Atomic and Nuclear Physics

Week #3

Atomic constituents, atomic models & hydrogen spectrum

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Over the slides

- ➢Atomic nuclear constituents
- ➢Plum pudding model of atom
- ➢Rutherford’s gold foil experiment
- ➢Origin of atomic spectra
- ➢Hydrogen spectrum

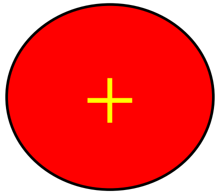
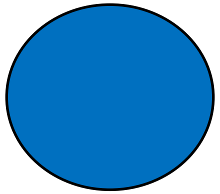
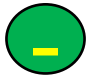
Atom – a brief history

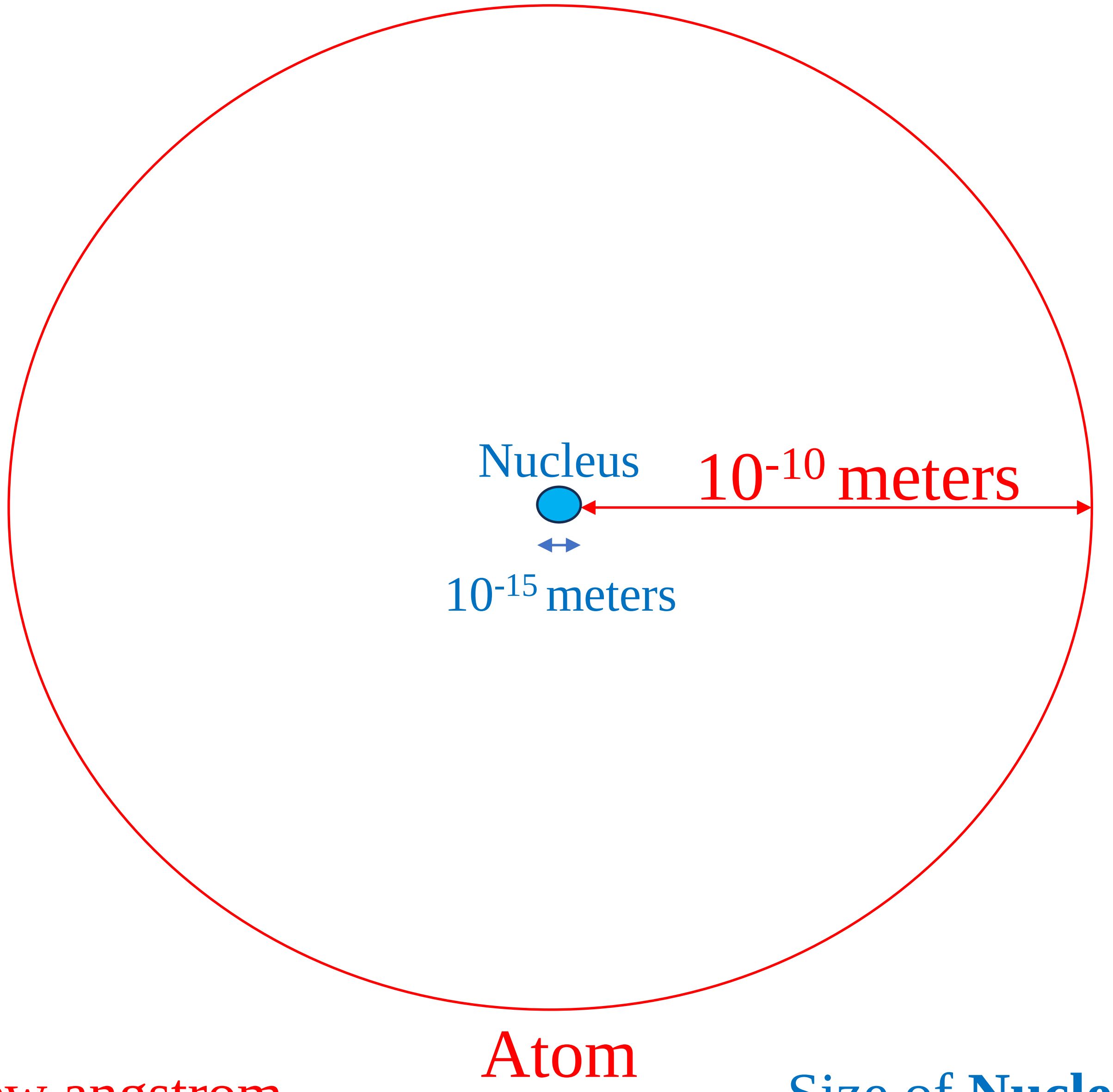
- Concept of **indivisible 'atom**' by Greek Philosophers (5th century BC) like **Democritus** –
- Purely philosophical notion with no experimental evidence.
- Idea of element – each element characterized by unique 'Atom' with specific masses (early 19th century) by **John Dalton**
- Provided more scientific basis.
- This theory laid the groundwork for modern atomic theory.

- Discovery of Subatomic Particles (late 19th century) **J.J. Thomson** discovered the electron in 1897
- Demonstrated that atoms are divisible
- Put forward "plum pudding" model of the atom.
- Nuclear Model (early 20th century) by Ernest Rutherford's famous gold foil experiment in 1909
- Experimental confirmation for the nuclear model of the atom, with electrons orbiting the nucleus.

- Quantum model of atom (1913) by Niels Bohr
- Electrons move in discrete energy levels or orbits around the nucleus and can jump between these levels by emitting or absorbing energy in the form of photons.
- This model explained the spectral lines of hydrogen and laid the foundation for quantum mechanics.
- Neutrons were discovered (1932) by Chadwick while alpha particle interaction with Beryllium nucleus

Building blocks of atoms

	Protons : Charge : Positive Mass $\approx$ 1 amu Number of protons (or electrons) determines the chemical element – The number is atomic number of the atom
	Neutrons: Charge : Nil Mass $\approx$ 1 amu Number of protons + neutrons determines the mass number
	electrons : Charge : Negative Mass $\approx$ (1/1836) amu Number of electrons (or protons) determines the chemical element

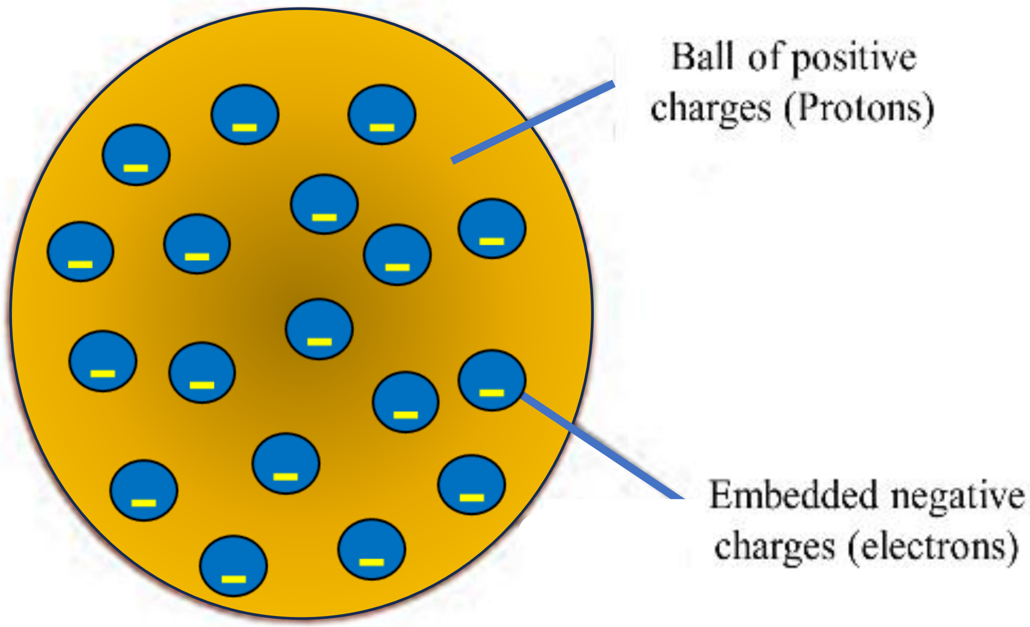


Size of **Atoms** is a few angstrom (Å) **10^{-10}** meters.

Size of **Nucleus** is a few femtometer (fm) **10^{-15}** meters.

- Plum pudding model of atom
- Electrons embedded on the cloud of protons
- J. J. Thompson’s atom model

Plum pudding model



- Plum pudding model is built up with reason that atom is neutral
- No evidence for clustering of positive or negative charges
- No idea of free space within the atom
- No other experimental evidence was available to refute the concept of ‘Plum pudding’ model
- Rutherford’s plan was to verify these facts experimentally

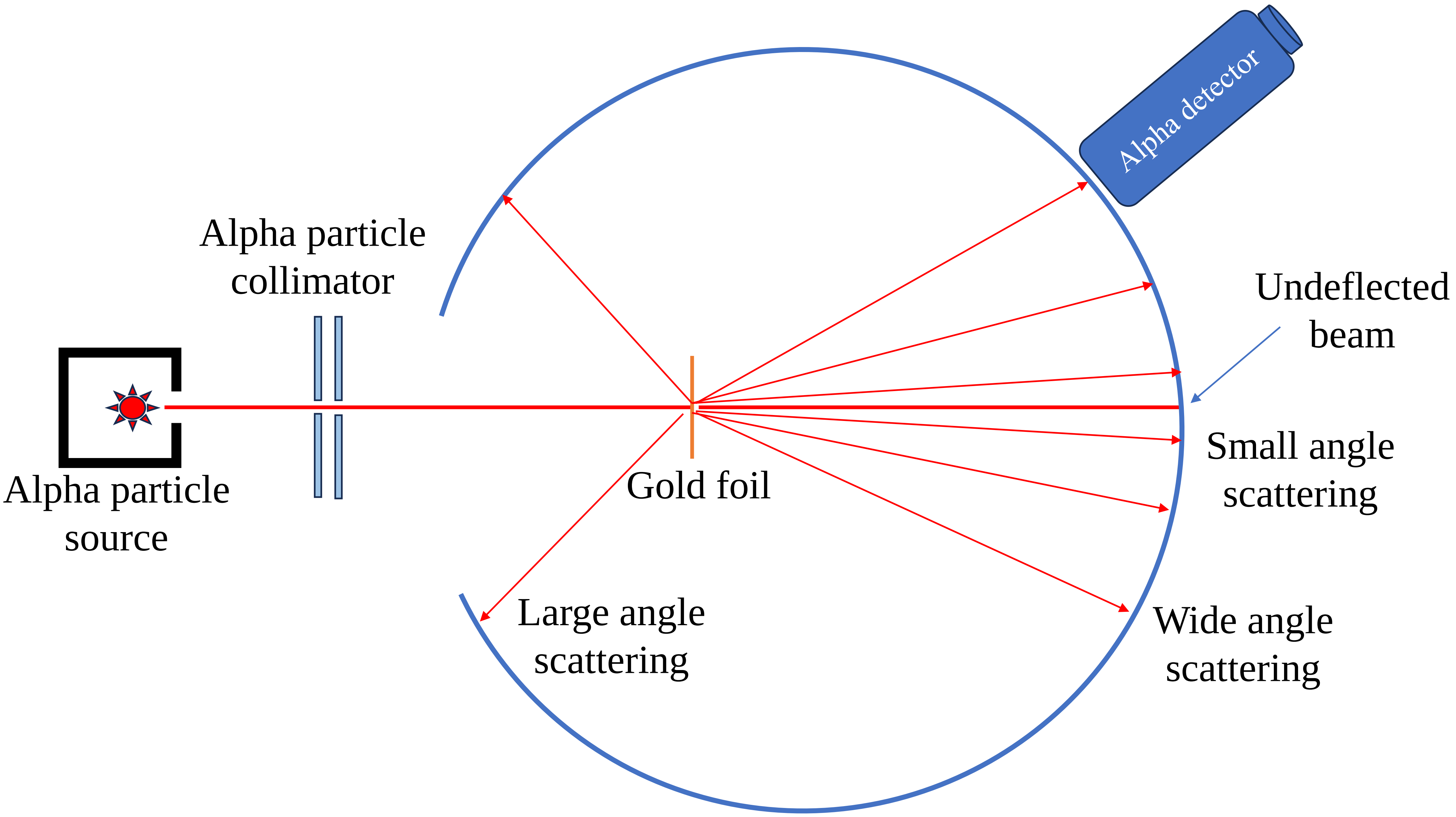
Alpha-particle scattering and the Rutherford model.

- ✓ The experiment was conducted by Ernest Rutherford in 1909 at the University of Manchester.
- ✓ The prevailing model at that time was the Thomson model, suggesting that the atom was a "plum pudding" with positive charge spread evenly throughout.
- ✓ Rutherford's experiment aimed to test this model and investigate the nature of the atomic structure.

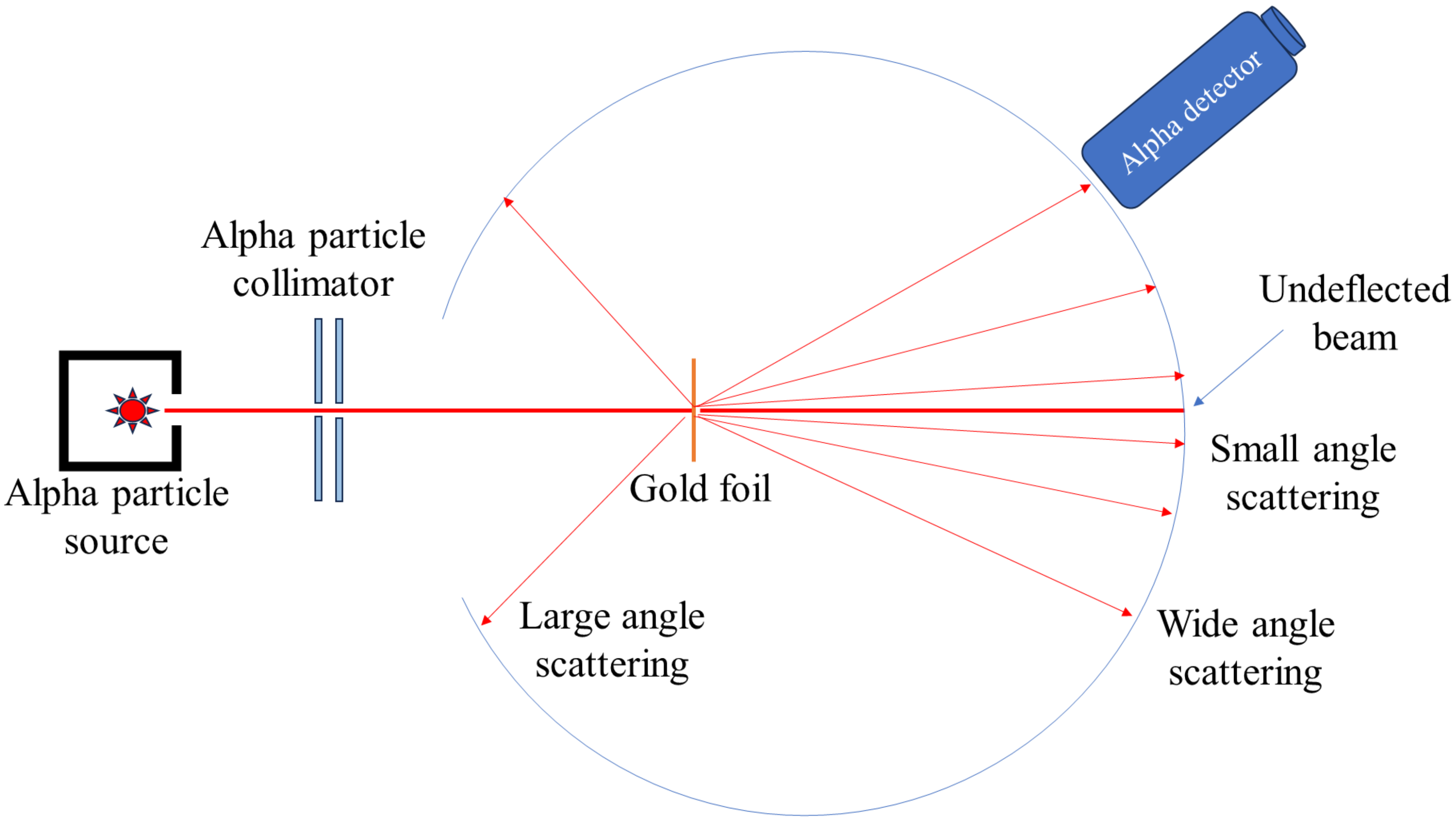
Experimental Setup

- Rutherford's team used a beam of alpha particles emitted from a radioactive source (radium).
- The alpha particles were directed towards a thin gold foil.
- Around the gold foil, they placed a circular fluorescent zinc sulfide screen to detect the scattered alpha particles.

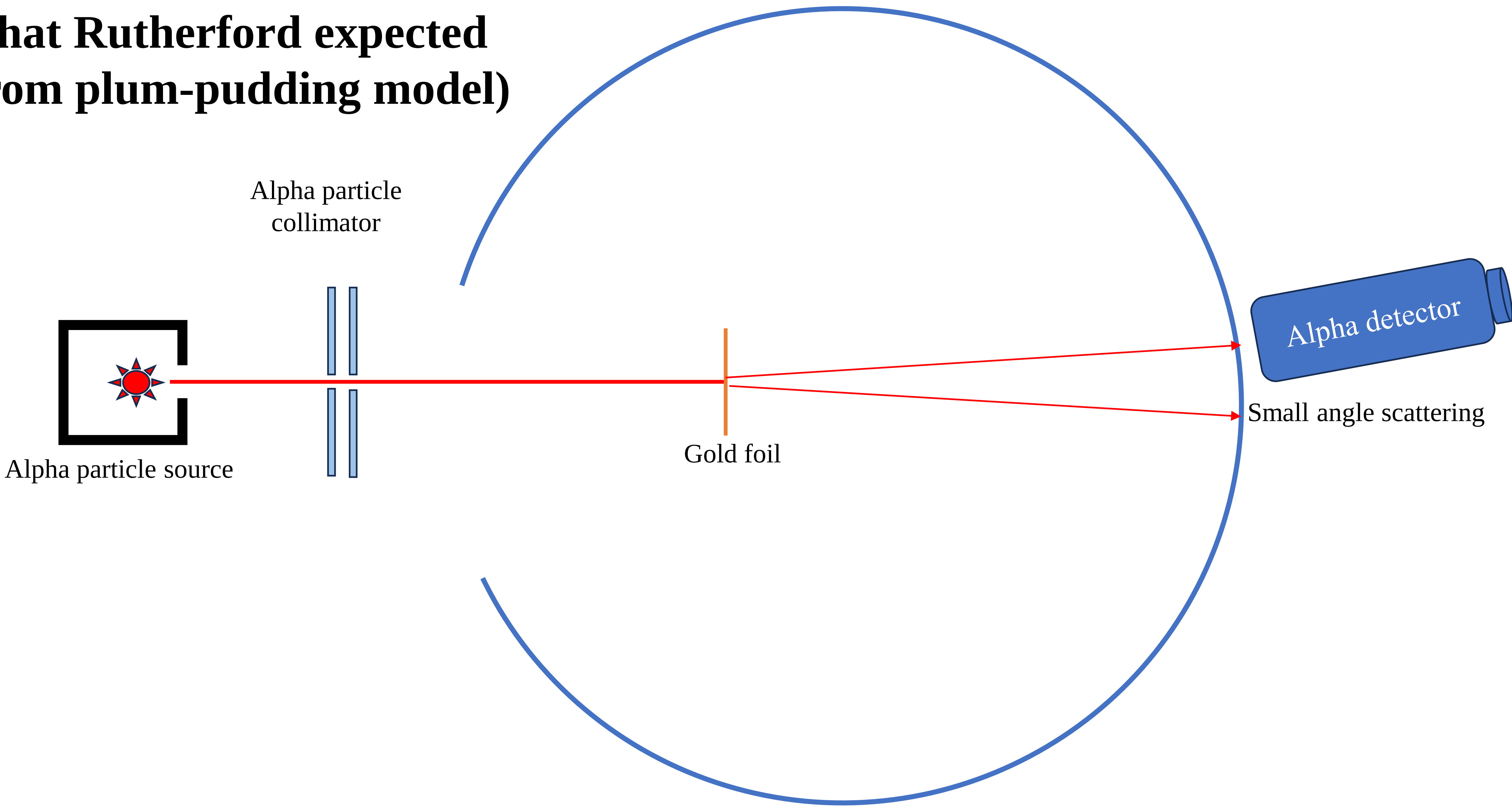
- According to Thomson's model, most alpha particles should pass through the gold foil with only minor deflections.
- This is because the positive charge was believed to be spread uniformly throughout the atom.


Alpha particle collimator
Alpha detector
Undeflected beam
Alpha particle source
Gold foil
Small angle scattering
Large angle scattering
Wide angle scattering

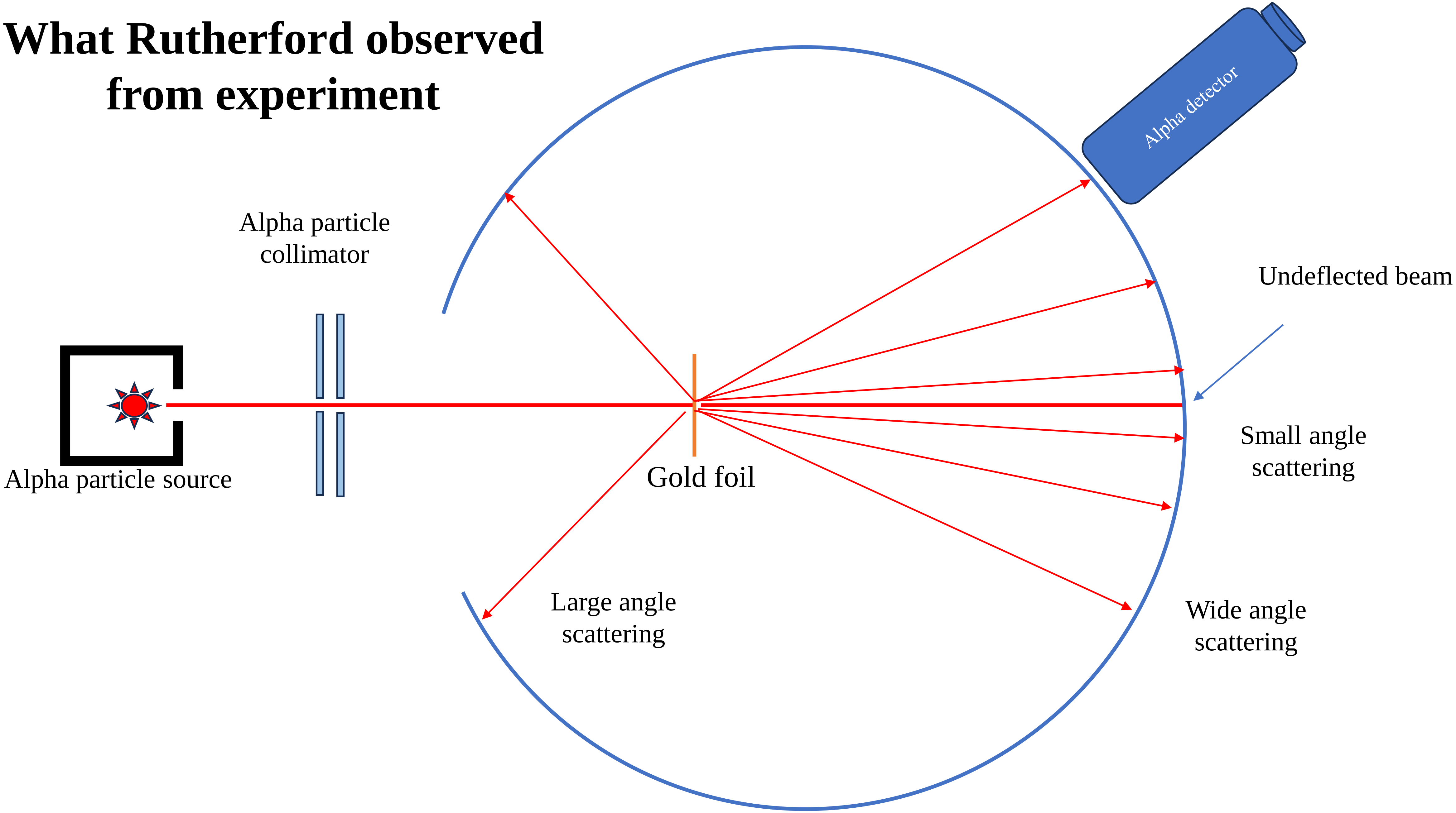
Rutherford expected only small angle scatterings from the gold foil, as the interaction would be purely week electrostatic.



What Rutherford expected (from plum-pudding model)



What Rutherford observed from experiment



Surprising Observations

- ✓ The **majority of alpha particles passed through the gold foil** with little or no deflection.

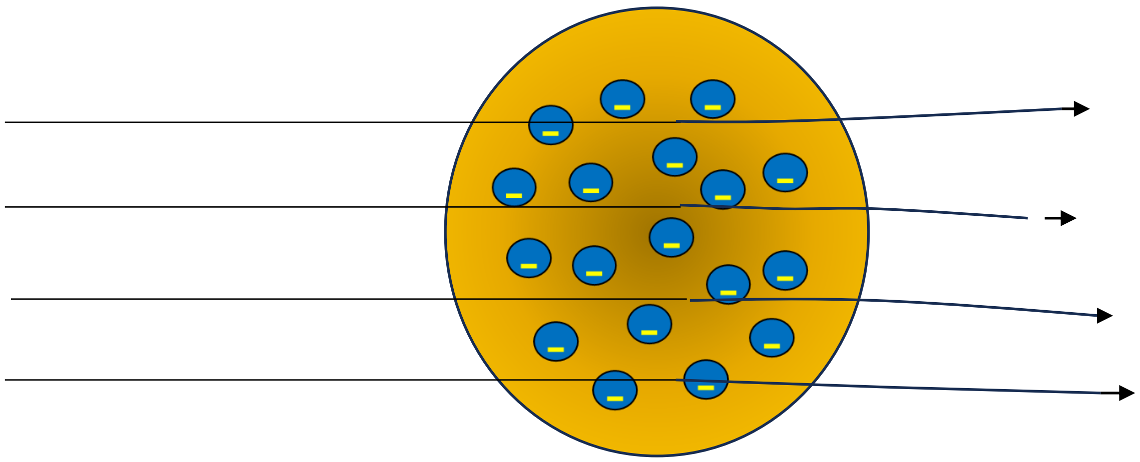
- ✓ However, a small fraction of alpha particles experienced **large-angle deflections or even bounced back** in the opposite direction.

Rutherford's Interpretation

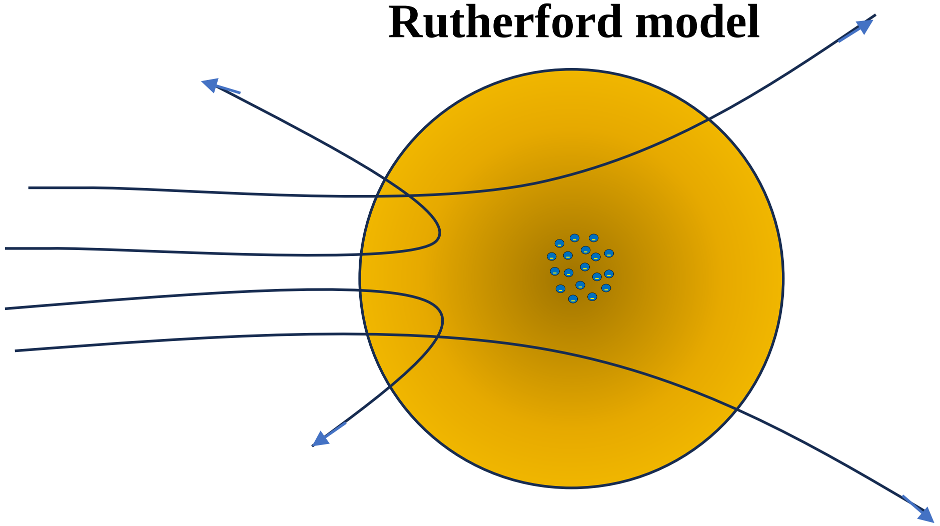
- Rutherford concluded that the alpha particles' deflections were due to a **concentrated positive charge in the atom's center, which he called the "nucleus.“**
- He proposed that the **nucleus was tiny** compared to the entire atom's size.
- This discovery led to the development of a new **nuclear model of the atom**.

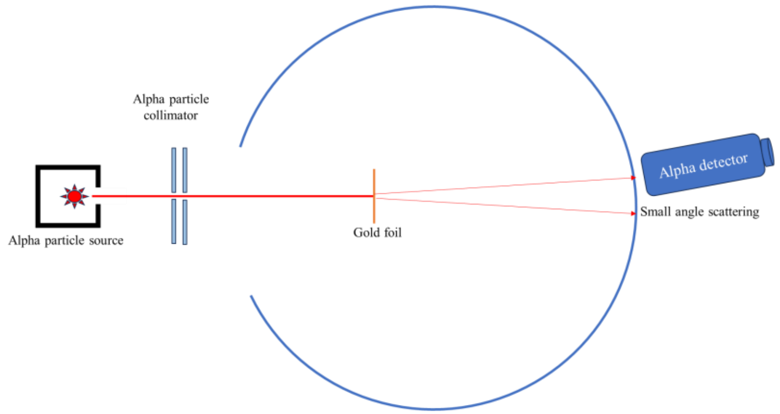
Alpha-particle scattering and the Rutherford model.

Plum pudding Model

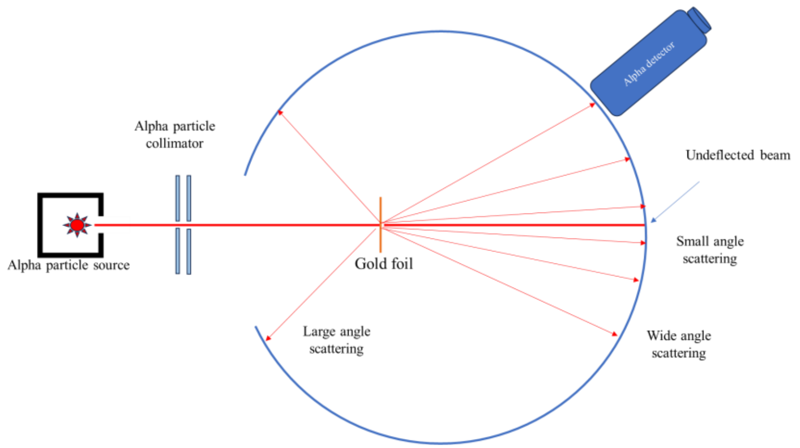


Rutherford model





Expected result



Observed result

Conclusions from the experiment

1. The results of this experiment **were not in agreement with the plum-pudding model** of the atom as suggested by Thomson.
2. Since **large number of alpha particles were passed un deflected**, he concluded that a good amount of atom is empty.
3. For **alpha particles to be deflected back**, there should be a positively charged region producing large repelling force. Therefore, he concluded that the positive charges of the atom are concentrated in the centre.
4. He then suggested the '**nuclear model of an atom**' wherein the entire positive charge and most of the mass of the atom is concentrated in the nucleus.

5. The alpha particles were deflected or scattered through a large angle on coming close to the nucleus.

6. Also, the electrons having negligible mass, do not affect the trajectory of these incident alpha particles.

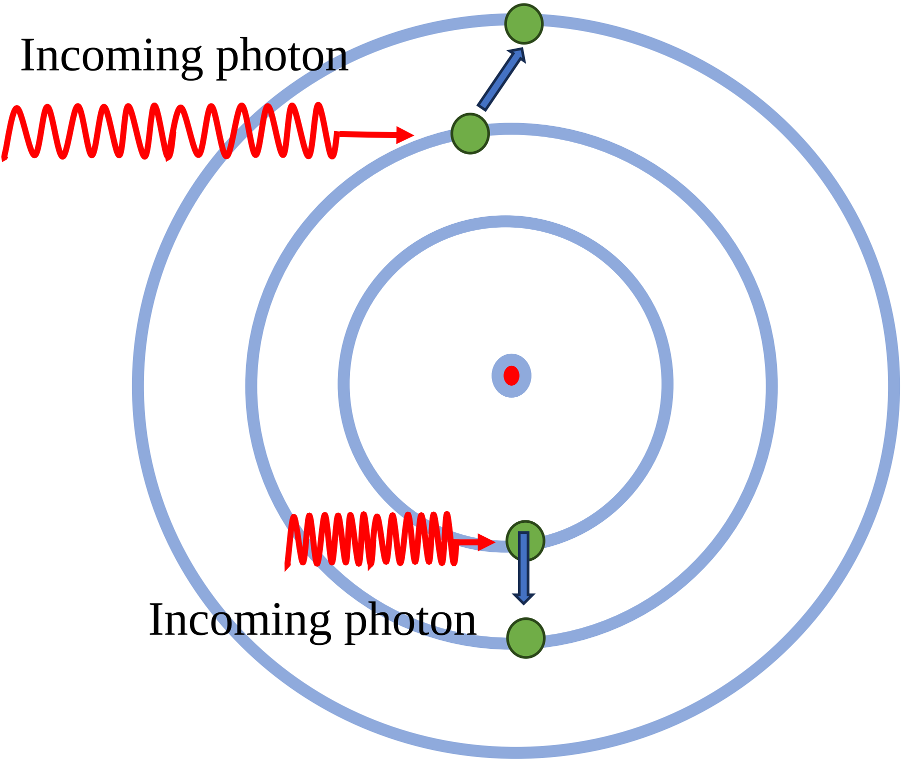
7. Also, the electrons are moving in orbits around the nucleus similar to the planets and the sun.

8. Finally, from his experimental data, Rutherford concluded that the size of the nucleus is between 10^{-15} and 10^{-14}m.

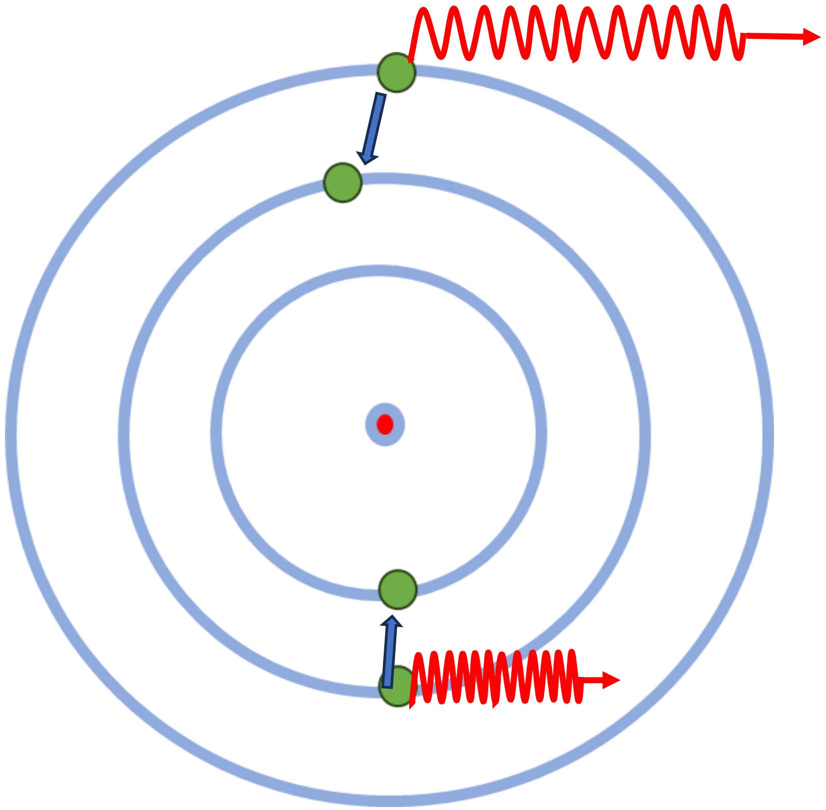
Atomic spectra

When a substance absorbs energy and radiates, it occurs in the form of a spectrum

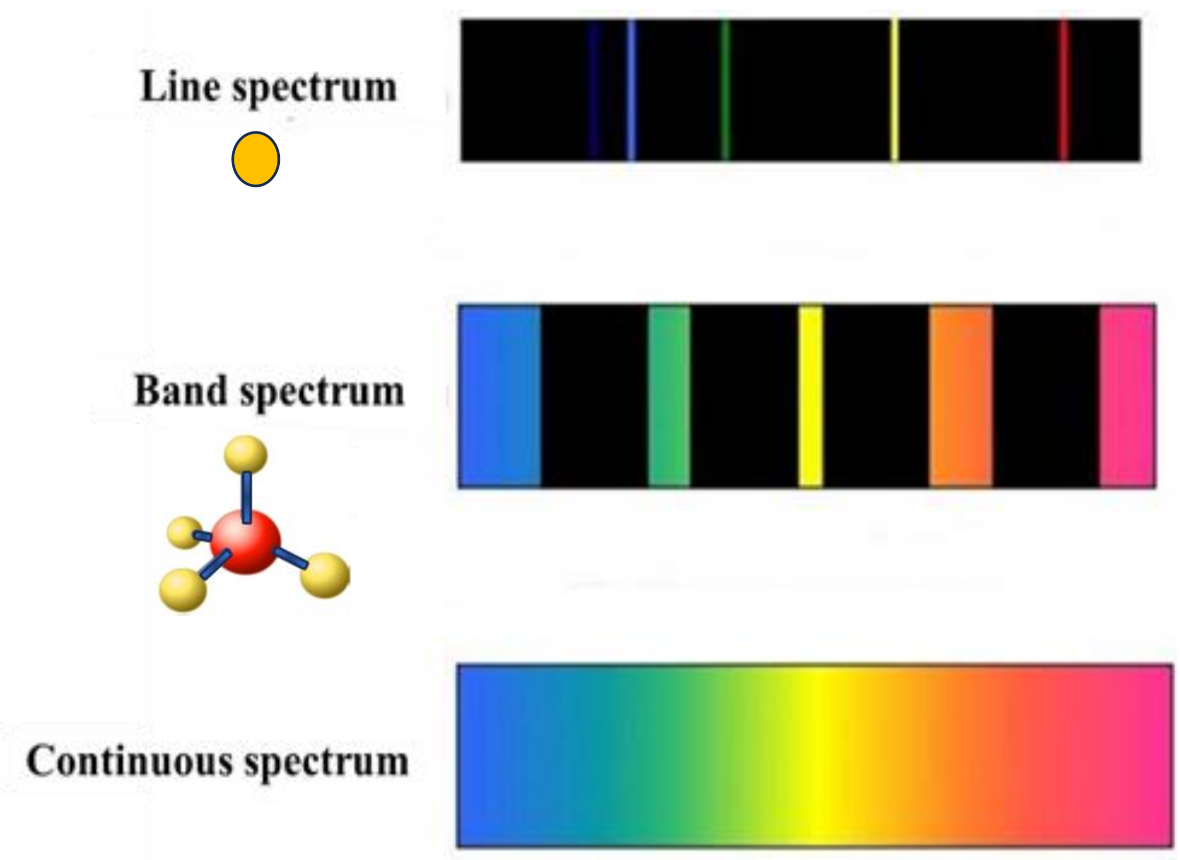
Absorption of photon by electrons



Emission of photon by electrons



- When atoms and molecules absorb energy from a source, they become excited and re-emit the absorbed energy in a very short period of time (typically in **10^{-8}s**).
- The spectrum of emitted light is called **emission spectrum**.
- Depending on the emitting substance medium, there can be **line, band or continuous**

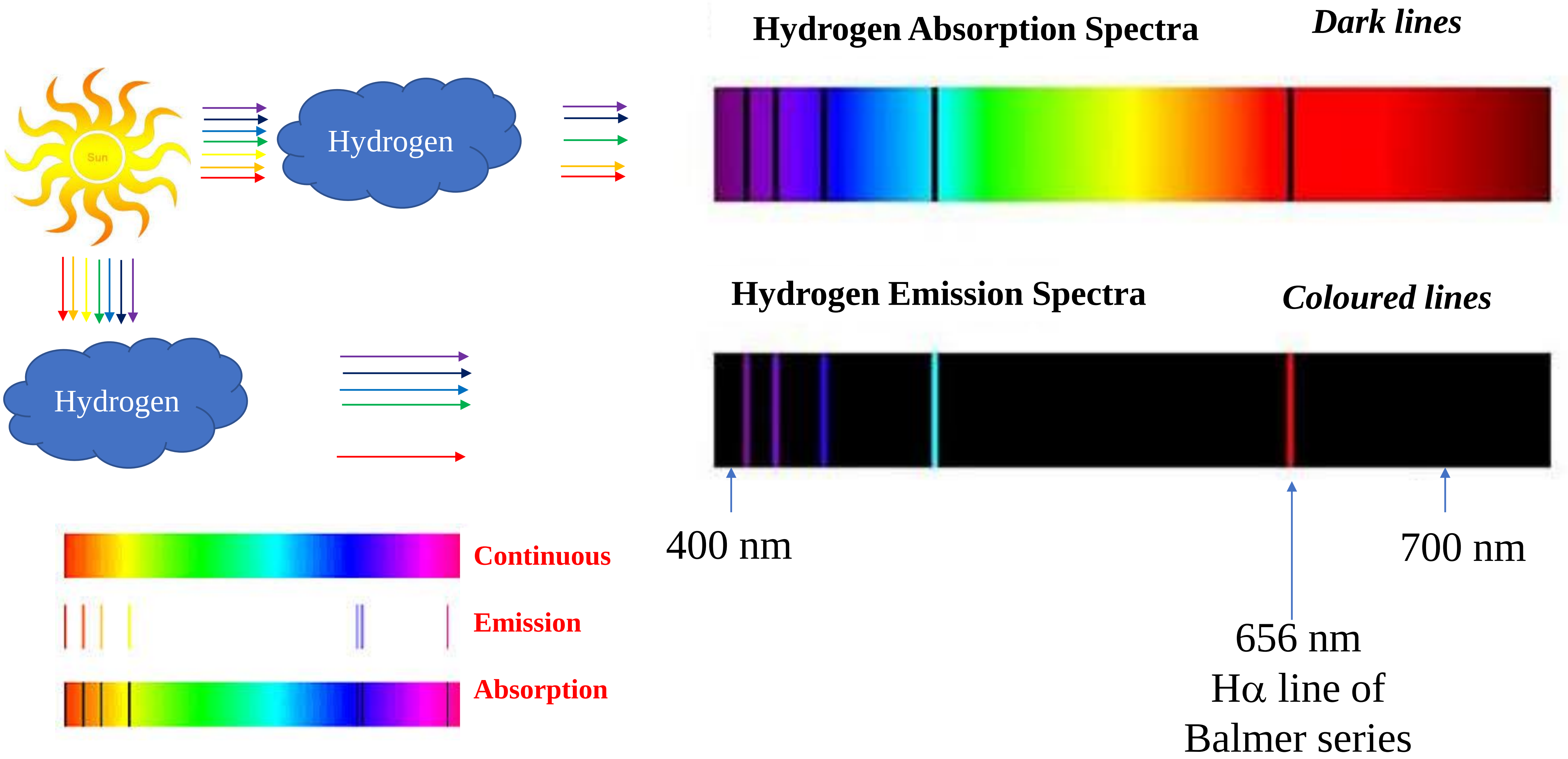


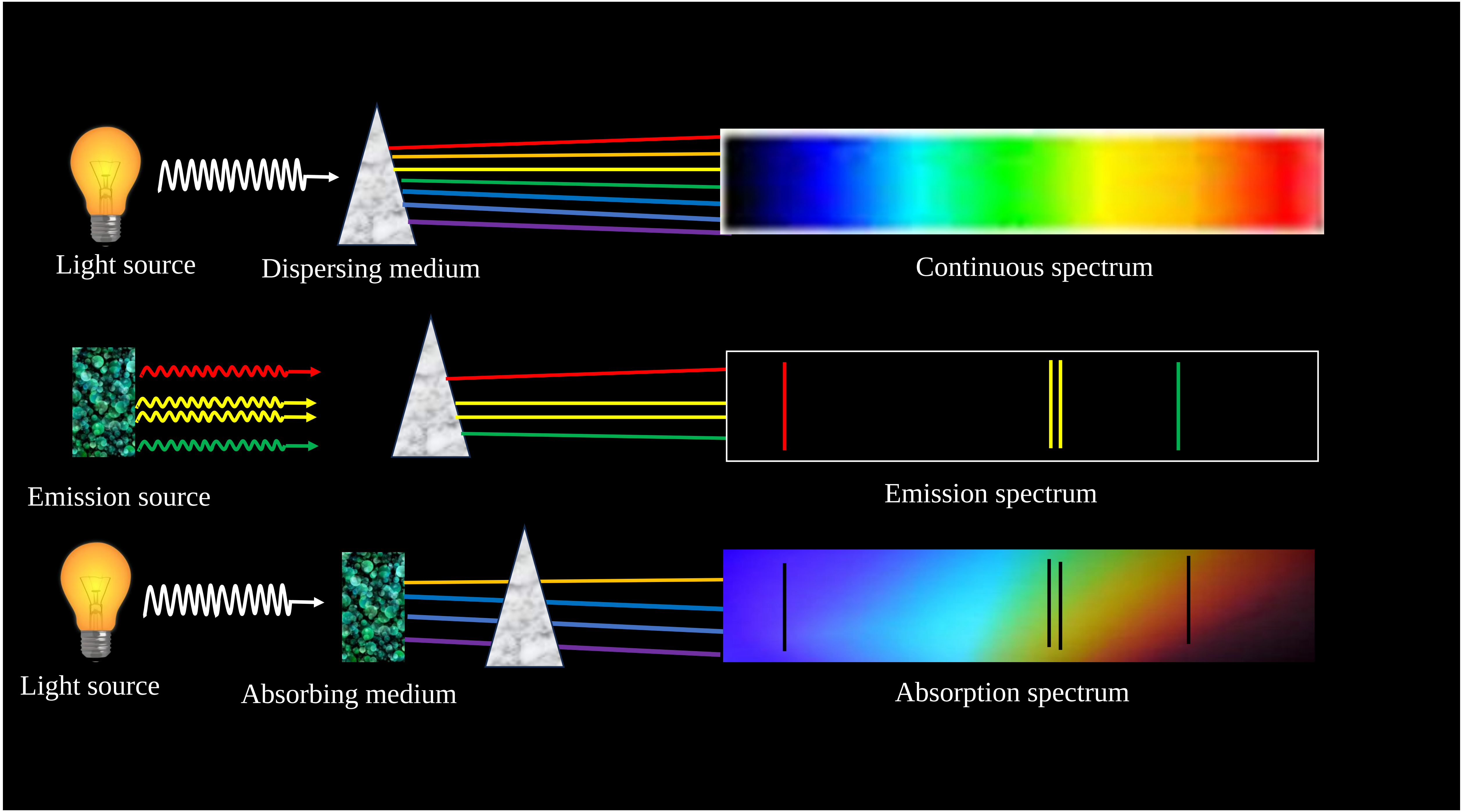
- Atomic system (Hydrogen) - line spectrum
- Specific wave lengths will be seen as coloured lines.

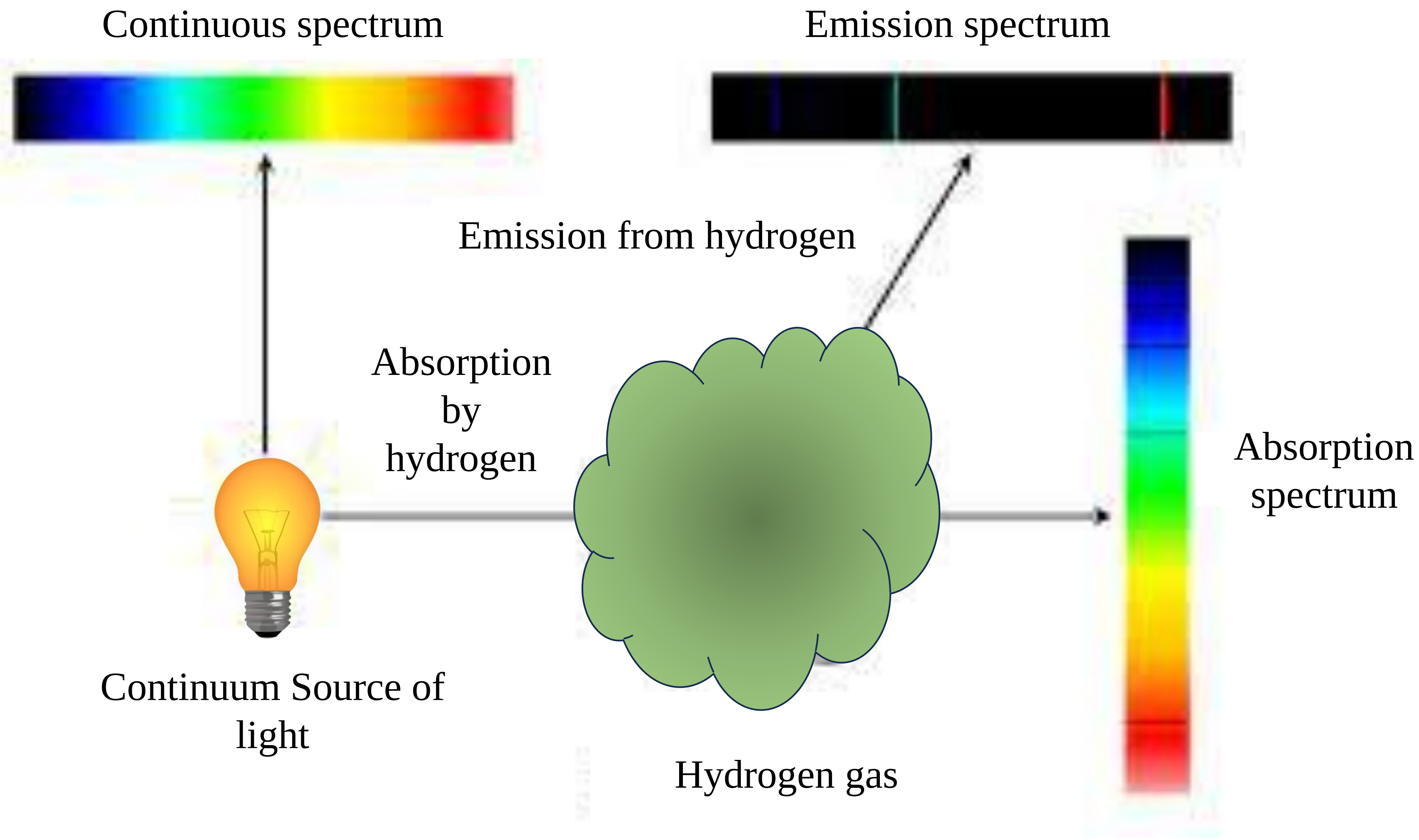
- For a molecular system (Benzene)
- Emitted spectrum - broader lines - bands.

- A hot cloud or incandescent lamp
- Continuous display of colours - continuous spectrum is produced.

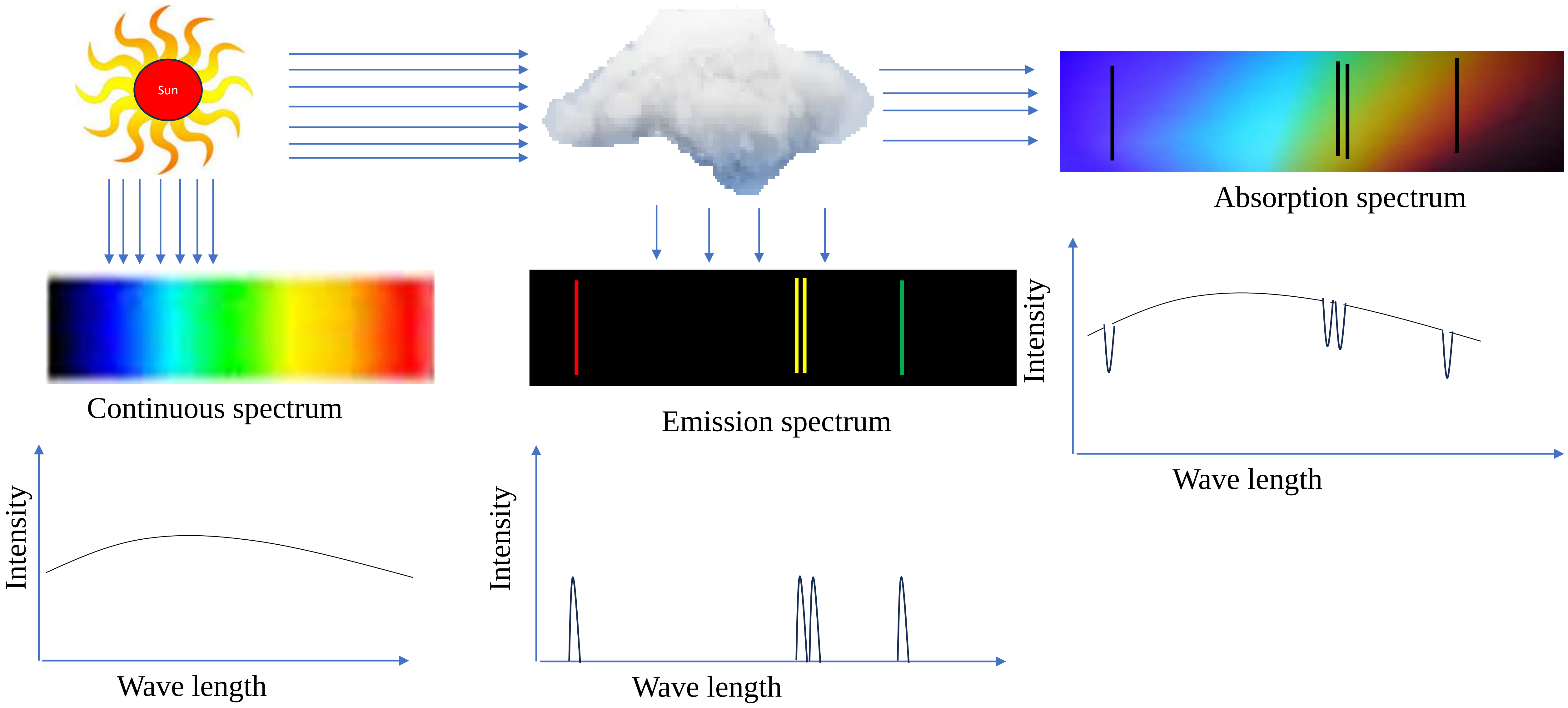
Emission & Absorption Spectra




Light source
Dispersing medium
Continuous spectrum
Emission source
Emission spectrum
Light source
Absorbing medium
Absorption spectrum



Continuous spectrum
Emission spectrum
Emission from hydrogen
Absorption
by
hydrogen
Absorption
spectrum
Continuum Source of
light
Hydrogen gas


Sun
Absorption spectrum
Continuous spectrum
Emission spectrum
Intensity
Wave length
Intensity
Wave length
Intensity
Wave length

Hydrogen Spectrum

1.The electron in a hydrogen atom travels around the nucleus in a circular orbit.

2.The energy of the electron in an orbit is proportional to its distance from the nucleus.

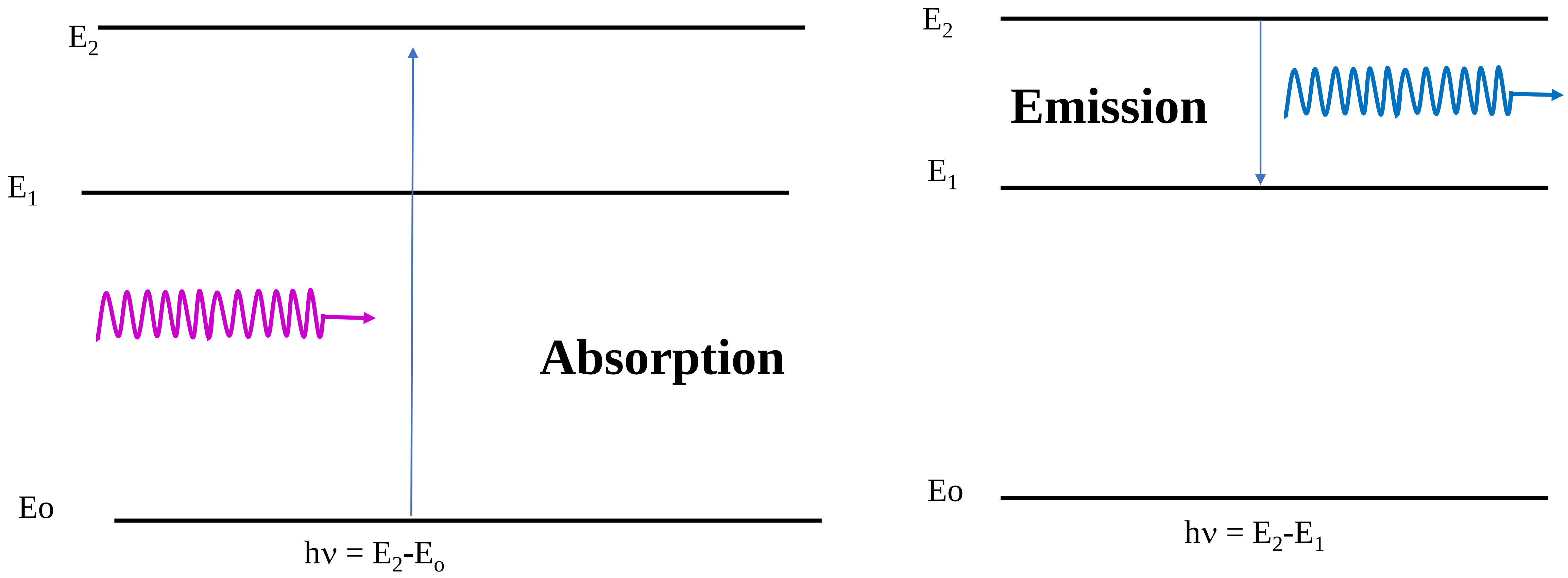
3.Only a limited number of orbits with certain energies are allowed. In other words, the orbits are quantized.

4.The orbits that are allowed are those for which the angular momentum of the electron is an integral multiple of Planck's constant divided by 2π.

5.Light is absorbed when an atomic electron excites to a higher energy orbit and emitted when an electron de-excite to a lower energy orbit.

6.The energy of the light emitted or absorbed is exactly equal to the difference between the energies of the orbits.

Bohr theory and the hydrogen spectrum.



- Bohr assumed that there are only a limited number of orbits in which the electron can reside.
- He based this assumption there are only a limited number of lines in the spectrum
- Bohr restricted the number of orbits on the hydrogen atom by limiting the allowed values of the angular momentum of the electron.
- Bohr assumed that the angular momentum of the electron is equal to an integer times reduced Planck's constant

That is, angular momentum, $\mathbf{L = mvr = n\frac{h}{2\pi} = n\hbar \quad (n = 1,2,3,4\ldots.)}$

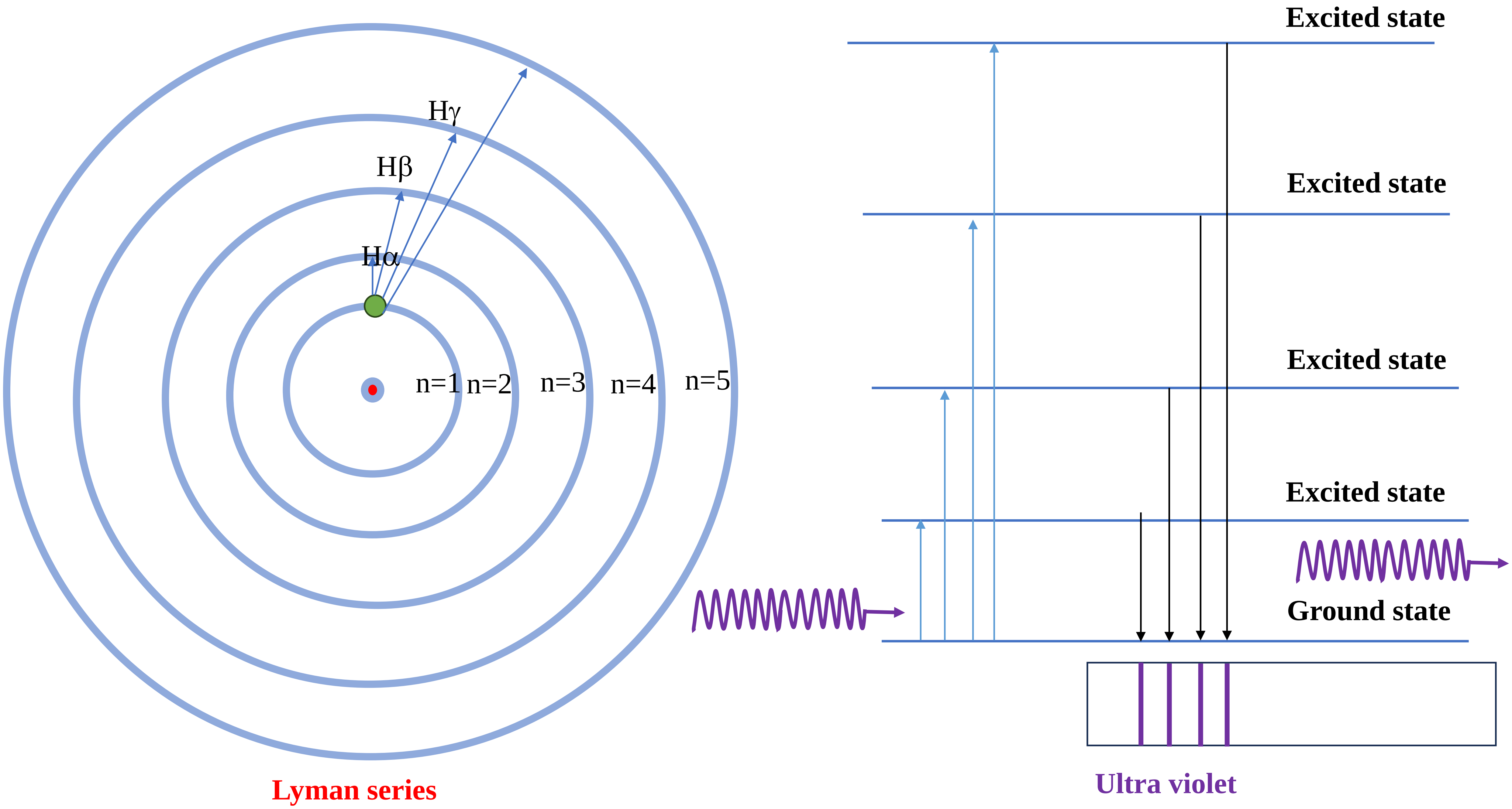
- ✓ An object moving in a circular orbit has an angular momentum equal to its mass (m) times the velocity (v) times the radius of the orbit (r).
- ✓ Bohr assumed :
- ✓ Angular momentum of the electron = (Integer × Planck's constant) /2π.
- ✓ That is, angular momentum, $L = mvr = n\frac{h}{2\pi} = n\hbar$ (n = 1,2,3,4….)

➢ Rydberg (1890) found that all the lines of the atomic hydrogen spectrum could be fitted to a simple empirical formula:

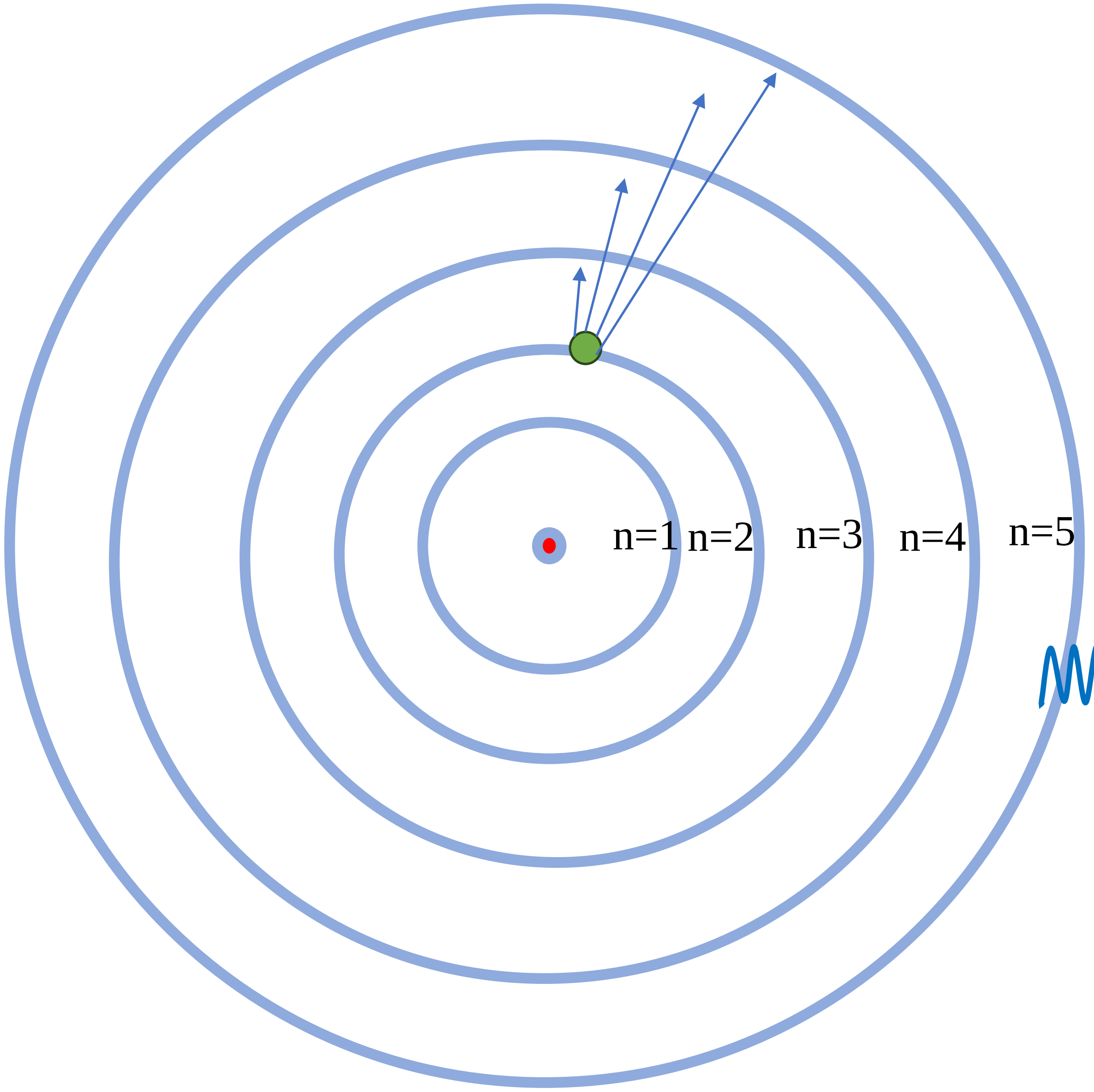
$$\bar{\nu} = \frac{1}{\lambda} = \mathbf{R}\left[\frac{1}{(n_1)^2} - \frac{1}{(n_2)^2}\right]$$

where n = 1,2,3,4……… and $n_2 > n_1$.

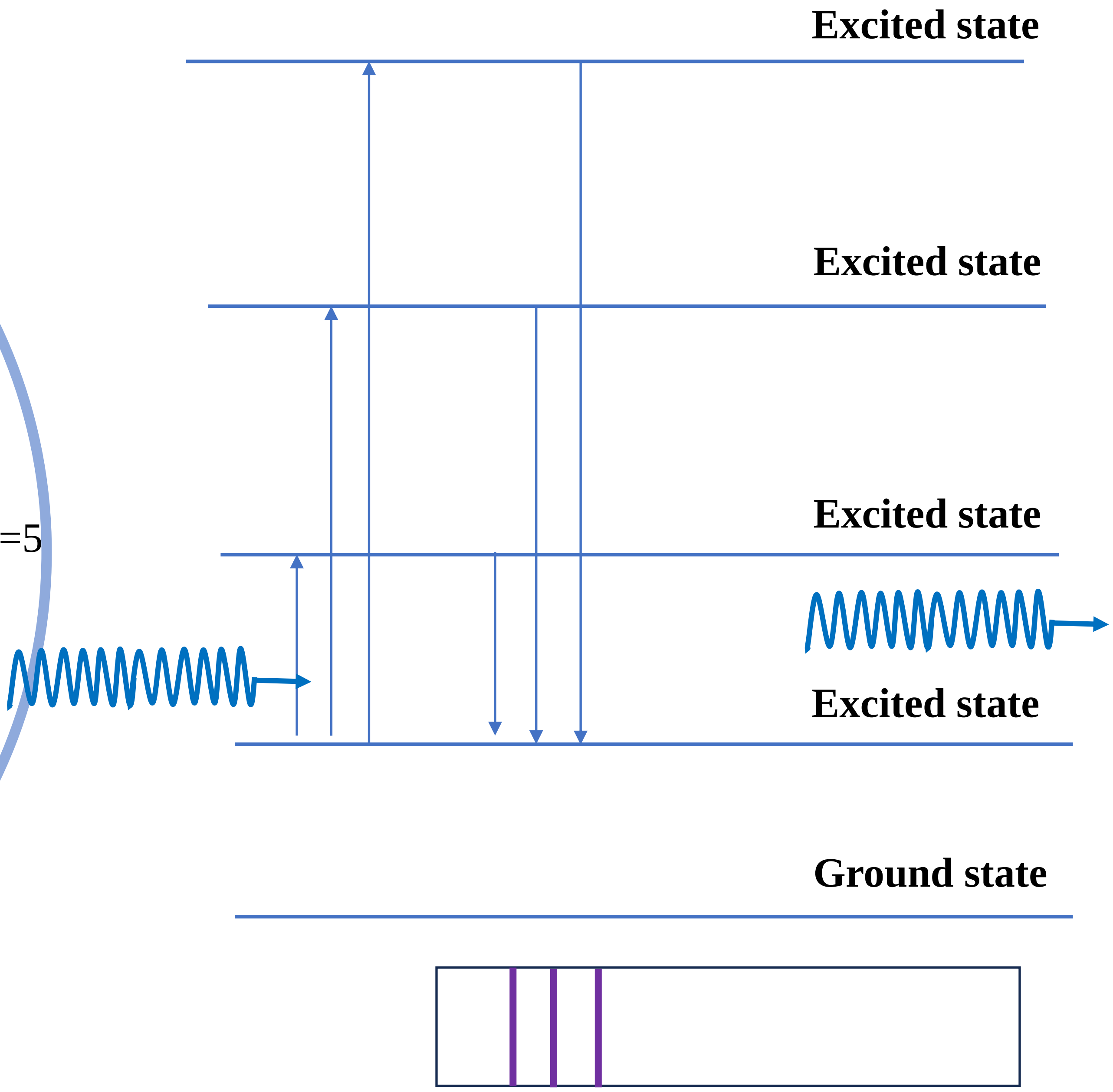
R is the Rydberg constant =109,677 cm^{-1} = $1.09\times10^{7}\ m^{-1}$.



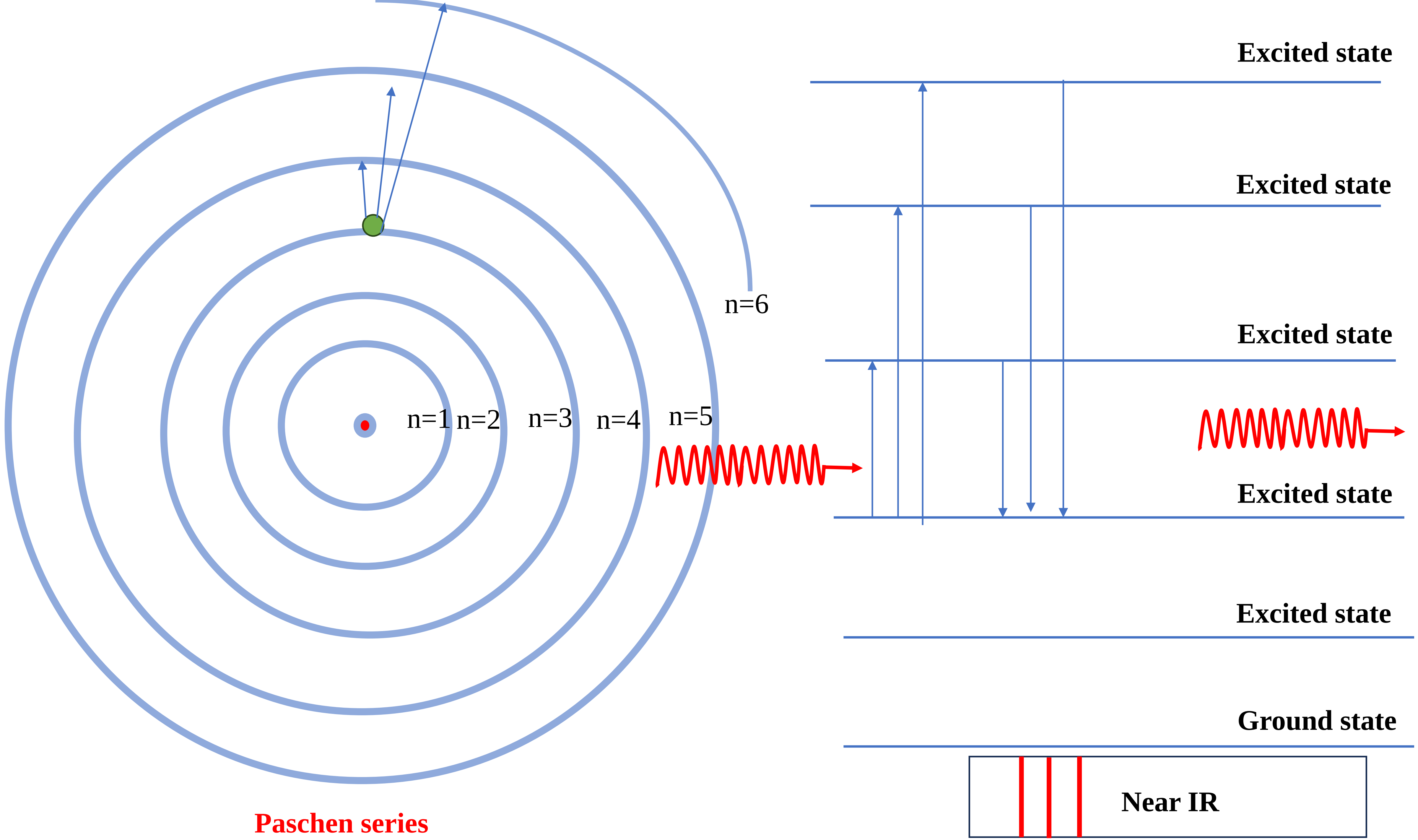
Hγ
Hβ
Hα
n=1
n=2
n=3
n=4
n=5
Lyman series
Excited state
Excited state
Excited state
Excited state
Ground state
Ultra violet



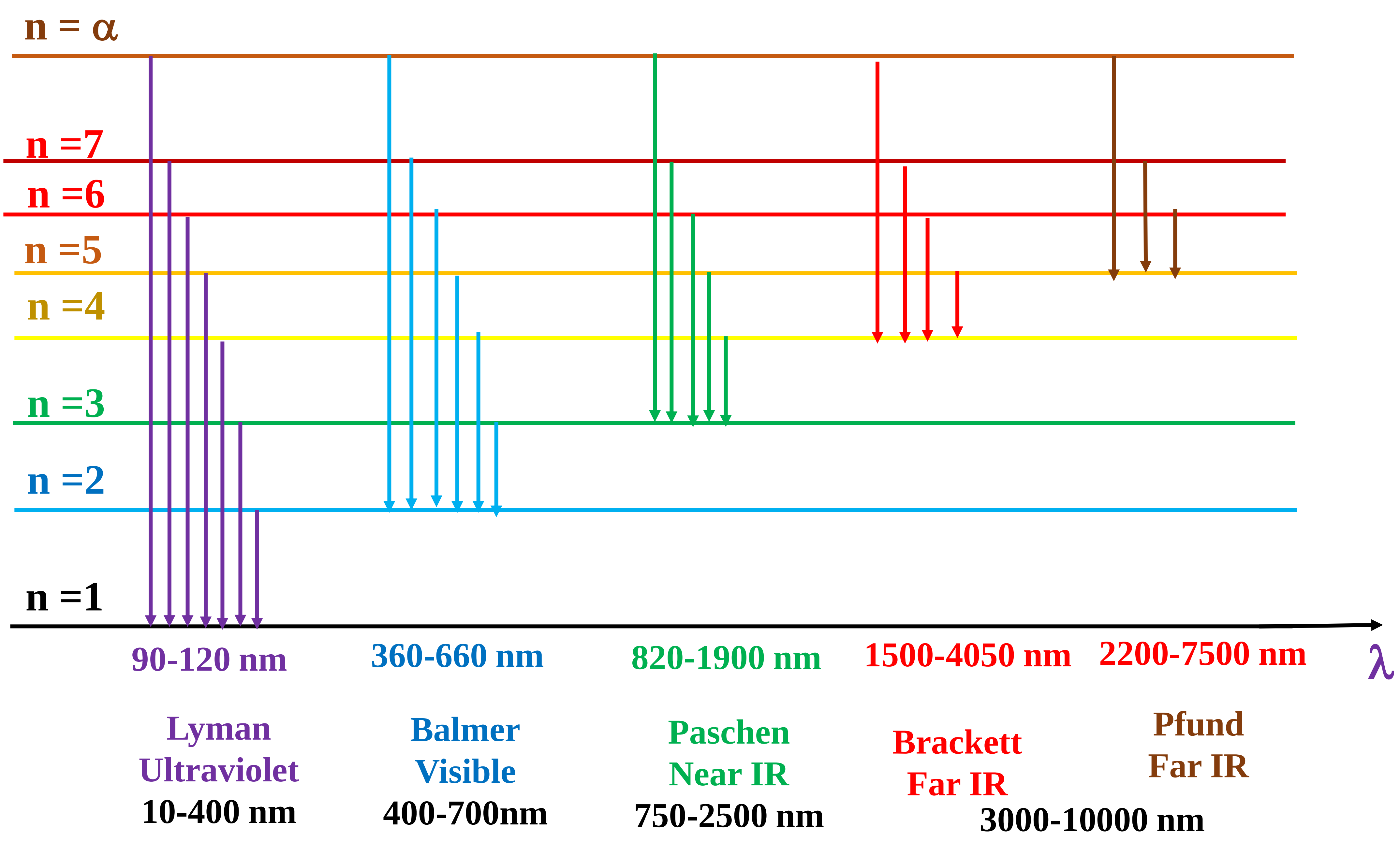
Balmer series



Visible / Ultraviolet



Excited state
Excited state
Excited state
Excited state
Excited state
Ground state
n=6
n=1
n=2
n=3
n=4
n=5
Near IR
Paschen series



n = α
n =7
n =6
n =5
n =4
n =3
n =2
n =1
90-120 nm
360-660 nm
820-1900 nm
1500-4050 nm
2200-7500 nm
λ
Lyman
Ultraviolet
10-400 nm
Balmer
Visible
400-700nm
Paschen
Near IR
750-2500 nm
Brackett
Far IR
Pfund
Far IR
3000-10000 nm

Series	Region of spectrum	Equation for wavenumber$(\bar{\nu})$
Lyman series	Ultraviolet	$\bar{\nu} = R_H \frac{1}{1^2} - \frac{1}{n^2}$, n =2,3,4,5...
Balmer series	Visible/ultraviolet	$\bar{\nu} = R_H \frac{1}{2^2} - \frac{1}{n^2}$, n =3,4,5,6..
Paschen series	Infrared	$\bar{\nu} = R_H \frac{1}{3^2} - \frac{1}{n^2}$, n =4,5,6,7...
Breckett series	Infrared	$\bar{\nu} = R_H \frac{1}{4^2} - \frac{1}{n^2}$, n =5,6,7,8...
Pfund series	Infrared	$\bar{\nu} = R_H \frac{1}{5^2} - \frac{1}{n^2}$, n =6,7,8...

$$\bar{\nu} = \frac{1}{\lambda} = \mathbf{R}\left[\frac{1}{(n_1)^2} - \frac{1}{(n_2)^2}\right]$$

Numerical problems

Problem 1 A wavelength of 4.653 μm is observed in a hydrogen spectrum for a transition that ends in the n_f =5 level. What was n_i for the initial level of the electron?

We have: $\frac{1}{\lambda} = R\left[\frac{1}{(n_1)^2} - \frac{1}{(n_2)^2}\right]$ where R = 109,677 cm^{-1}.

Wave length, 4.653 μm = 4.653×10^{-6} m = 4.653×10^{-4} cm

$$\frac{10^4}{4.653} = 109{,}677\ cm^{-1}\left[\frac{1}{(n_i)^2} - \frac{1}{(5)^2}\right]$$

Solving for, n_i the nearest integer is 4.

Problem 2. Calculate the wave lengths of first two lines of Balmer series in H spectrum?

We have: $\frac{1}{\lambda} = R\left[\frac{1}{(n_1)^2} - \frac{1}{(n_2)^2}\right]$ For Balmer series final level is n = 2.

Therefore, first line will be $\frac{1}{\lambda} = R\left[\frac{1}{(2)^2} - \frac{1}{(3)^2}\right]$ and

second line will be $\frac{1}{\lambda} = R\left[\frac{1}{(2)^2} - \frac{1}{(4)^2}\right]$

be $\frac{1}{\lambda} = R\left[\frac{9-4}{36}\right] = 109{,}677 \times \frac{5}{36} = 15233\ cm^{-1}$ or $\lambda = 6.565 \times 10^{-5}$ cm = 6565 ^{o}A

$\frac{1}{\lambda} = R\left[\frac{16-4}{64}\right] = 109{,}677 \times \frac{12}{64} = 20564\ cm^{-1}$ or $\lambda = 4.863 \times 10^{-5}$ cm = 4863 ^{o}A

Problem 3. Calculate the wave length of H_α and H_β lines of the Lyman series in H spectrum?

For Lyman series, $n_1 = 1$ and $n_2 = 2$ for H_α

$n_1 = 1$ and $n_2 = 3$ for H_β

Therefore, Wavelength of H_α is given by: $\frac{1}{\lambda} = R\left[\frac{1}{(1)^2} - \frac{1}{(2)^2}\right] = R\left[\frac{3}{4}\right]$

$$\lambda = \frac{4}{3R} = \frac{1.333}{1.09\times10^7} = 1.223\times10^{-7}\ \text{m} = 1223^{o}\text{A}$$

Wavelength of H_β is given by: $\frac{1}{\lambda} = R\left[\frac{1}{(1)^2} - \frac{1}{(3)^2}\right] = R\left[\frac{8}{9}\right]$

$$\lambda = \frac{9}{8R} = \frac{1.125}{1.09\times10^7} = 1.125\times10^{-7}\ \text{m} = 1125^{o}\text{A}$$

Thank you