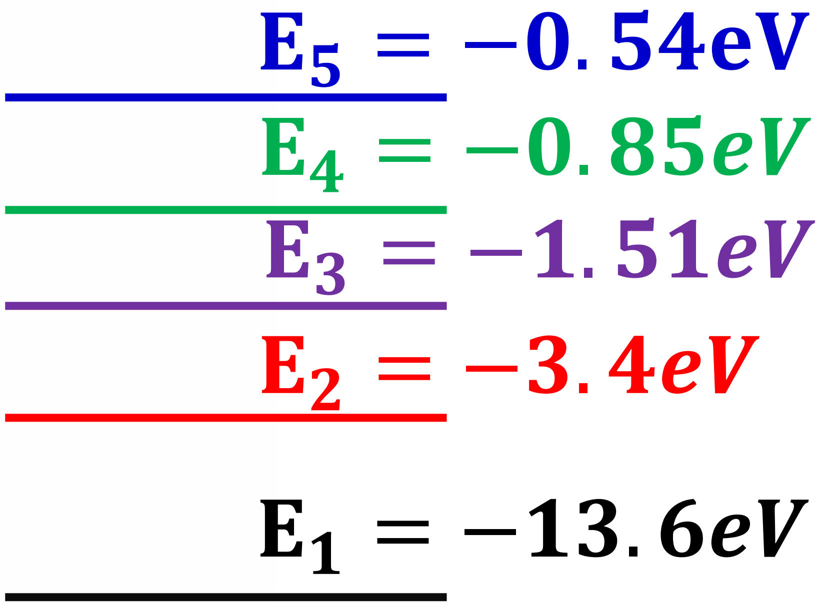
$$\lambda_{5\to2} = 4.34 \times 10^{-7} m = 434nm$$

Emission and absorption of a photon

3) This photon is visible why?

The photon is visible because its wavelength is in the visible region

$$400nm \leq \lambda_{ph} = 434nm \leq 800nm$$



4)The hydrogen atom is in the ground state.

a)The atom is hit by a photon of energy 11eV. Specify whether this photon is absorbed.

$$E_{ph} = E_h - E_l \Rightarrow 11 = E_h - (-13.6) \Rightarrow E_h = -2.6eV$$

Since $E_h = -2.6eV$ does not equal any energy of the given levels then the photon is not absorbed.

Emission and absorption of a photon

b)Deduce the state of the atom.

Since the photon is not absorbed then the atom remains in the ground state.

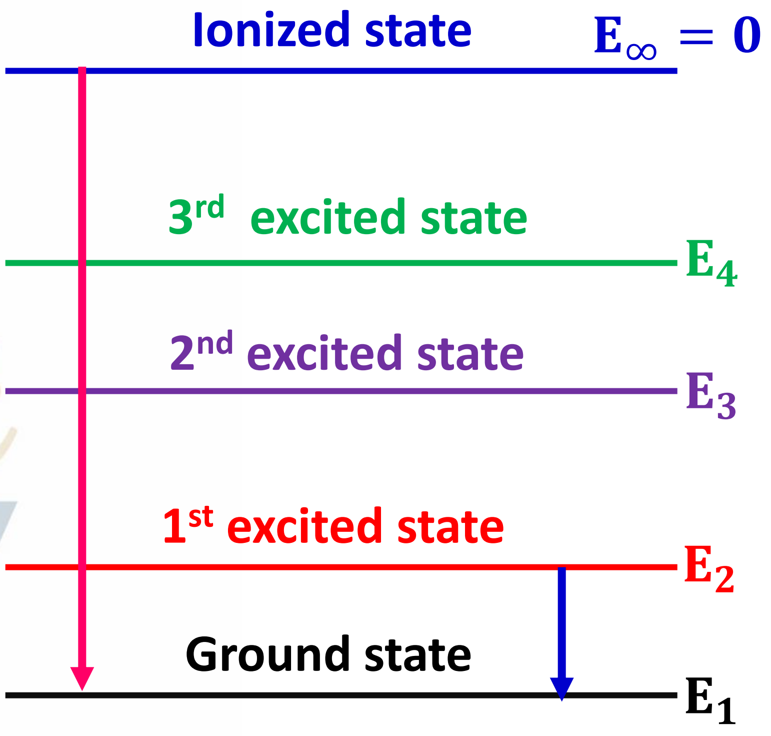
Emission and absorption of a photon

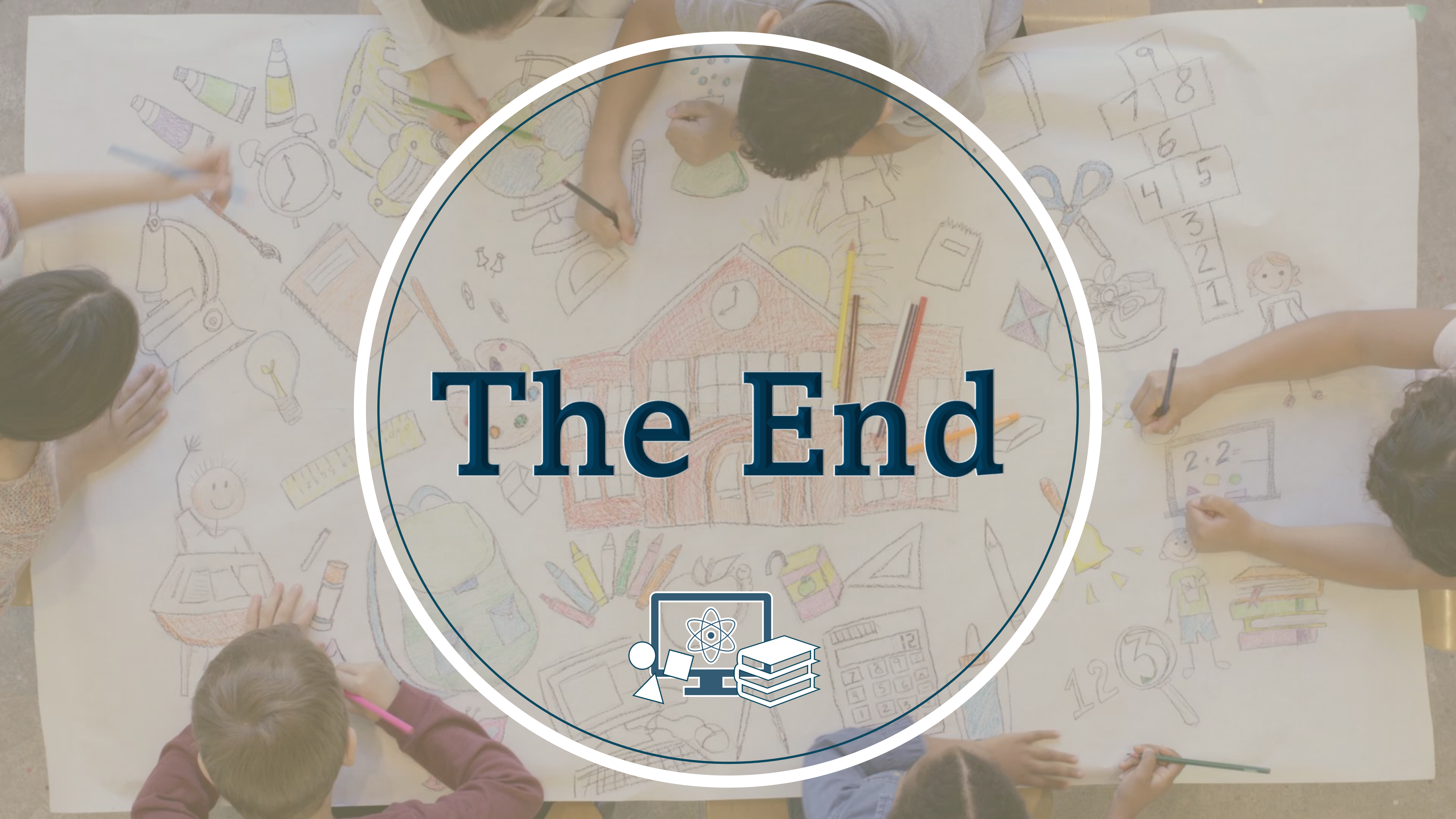
1)The emitted photon is with **maximum frequency (minimum λ)** if the electron transmitted from n=∞ to n=1 then:

$$h\nu_{max} = E_{\infty} - E_1$$

2)The emitted photon is with **minimum frequency (maximum λ)** if the electron transmitted from n=2 to n=1

$$h\nu_{min} = E_2 - E_1$$

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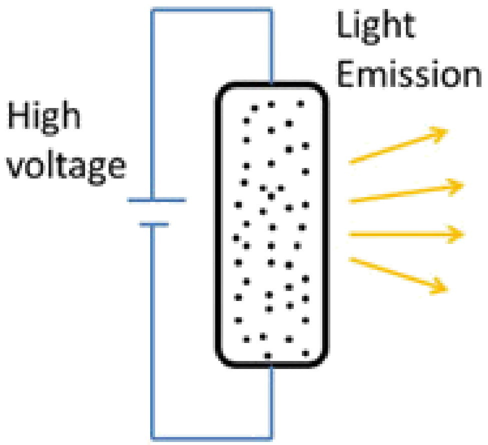
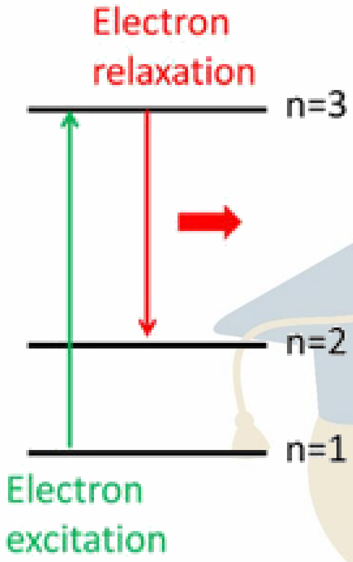
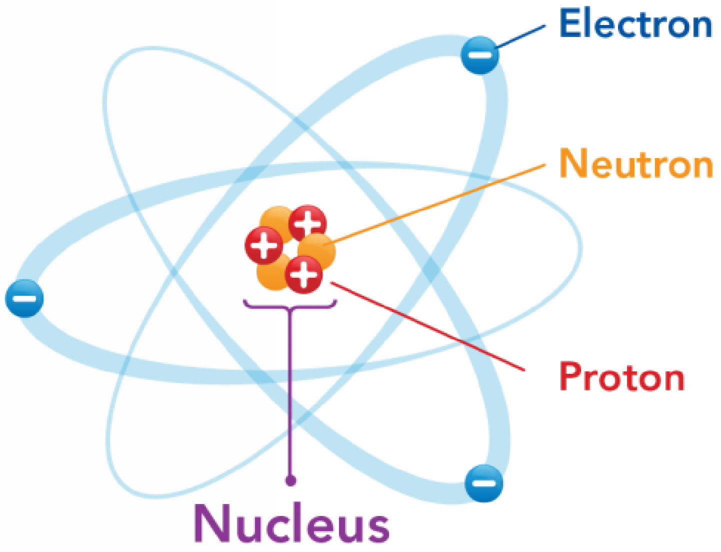






Chapter 17

The atom

Prepared & Presented by: Mr. Mohamad Seif

OBJECTIVES

4 **The spectral series of the hydrogen atom**

5 **Expression of energy levels of the hydrogen atom**

The spectral series of the hydrogen atom

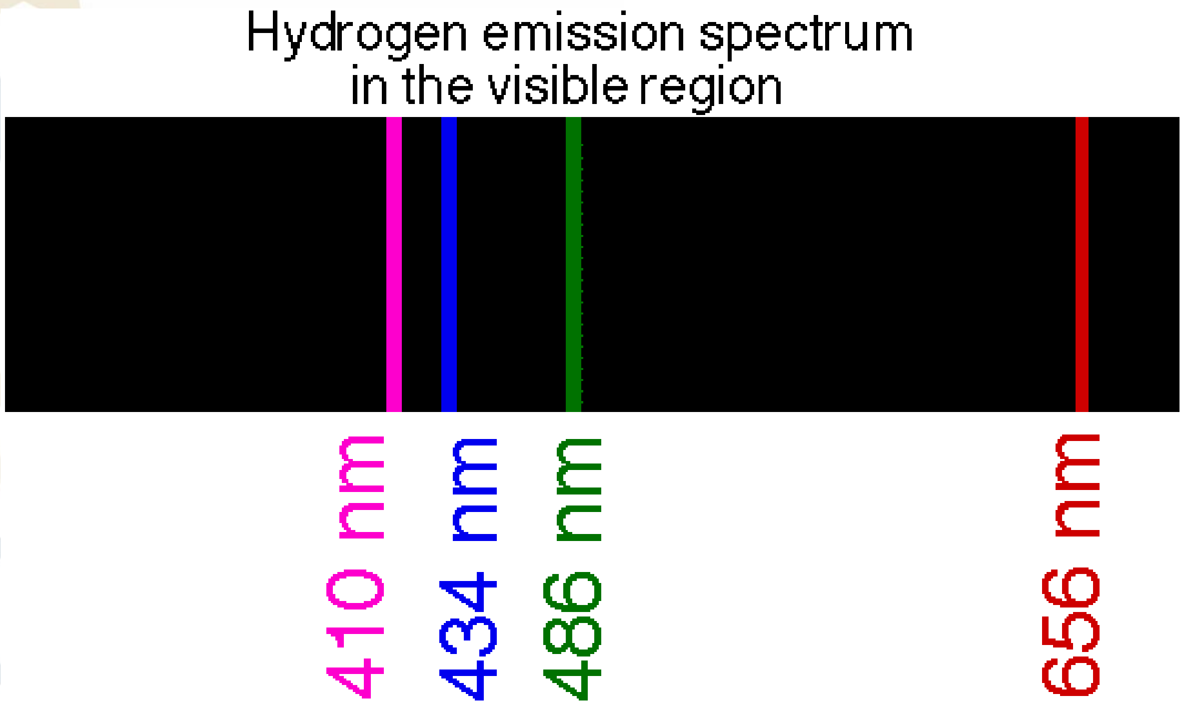
The emission spectrum of hydrogen gas shows a series of four spectral lines (4 colors) whose wavelengths in air are:

$H_\alpha: \lambda_\alpha = 656.5nm.$

$H_\beta: \lambda_\beta = 486.1nm.$

$H_\gamma: \lambda_\gamma = 434.1nm.$

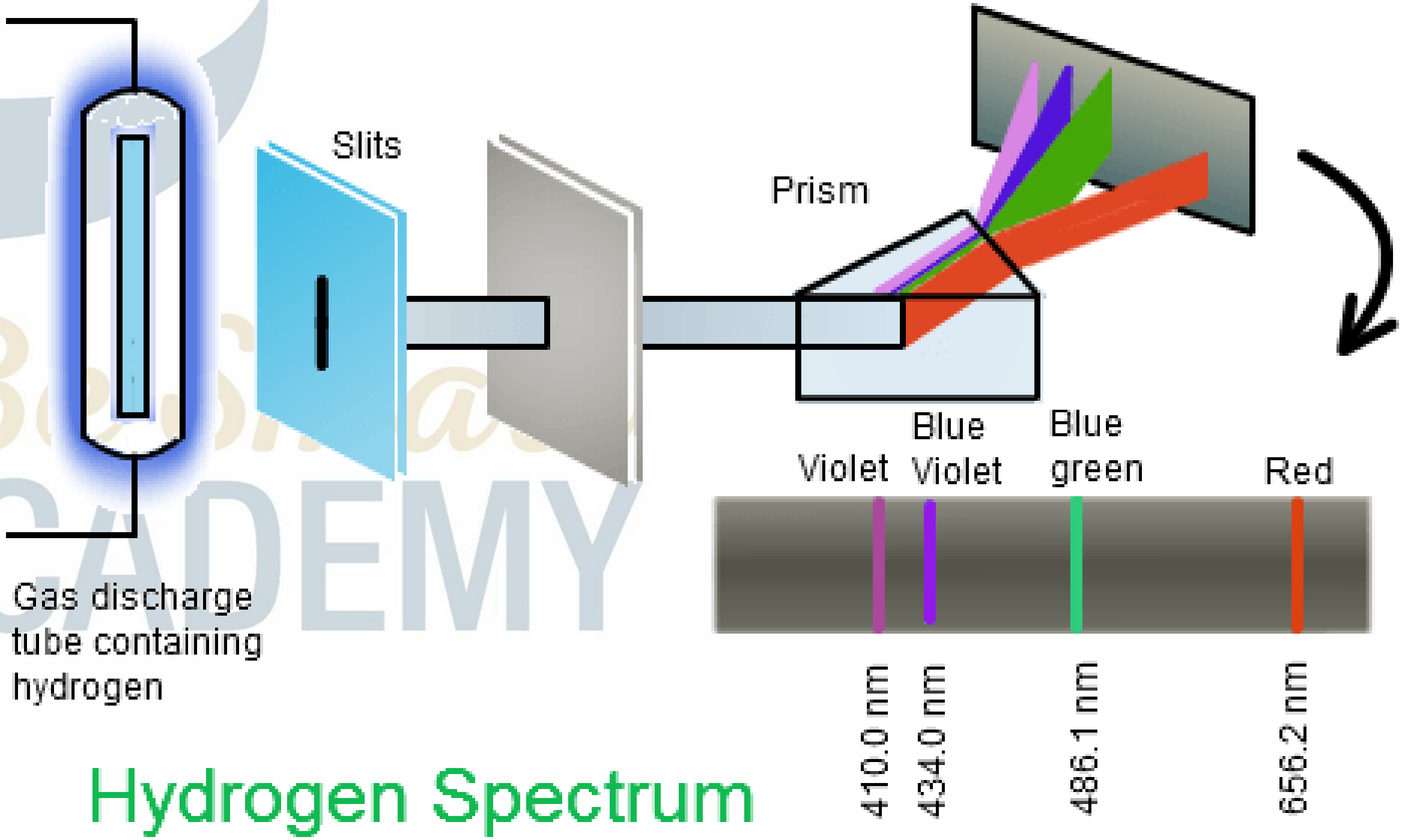
$H_\delta: \lambda_\delta = 410.2nm.$



The spectral series of the hydrogen atom

The equation that covers all transitions of hydrogen atom in which the atom moves from a high energy level n to a lower energy level m is given by:

$$\frac{1}{\lambda} = R\left(\frac{1}{m^2} - \frac{1}{n^2}\right)$$



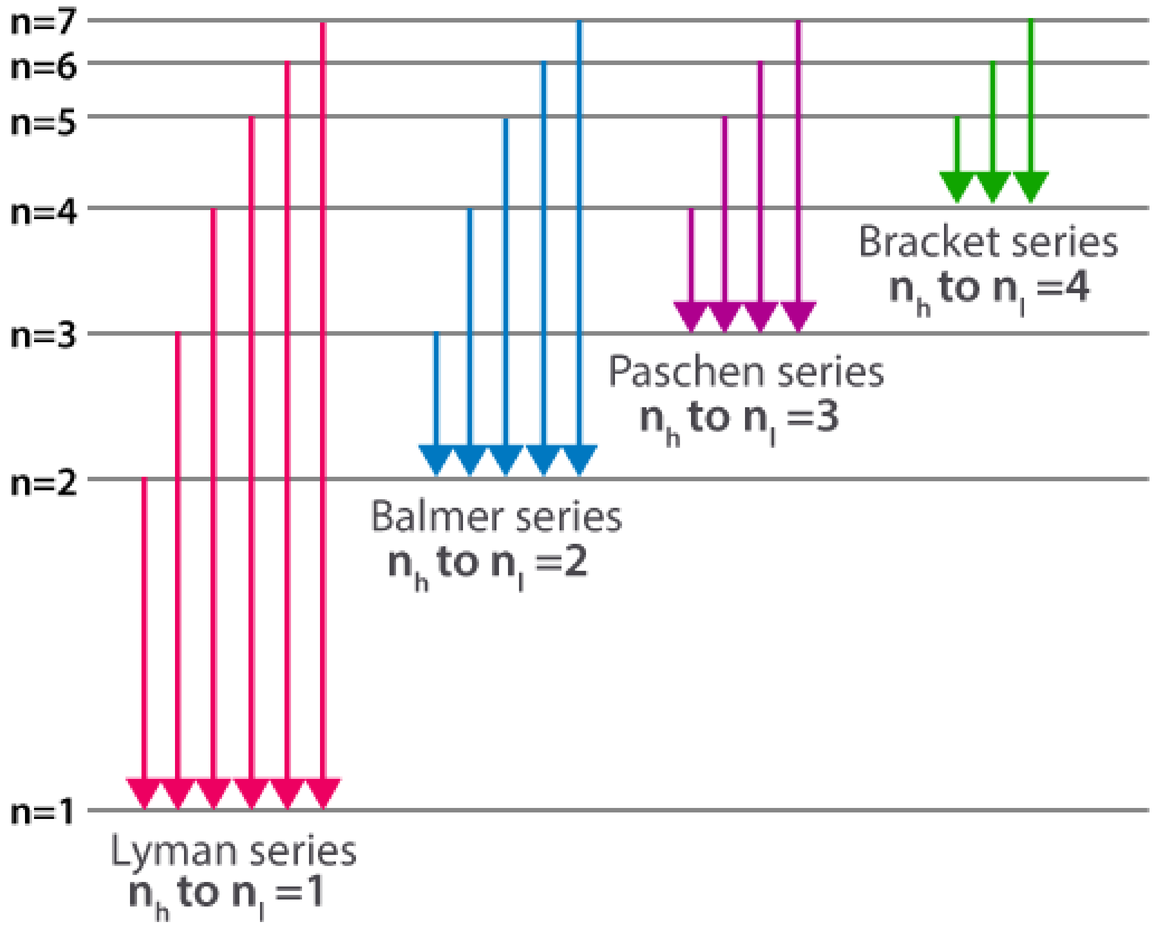
The spectral series of the hydrogen atom

1)Lyman series:

Lyman series corresponds to downward transition from an energy level of $n \geq 2$ to energy level of $n = 1$.

$$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{n^2}\right)$$

Where $n = 2, 3, 4 \dots$



All the Wavelengths in Lyman series lie in the Ultraviolet range

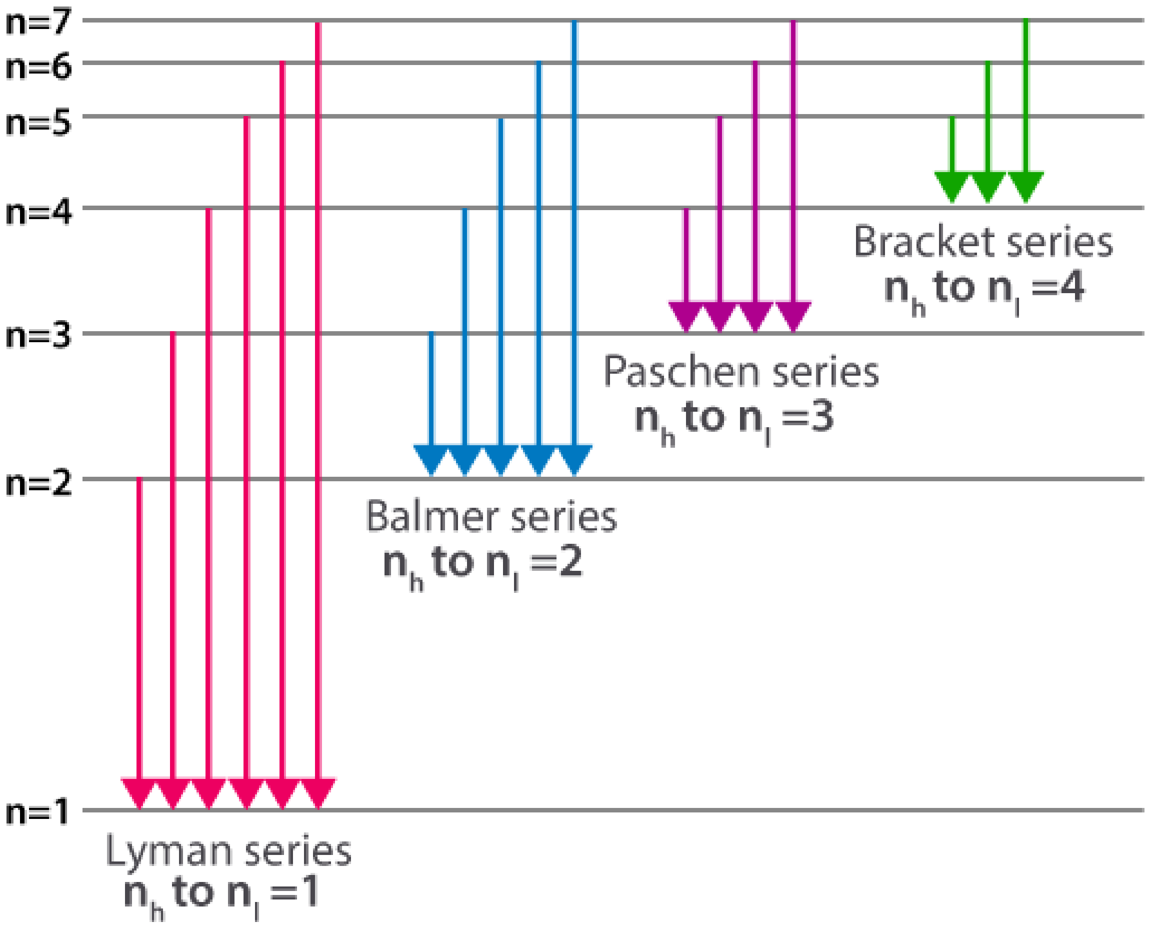
The spectral series of the hydrogen atom

2)Balmer series:

Balmer series corresponds to downward transition from an energy level of $n \geq 3$ to energy level of $n = 2$.

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

Where $n = 3, 4, 5 \ldots$



All the Wavelengths in Balmer series lie in the visible range

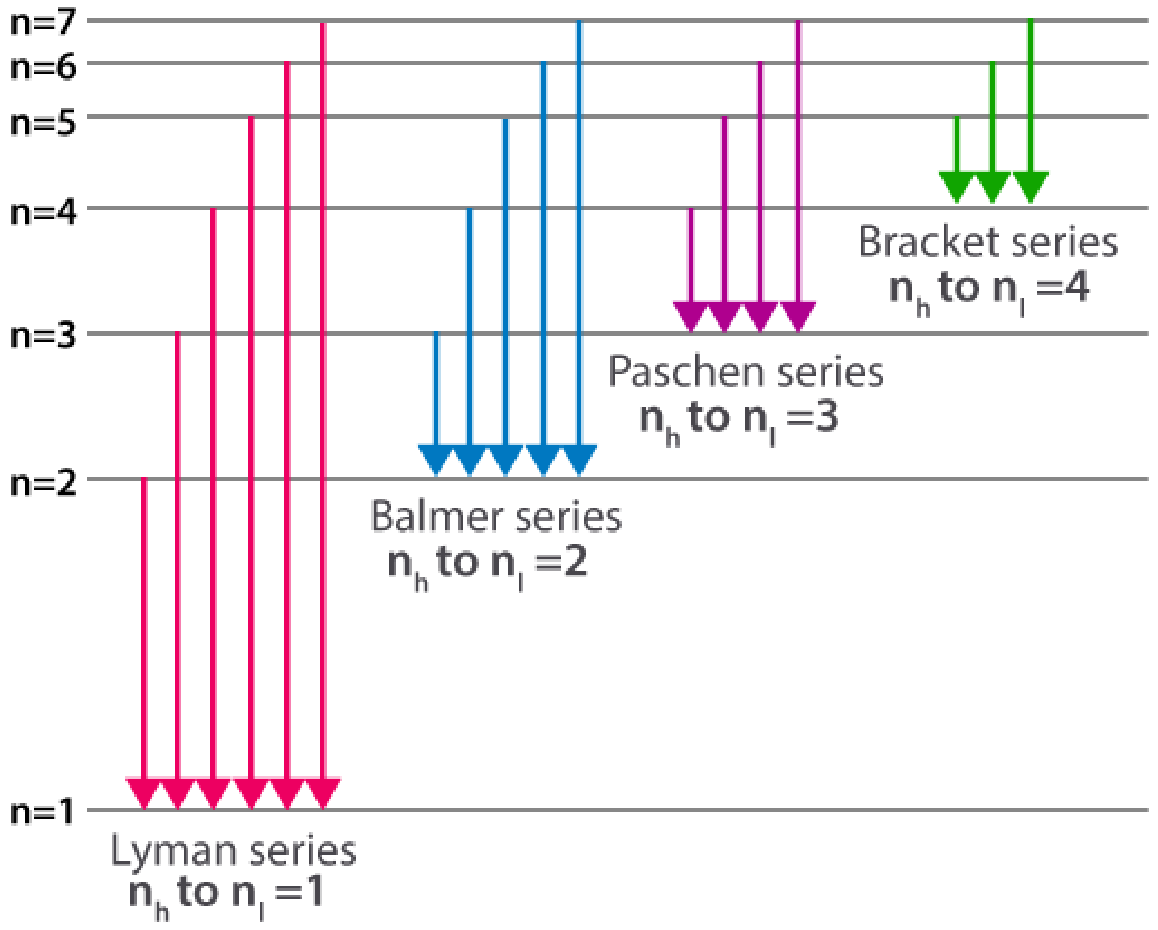
The spectral series of the hydrogen atom

3)Paschen series:

Paschen series corresponds to downward transition from an energy level of $n \geq 4$ to energy level of $n = 3$.

$$\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{n^2}\right)$$

Where $n = 4, 5, 6 \ldots$



All the Wavelengths in Paschen series lie in the infrared range

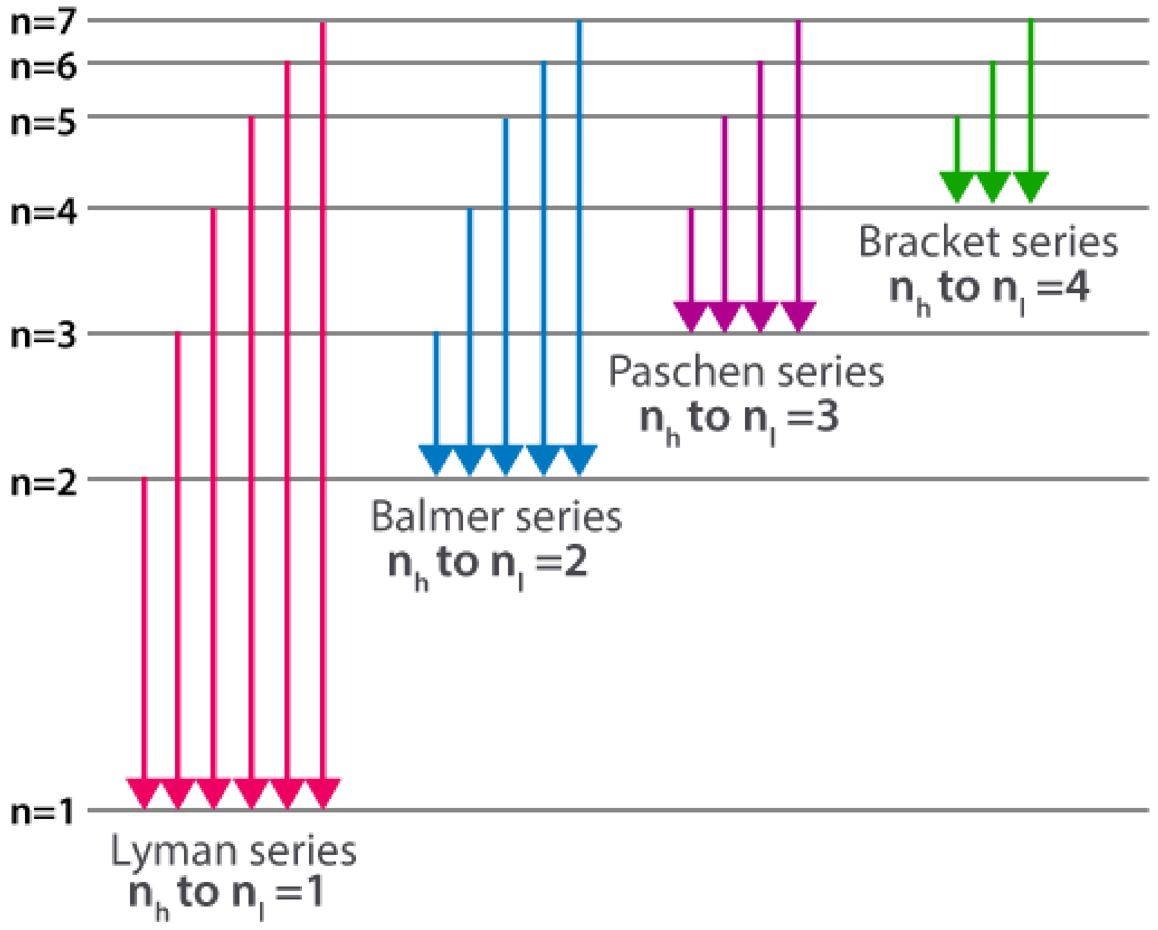
The spectral series of the hydrogen atom

4)Brackett series:

Brackett series corresponds to downward transition from an energy level of $n \geq 5$ to energy level of $n = 4$.

$$\frac{1}{\lambda} = R\left(\frac{1}{4^2} - \frac{1}{n^2}\right)$$

Where $n = 5, 6, 7 \ldots$



All the Wavelengths in Brackett series lie in the infrared range

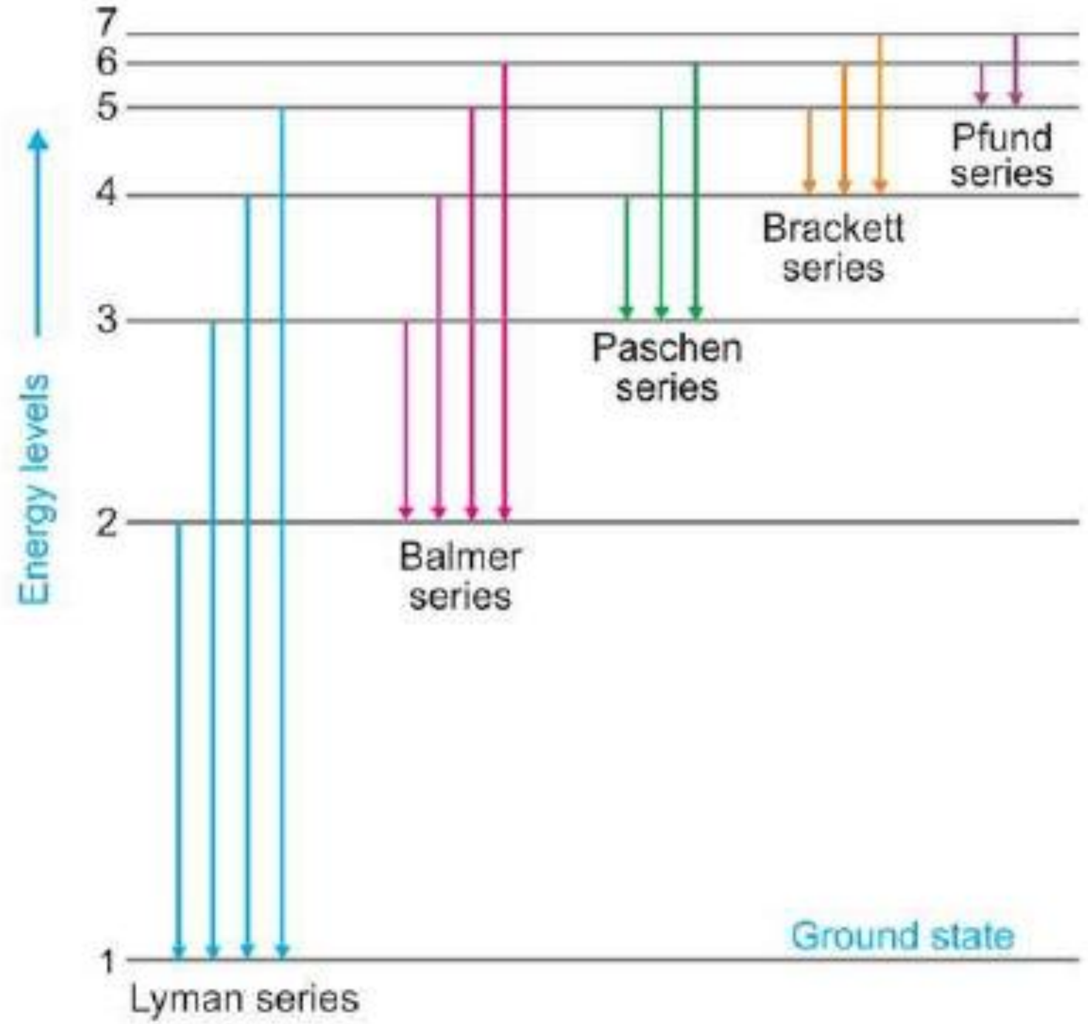
The spectral series of the hydrogen atom

5)Pfund series:

Pfund series corresponds to downward transition from an energy level of $n \geq 6$ to energy level of $n = 5$.

$$\frac{1}{\lambda} = R\left(\frac{1}{5^2} - \frac{1}{n^2}\right)$$

Where $n = 6, 7, 8 \ldots$



All the Wavelengths in Pfound series lie in the infrared range

Expression of energy levels of the hydrogen atom

According to Bohr's model, when the atom makes a transition from a higher level n to a lower level m, the energy of the emitted photon is:

$$E_{ph} = \frac{hc}{\lambda} = E_n - E_m$$

Substituting Rydberg formula, $\frac{1}{\lambda} = R\left(\frac{1}{m^2} - \frac{1}{n^2}\right)$ **in the above equation:**

$$E_{ph} = hcR\left(\frac{1}{m^2} - \frac{1}{n^2}\right) = E_n - E_m$$

Expression of energy levels of the hydrogen atom

$$E_{ph} = hcR\left(\frac{1}{m^2} - \frac{1}{n^2}\right) = E_n - E_m$$

$$E_{ph} = \frac{hcR}{m^2} - \frac{hcR}{n^2} = E_n - E_m$$

$$-\frac{hcR}{n^2} + \frac{hcR}{m^2} = E_n - E_m$$

$$E_n = -\frac{hcR}{n^2}$$

And

$$E_m = -\frac{hcR}{m^2}$$

Expression of energy levels of the hydrogen atom

The energy levels of the hydrogen atom are negative and are given by: $E_n = -\frac{hcR}{m^2}$

$$E_n = -\frac{6.62 \times 10^{-34} \times 3 \times 10^{8} \times 1.097 \times 10^{7}}{n^2} \Rightarrow E_n = -\frac{2.179 \times 10^{-18}}{n^2}$$

$$1eV = 1.6 \times 10^{-19} J \Rightarrow E_n = -\frac{2.179 \times 10^{-18}}{1.6 \times 10^{-19} \times n^2}$$

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

Expression of energy levels of the hydrogen atom

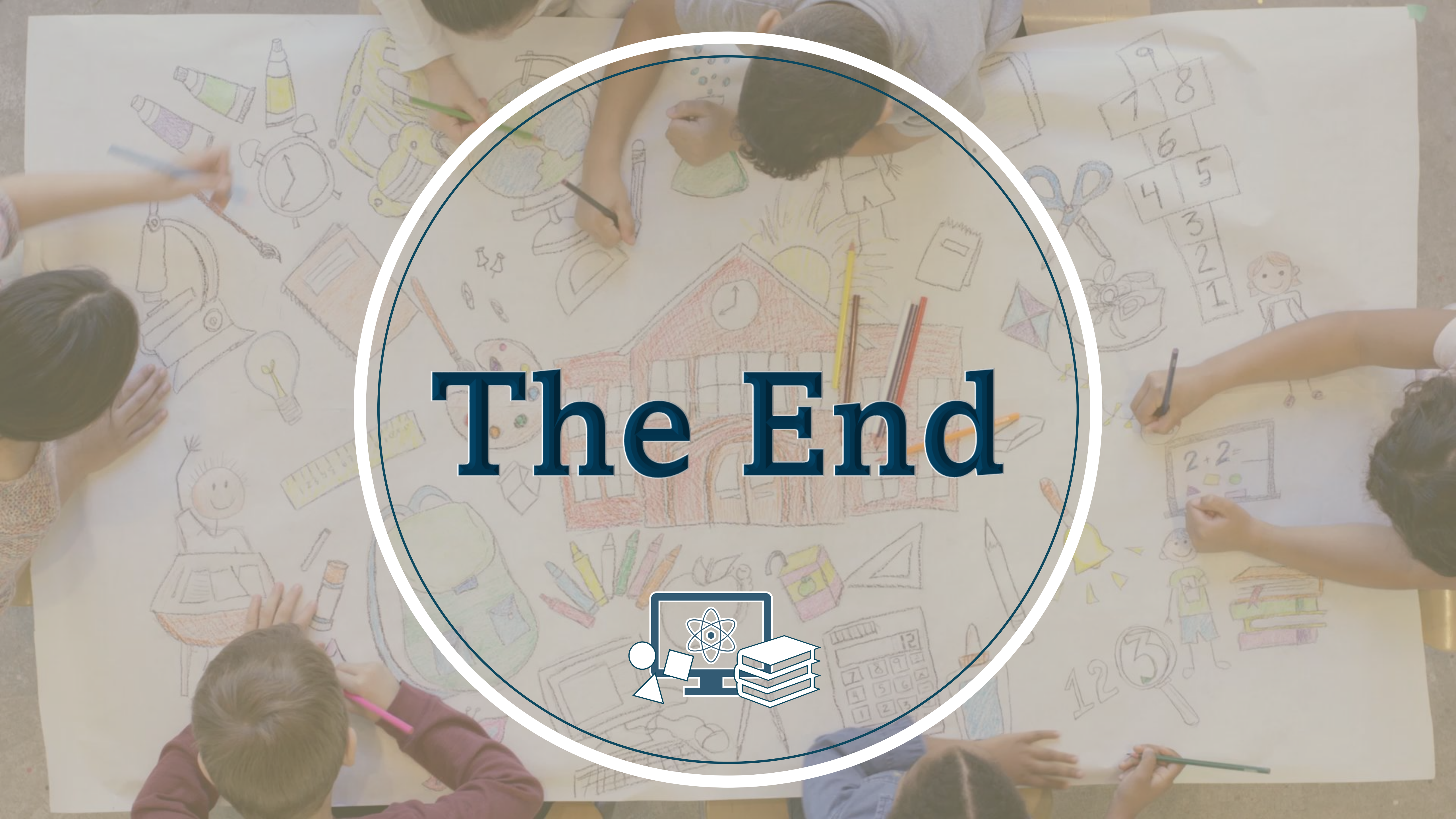
$$E_n = -\frac{13.6}{n^2}$$

For n=1: $E_1 = -\frac{13.6}{(1)^2} = -13.6eV$ — Ground state

For n=2: $E_2 = -\frac{13.6}{(2)^2} = -3.4eV$ — 1st excited state

For n=3: $E_3 = -\frac{13.6}{(3)^2} = -1.51eV$ — 2nd excited state

For n=∞: $E_\infty = -\frac{13.6}{\infty} = 0eV$ — Reference (ionized) state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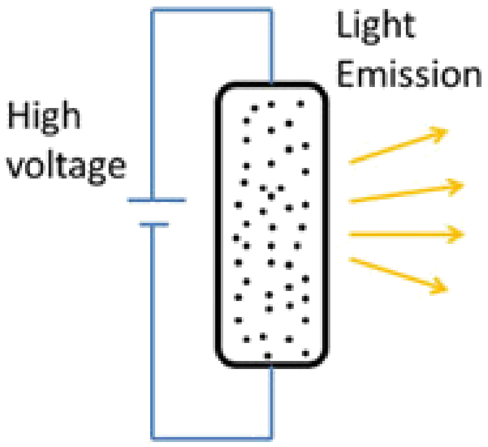
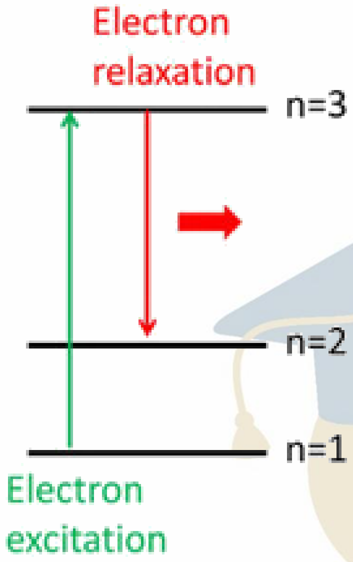
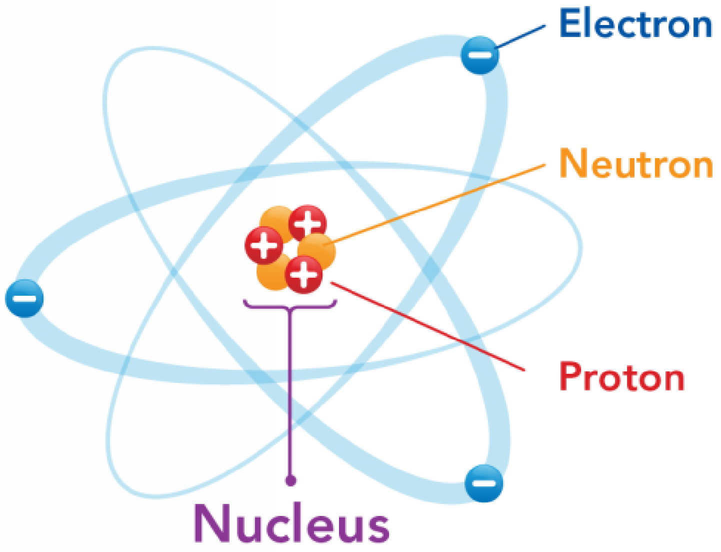






Chapter 17

The atom

Prepared & Presented by: Mr. Mohamad Seif

OBJECTIVES

1. Ionization energy
2. K.E of extracted electron
3. Emission and absorption spectrum

Ionization energy

Ionization energy (W_{ion}):

Is the minimum energy needed to liberate an electron from the ground state of the atom.

$$W_{ion} = E_{\infty} - E_1$$

For hydrogen atom: $\mathbf{E_{\infty} = 0eV}$ **and** $\mathbf{E_1 = -13.6eV}$

$$W_{ion} = E_{\infty} - E_1 \Rightarrow W_{ion} = 0 - (-13.6)$$

$$W_{ion} = 13.6eV$$

K.E of extracted electron

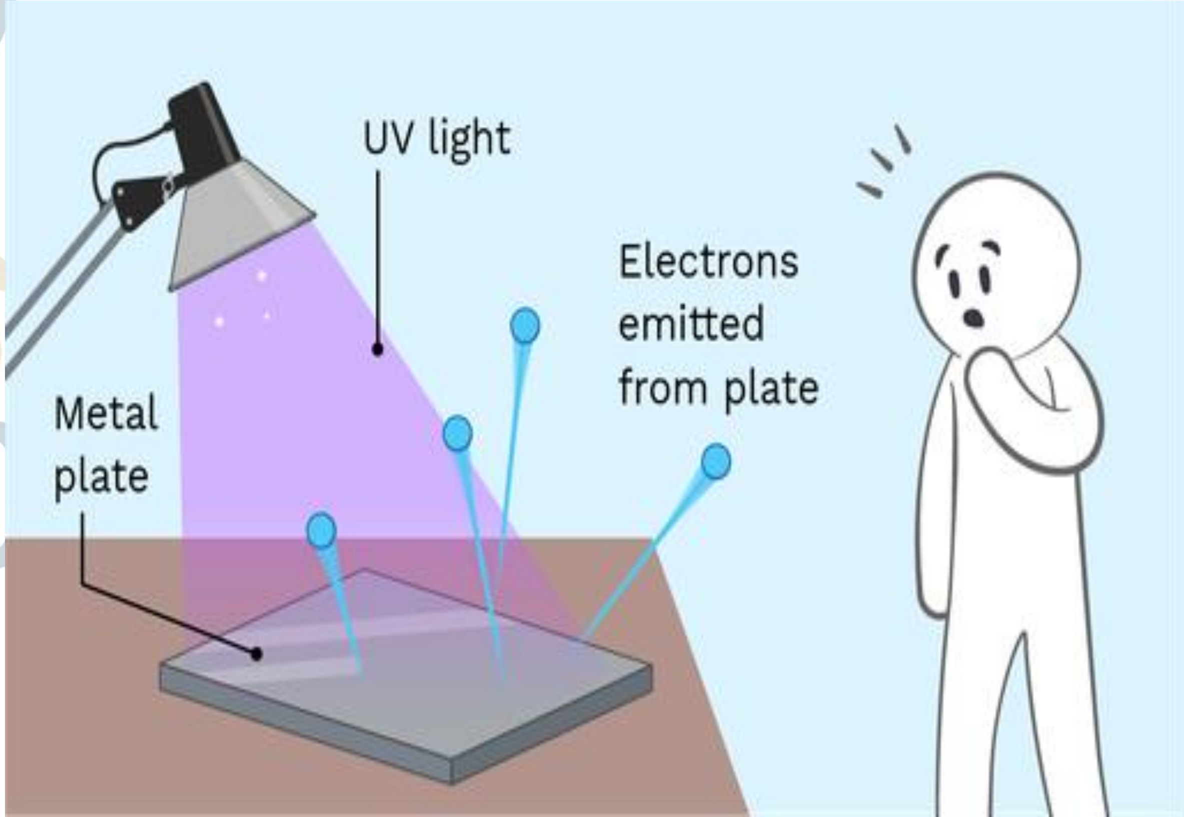
Kinetic energy of extracted electron:

If the energy of the photon is greater than W_{ion}: $E_{ph} > W_{ion}$:

The electron is extracted with kinetic energy determined by:

$$K.E = E_{ph} - W_{ion}$$

This formula was mentioned in photoelectric effect lesson



Application 2:

The energy of the hydrogen atom is given by $E_n = -\frac{13.6}{n^2}$, in eV

1)Calculate in eV the energies of the first two energy levels.

2)The hydrogen atom is de-excited from the first excited state to ground state. Calculate the wavelength.

3)To which series does this transition belong?

4)The hydrogen atom is hit by a photon of energy of 14eV. Calculate the K. E of extracted electron.

Given: $R = 1.097 \times 10^7 m^{-1}$; $\mathrm{h} = 6.62 \times 10^{-34} J.s$; $c = 3 \times 10^8 m/s$; $1\ eV = 1.6 \times 10^{-19} J$;

$R = 1.097 \times 10^{7} m^{-1}$; $h = 6.62 \times 10^{-34} J.s$; $c = 3 \times 10^{8} m/s$; $1\ eV = 1.6 \times 10^{-19} J$;

1)Calculate in eV the energies of the first two energy levels.

For n=1: $$E_1 = -\frac{13.6}{(1)^2} = -13.6\text{eV}$$

For n=2: $$E_2 = -\frac{13.6}{(2)^2} = -3.4\text{eV}$$

$R = 1.097 \times 10^{7} m^{-1}$; $h = 6.62 \times 10^{-34} J.s$; $c = 3 \times 10^{8} m/s$; $1\ eV = 1.6 \times 10^{-19} J$;

2)The hydrogen atom is de-excited from the first excited state to ground state. Calculate the wavelength.

$$\frac{1}{\lambda} = R\left(\frac{1}{m^2} - \frac{1}{n^2}\right) \Rightarrow \frac{1}{\lambda} = 1.097 \times 10^{7}\left(\frac{1}{(1)^2} - \frac{1}{(2)^2}\right)$$

$$\frac{1}{\lambda} = 1.097 \times 10^{7} \times 0.75 \Rightarrow \lambda = 1.215 \times 10^{-7} m$$

$R = 1.097 \times 10^{7} m^{-1}$; $h = 6.62 \times 10^{-34} J.s$; $c = 3 \times 10^{8} m/s$; $1\ eV = 1.6 \times 10^{-19} J$;

3)To which series does this transition belong?

The transition to ground state refers to Lyman series

4)The hydrogen atom is hit by a photon of energy of 14eV. Calculate the K. E of extracted electron.

$W_{ion} = E_{\infty} - E_{1} \Rightarrow W_{ion} = 0 - (-13.6) \Rightarrow W_{ion} = 13.6 eV$

$K.E = E_{ph} - W_{ion} \Rightarrow K.E = 14 - 13.6 \Rightarrow K.E = 0.4 eV$

Emission and absorption spectrum

Emission and absorption spectrum characterize the element. Each element has its own emission and absorption spectra

Emission spectrum:

Is the set of wavelengths that constitute the radiation emitted by this element when it is excited. It depends on the source.

Types of emission spectrum:

1)Continuous emission spectrum:

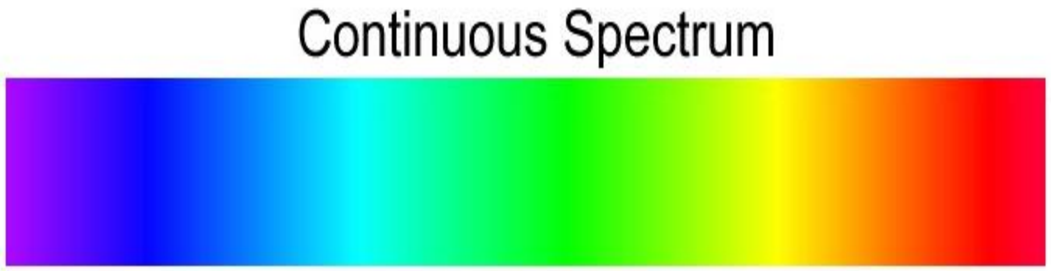
2)Discrete emission spectrum

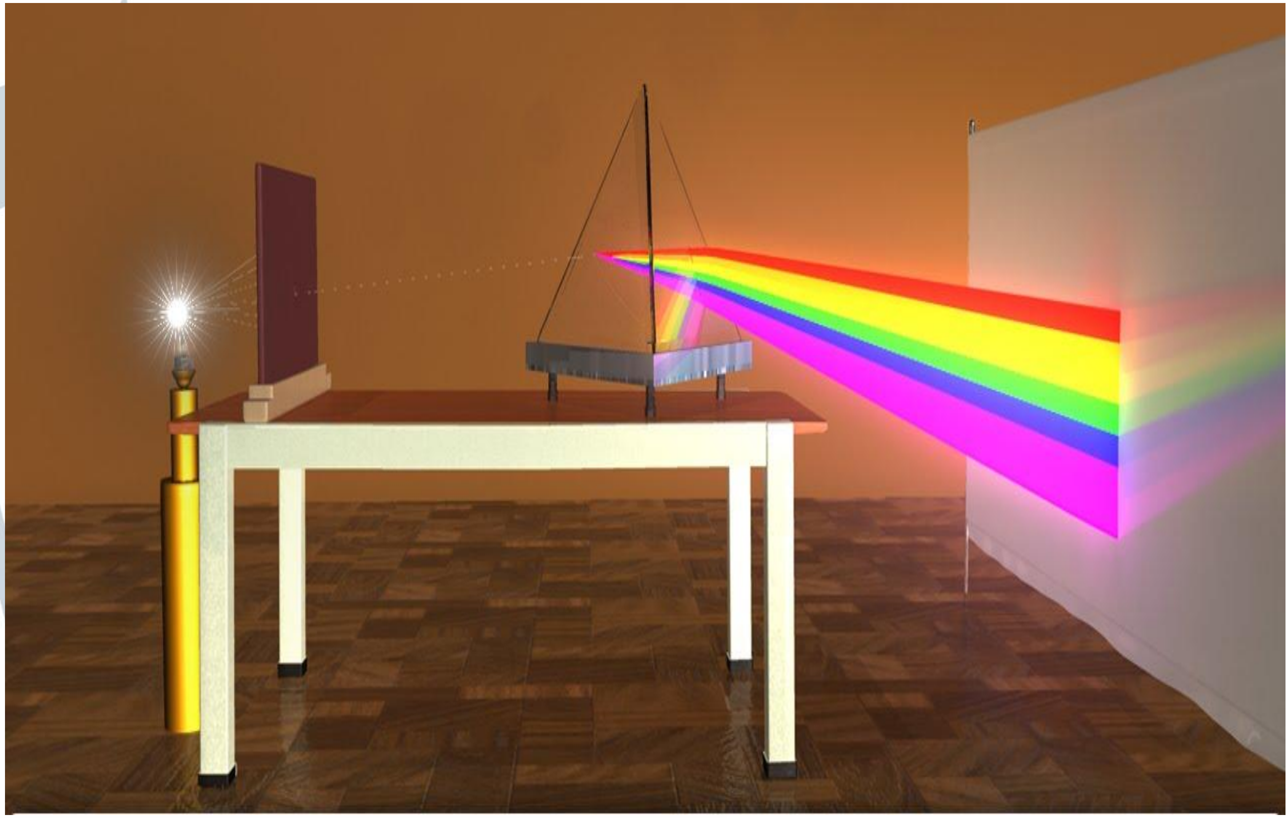
Emission and absorption spectrum

1)Continuous emission spectrum:

Continuous band of colors without dark lines.

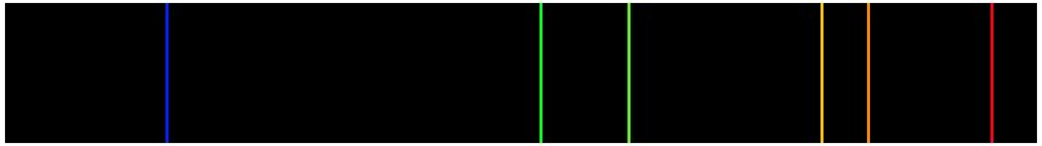
When gases, solids and liquids are heated under high pressure, they have continuous spectra of light.





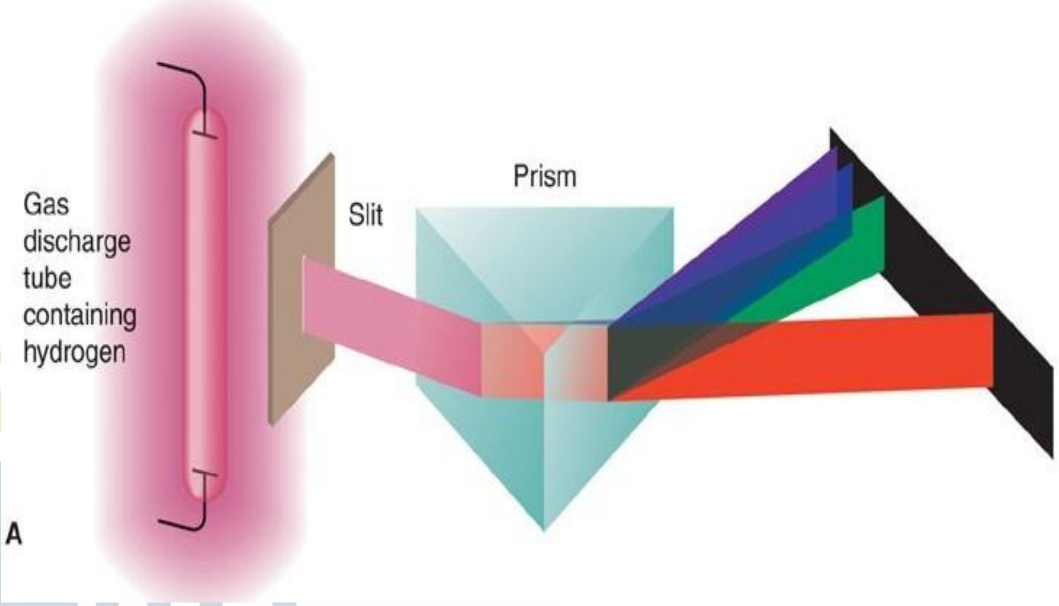
Emission and absorption spectrum

2)Discrete emission spectrum:



Set of discrete bright lines on a dark background

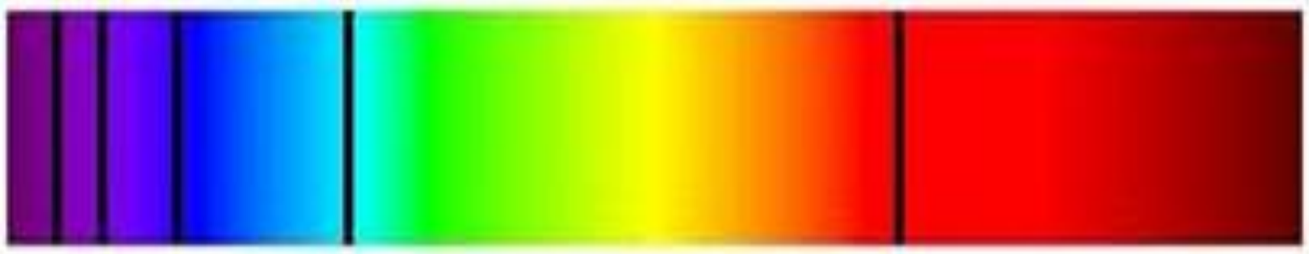
Each bright line in the emission spectrum corresponds to a wavelength of a certain photon emitted by the atom when de-excitation occur.



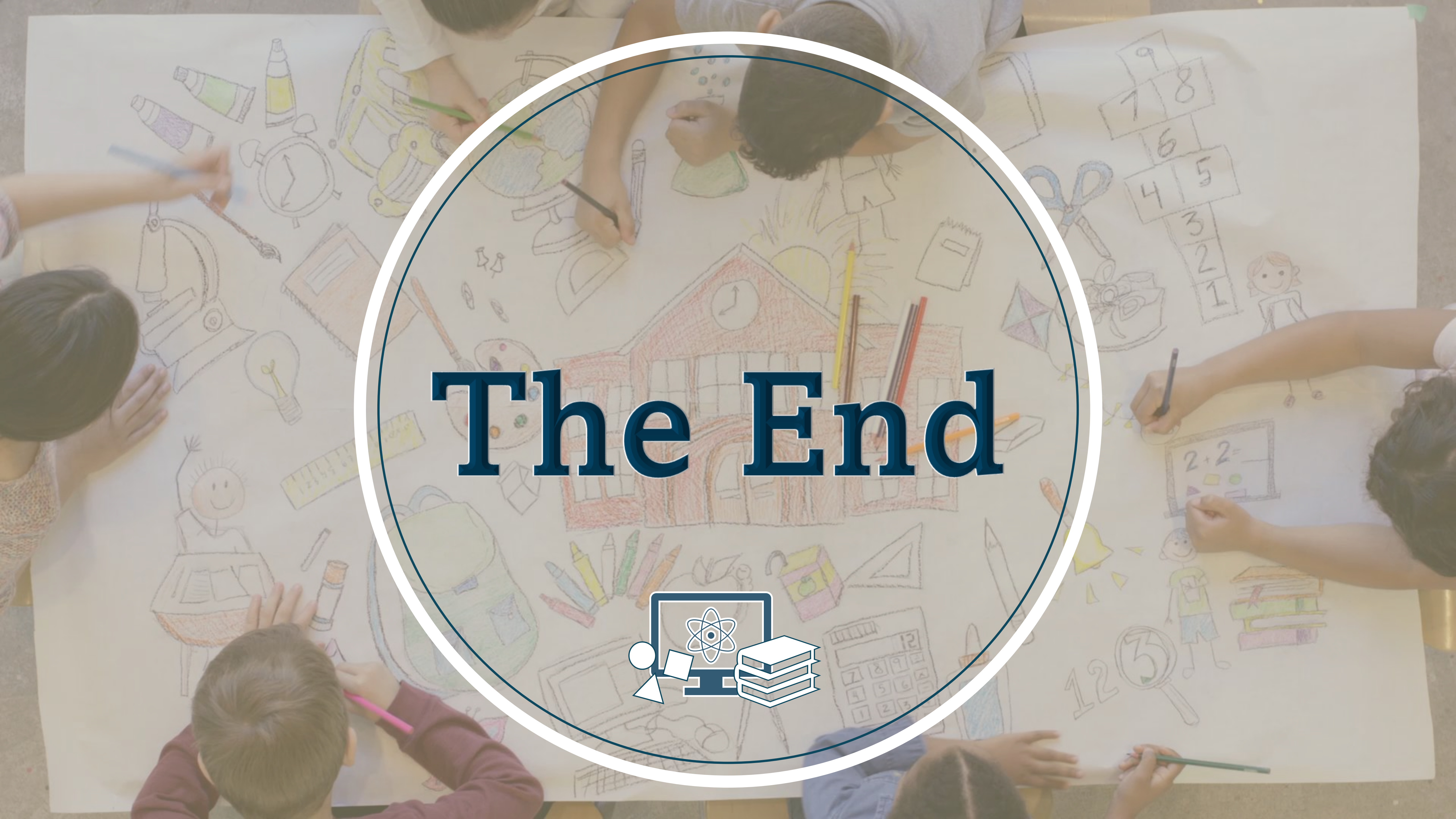
Emission and absorption spectrum

Absorption spectrum:

Set of discrete dark lines on a bright band background.



Each dark line in the absorption spectrum corresponds to a wavelength of a certain photon absorbed by the atom when excitation occur.



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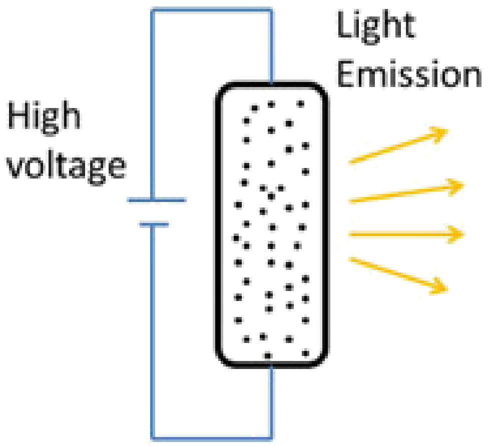
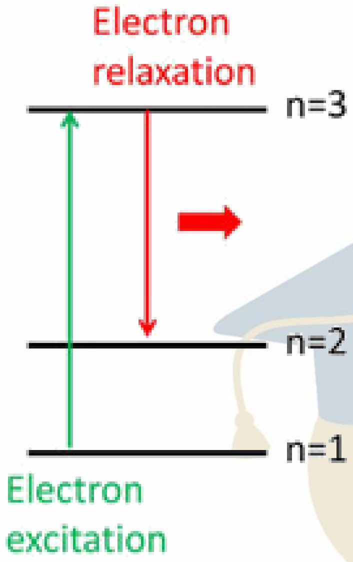
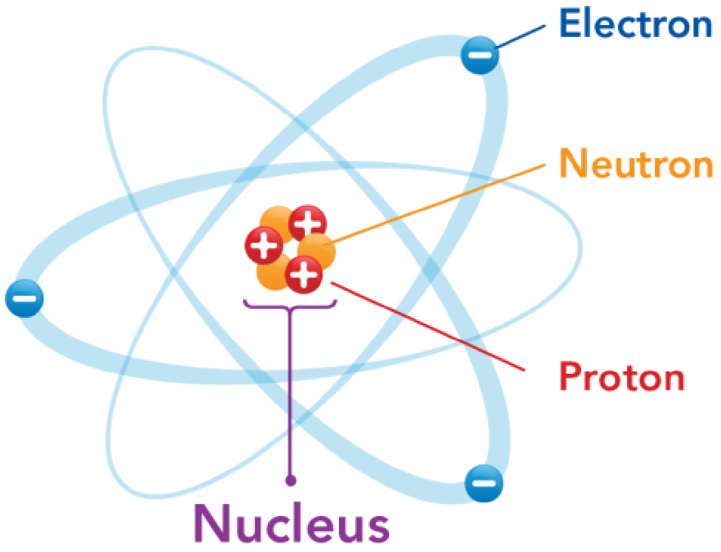






Chapter 17

The atom

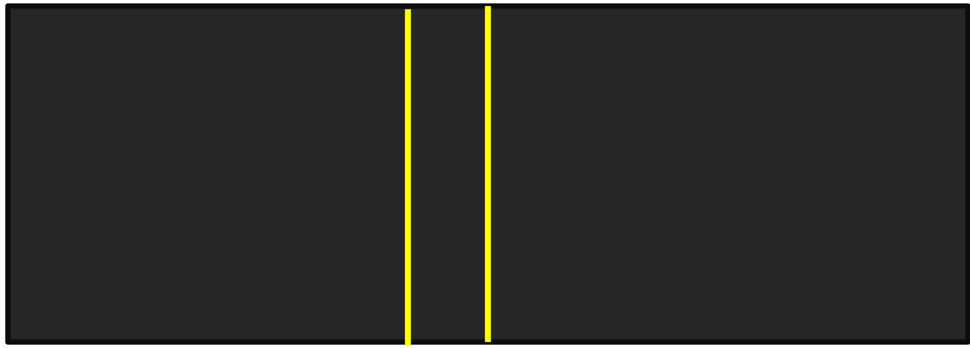
Prepared & Presented by: Mr. Mohamad Seif

OBJECTIVES

1. **Doublet – lines**
2. **Excitation of the atom by an electron**

Doublet – lines

An energy level is called a doublet, if it consists of two closely spaced levels.



The absorption and emission spectra of an atom having such a level include two closely spaced lines called doublet lines.

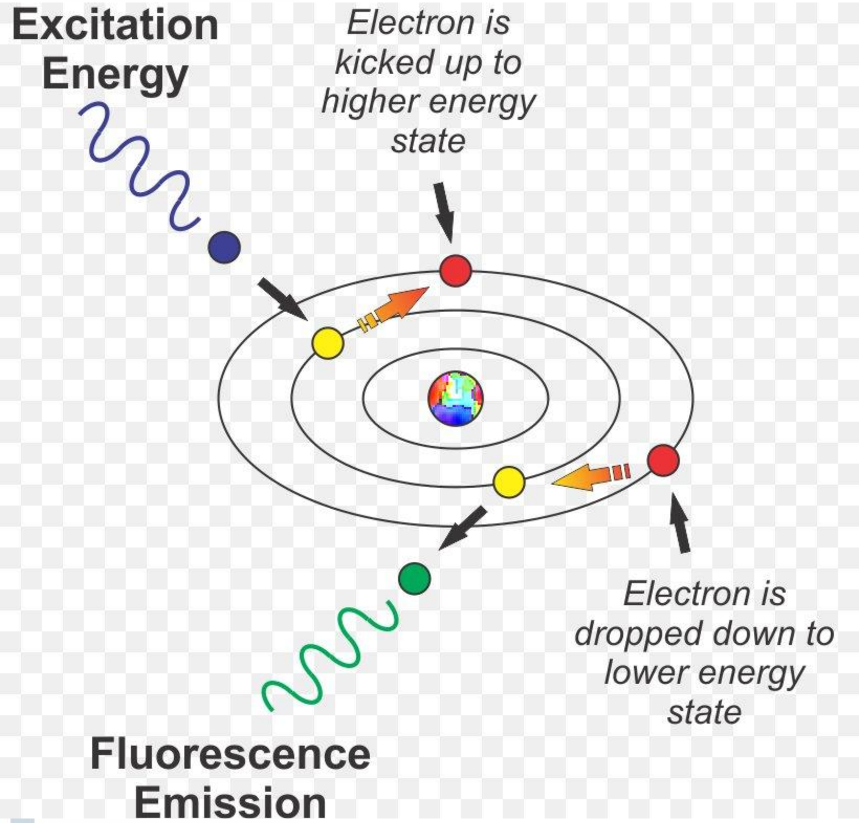
These two lines have very close wavelengths.

The sodium emission spectrum is dominated by the bright doublet known as the Sodium D-lines at $588.995nm$ and $589.592nm$

Excitation of the atom by an electron

The atom can be excited when bombarded by an electron.

An electron causes the transition of an atom from an energy level E_l, to a higher energy level E_n if its kinetic energy KE before, is at least equal to the difference $(E_h - E_l)$ between the energies of these two levels.



$$KE_{before} \geq E_h - E_l$$

Excitation of the atom by an electron

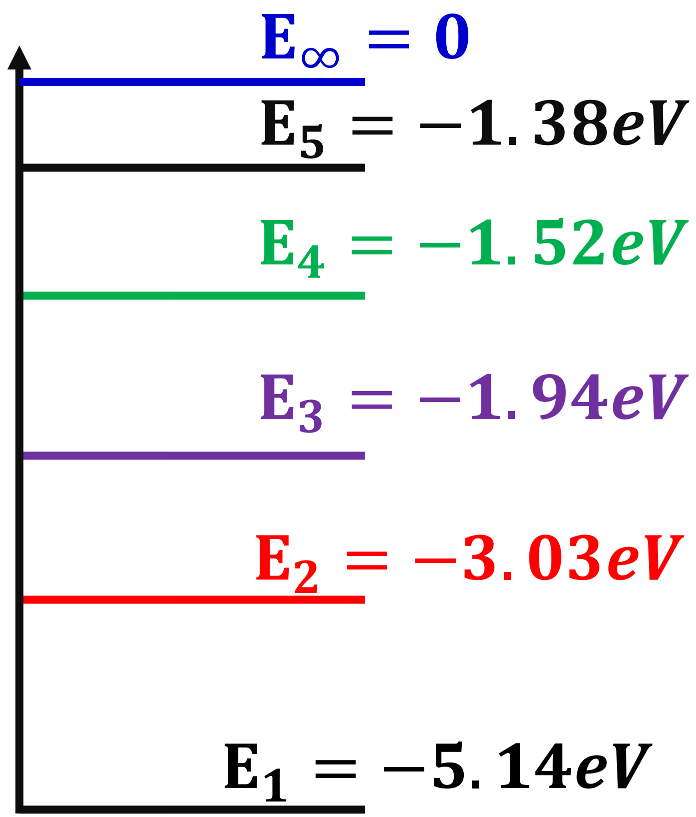
The atom absorbs from the electron an amount of energy enough to ensure a transition.
The rest of the energy is carried by the electron as kinetic energy. The kinetic energy carried by the electron after exciting the atom is:

$$KE_{after} = KE_{before} - (E_h - E_l)$$

Excitation of the atom by an electron

Application 3:

Given the energy-level diagram of the sodium atom. $1eV = 1.6 \times 10^{-19} J$



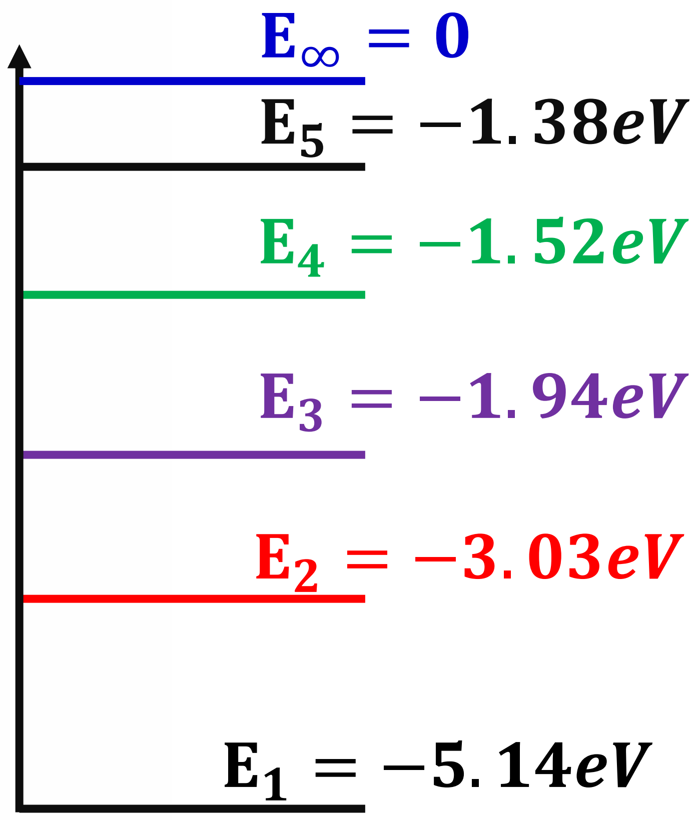
1) An electron of kinetic energy KE, hits the sodium atom when it is in the ground state. Determine what would happen to the sodium atom if $KE_e = 6eV$

2) An electron of kinetic energy $KE_e = 16eV$ hits the sodium atom when it is in the first excited state. Determine the possible values of the kinetic energy carried by the leaving electron.

Excitation of the atom by an electron

$1eV = 1.6 \times 10^{-19} J$; $KE_e = 6eV$

1) An electron of kinetic energy KE, hits the sodium atom when it is in the ground state.



Determine what would happen to the sodium atom if $KE_e = 6eV$

$W_{ion} = E_\infty - E_1 = 0 - (-5.14) = 5.14eV$

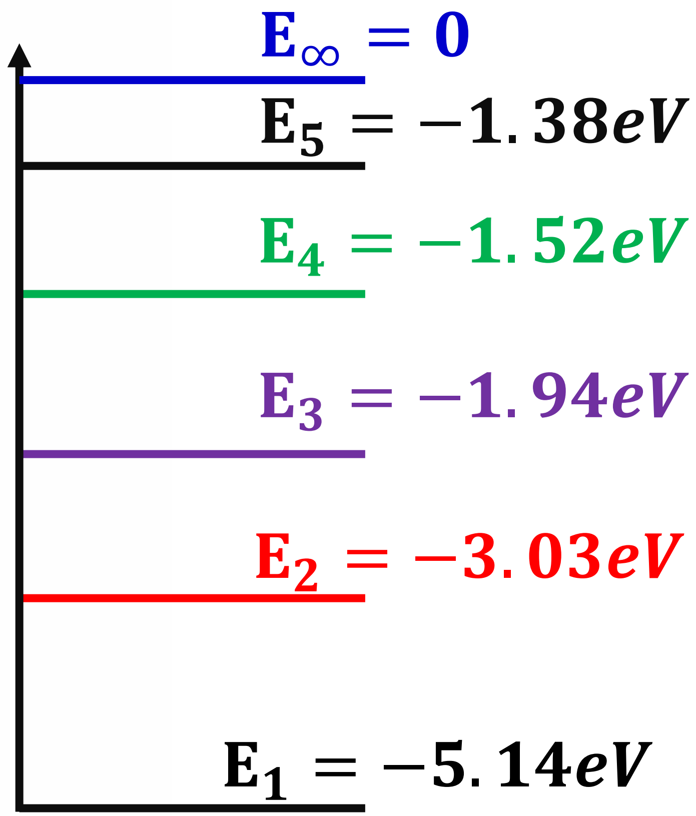
Since $KE_e = 6eV > W_{ion} = 5.14eV$ then:

The Sodium atom may undergo an electronic transition to a higher energy level

Excitation of the atom by an electron

$1eV = 1.6 \times 10^{-19} J$; $KE_e = 6eV$

2)An electron of kinetic energy $KE_e = 16eV$ hits the sodium atom when it is in the first excited state.



Determine the possible values of the kinetic energy carried by the leaving electron.

$KE_e = E_h - E_l$ ⇨ $E_h = KE_e + E_l$

$$E_h = 1.6 + (-3.03) = -1.43eV$$

We notice that $E_4 < E_h < E_5$ then sodium may be excited to E_3 or E_4

Excitation of the atom by an electron

If the atom excited to E_3; then: $KE_{after} = KE_{before} - (E_h - E_l)$

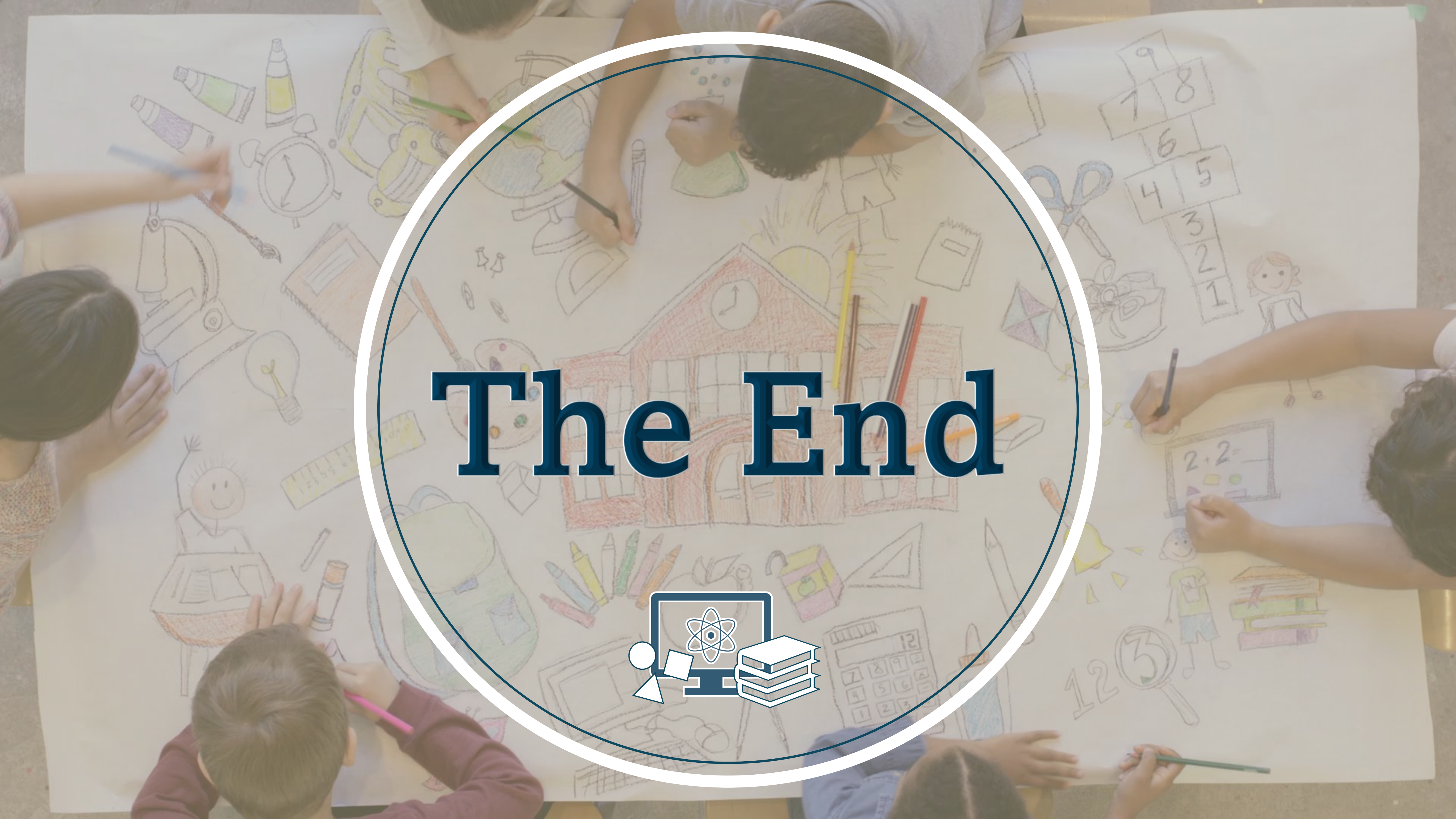
$$KE_{after} = 1.6 - (-1.94 - (-3.03))$$

$$KE_{after} = 0.51eV$$

If the atom excited to E_4; then: $KE_{after} = KE_{before} - (E_h - E_l)$

$$KE_{after} = 1.6 - (-1.52 - (-3.03))$$

$$KE_{after} = 0.09eV$$



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