

Hydrogen spectrum-definition, types, Practice problems,FAQs

Rick is a mad physicist who is obsessed with finding new things. One day, his close scientist friend Ben stumbles upon an element whose composition he doesn't know. In order to find this out, he calls up Rick and they both bring a certain piece of equipment called a "spectrometer". When the element was kept inside the spectrometer, the individual components were detected and measured. How ? In order to understand this, we must understand the concept of "atomic transition"; which denotes electrons jumping between different levels in an atom. In this process, photons (quantum of radiation) having energy and frequency are emitted. This helps to analyse the elements. Experiments conducted in the late 1800s proved that when gas molecules are excited and when subjected to an electric discharge, the light emitted is analysed using a diffraction grating. The resulting diffraction pattern consists of a series of lines called "spectrum". In other words, the spectrum consisted of a continuous band of light, but of individual lines with well-defined wavelengths. Hydrogen is no exception when it comes to showing a distinct spectrum due to atomic transitions. In this article, let's explore the hydrogen spectrum in detail.

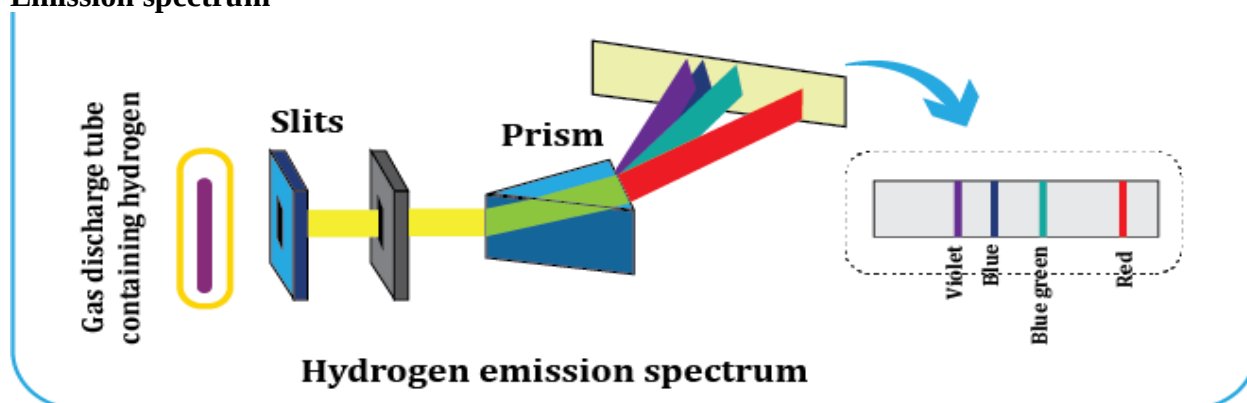
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What is the hydrogen spectrum?

When an electric discharge is applied to hydrogen, its molecules start to dissociate; as a result, they emit electromagnetic radiation. The hydrogen spectrum consists of bright lines in a dark background of discrete frequencies; these frequencies are named after the scientists who discovered them.

Emission spectrum



An emission spectrum is the series of lines which are emitted when an electron in the higher energy state transitions into a lower energy level thereby emitting energy in the forms of photons. It consists of bright lines in a dark background.

1) Continuous spectrum

A continuous spectrum contains all the wavelengths. They go from violet to red and when they lie in the visible region, can be seen by the naked eye.

2) Line spectrum

A line spectrum has discrete lines with dark spacing between them and does not consist of all the wavelengths like the continuous spectrum.



Fig showing emission spectrum

Absorption spectrum

An absorption spectrum is the series of lines which are emitted when an electron in the lower energy level transitions into a higher energy level by absorbing energy in the forms of photons in the process. It consists of dark lines in a bright background. The dark patches are the light waves of those wavelengths which are absorbed by the electrons.

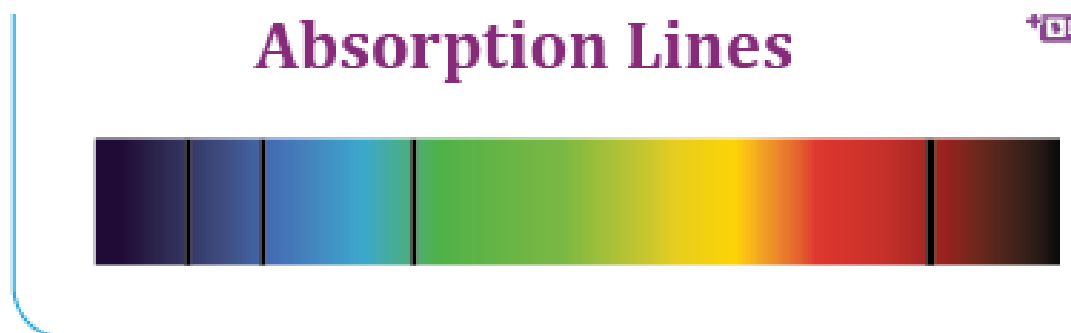


Fig showing absorption spectrum

Energy levels of an atom

The total energy of an electron in the n^{th} orbit is given by,

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

$R = 1.097 \times 10^7 \text{ m}^{-1}$ is called the Rydberg's constant.

$h = 6.626 \times 10^{-34} \text{ Js}$ denotes Planck's constant.

$c = 3 \times 10^8 \frac{\text{m}}{\text{s}}$ denotes the speed of light.

Z- atomic number of the element.

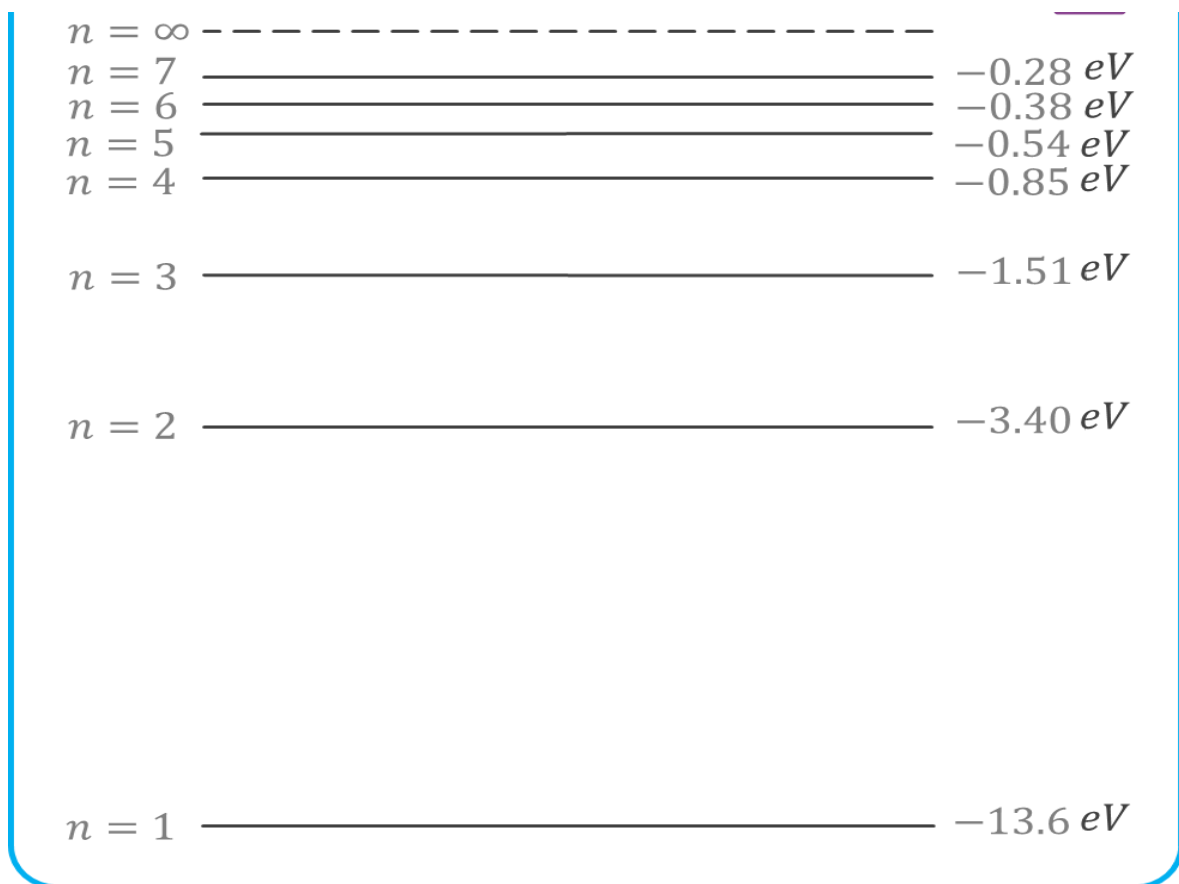


Fig showing different energy levels of an atom

Substituting the above values, we get $E_n = -\frac{13.6 Z^2}{n^2} eV$

eV stands for electron volt. $1 eV = 1.6 \times 10^{-19} J$.

For an electron in the ground state of a hydrogen atom, $Z=1, n=1$

$$E_1 = -\frac{13.6 (1)^2}{(1)^2} = -13.6 eV$$

In the first excited state, $n=2$

$$E_2 = -\frac{13.6 (1)^2}{(2)^2} = -3.40 eV$$

In the second excited state, $n=3$

$$E_3 = -\frac{13.6 (1)^2}{(3)^2} = -1.51 eV$$

As the value of n increases, the energy of the electron increases.

Spectral lines of hydrogen and Rydberg's formula

Hydrogen has its own emission spectrum caused by atomic transitions. Let n_f be the final level and n_i be the initial level from where the electron jumps, then the wavelength of the emitted photon can be calculated as

$$\frac{1}{\lambda} = R \left[\frac{1}{n_f^2} - \frac{1}{n_i^2} \right]$$

Here, R= Rydberg constant

$$R = 1097 \times 10^4 \text{ m}^{-1}$$

Note:

$\frac{1}{\lambda} = \underline{\nu}$ is called the wave number.

1) For Lyman series $n_f = 1$, $n_i = 2, 3, 4 \dots n$

$$\frac{1}{\lambda} = R \left[\frac{1}{1^2} - \frac{1}{n^2} \right]$$

The Lyman series lies in the Ultraviolet region of the EM spectrum.

2) For Balmer series $n_f = 2$, $n_i = 3, 4, 5 \dots n$

$$\frac{1}{\lambda} = R \left[\frac{1}{2^2} - \frac{1}{n^2} \right]$$

It lies in the visible region of the EM spectrum.

3) For Paschen series $n_f = 3$, $n_i = 4, 5, 6 \dots n$

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

It lies in the infrared region of the EM spectrum.

4) For Brackett series $n_f = 4$, $n_i = 5, 6, 7 \dots n$

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

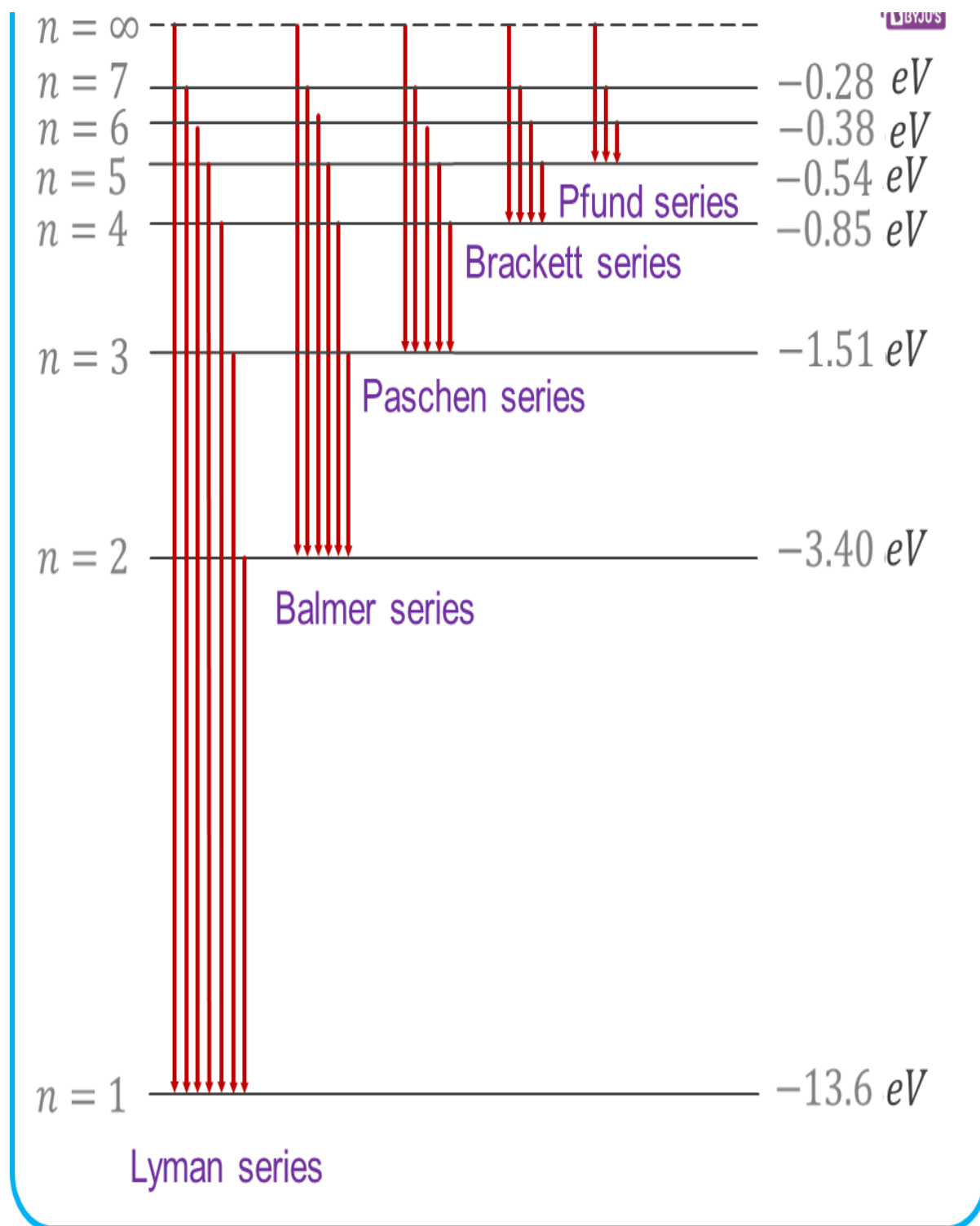
It lies in the infrared region of the EM spectrum.

5) For Pfund series $n_f = 5$, $n_i = 6, 7, 8 \dots n$

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

It lies in the infrared region of the EM spectrum.

For the series limit of any series, $n_i = \infty$.



Practice problems

Q1. Calculate the largest and shortest wavelengths in the Brackett series for hydrogen ?

For the Brackett series,

$$\frac{1}{\lambda} = R \left[\frac{1}{n_f^2} - \frac{1}{n_i^2} \right]$$

$$n_f = 4, n_i = 5, 6, \dots$$

Let $\lambda_{longest}$ indicate the largest wavelength and shortest $\lambda_{longest}$ indicate the shortest wavelength. Then

For the shortest wavelength, wave number should be maximum $\Rightarrow n_f = 4, n_i = \infty$

$$\frac{1}{\lambda_{shortest}} = R \left(\frac{1}{4^2} - \frac{1}{\infty} \right) = R \left(\frac{1}{16} - \frac{1}{\infty} \right)$$

$$\lambda_{shortest} = 1.4585 \times 10^{-6} \text{ m} = 14585 \text{ \AA}$$

Let $\lambda_{longest}$ indicate the maximum wavelength. For the longest wavelength, wave number should be minimum $n_f = 4, n_i = 5$

$$\frac{1}{\lambda_{longest}} = R \left(\frac{1}{4^2} - \frac{1}{5^2} \right) = R \left(\frac{1}{16} - \frac{1}{25} \right)$$

$$\lambda_{longest} = 40.650 \times 10^{-7} \text{ m} = 40560 \text{ \AA}$$

Q2. Calculate the ratio of the largest to the shortest wavelengths in the Paschen series of hydrogen.

(a) $\frac{25}{9}$ (b) $\frac{17}{6}$ (c) $\frac{16}{7}$ (d) $\frac{17}{6}$

Answer. c

For the Paschen series, $n_f = 3, n_i = 4, 5, \dots \infty$

Let $\lambda_{longest}$ indicate the largest wavelength and shortest $\lambda_{longest}$ indicate the shortest wavelength. Then

For the shortest wavelength, $n_f = 3, n_i = \infty$

$$\frac{1}{\lambda_{shortest}} = R \left(\frac{1}{3^2} - \frac{1}{\infty} \right) = R \left(\frac{1}{9} - \frac{1}{\infty} \right)$$

For the longest wavelength, $n_f = 3, n_i = 4$

$$\frac{1}{\lambda_{longest}} = R \left(\frac{1}{3^2} - \frac{1}{4^2} \right) = R \left(\frac{1}{9} - \frac{1}{16} \right)$$

$$\frac{\lambda_{longest}}{\lambda_{shortest}} = \frac{\left(\frac{1}{9} - \frac{1}{\infty} \right)}{\left(\frac{1}{9} - \frac{1}{16} \right)} = \frac{16}{7}$$

Q3. Find the ratio of the longest wavelength corresponding to Lyman and the shortest wavelength in the Balmer series?

- (a) $\frac{1}{3}$ (b) $\frac{5}{9}$ (c) $\frac{1}{4}$ (d) $\frac{5}{27}$

Answer. a

For the longest wavelength of the Lyman series λ_{Lyman} , we have $n_f = 1$, $n_i = 2$

For the shortest wavelength of the Balmer series, (λ_{Balmer}) we have $n_f = 2$, $n_i = \infty$

$$\frac{1}{\lambda_{Lyman}} = R \left(\frac{1}{1^2} - \frac{1}{2^2} \right) = \frac{3R}{4}; \frac{1}{\lambda_{Balmer}} = R \left(\frac{1}{2^2} - \frac{1}{\infty^2} \right) = \frac{R}{4}$$

$$\frac{\lambda_{Lyman}}{\lambda_{Balmer}} = \frac{R}{4} \times \frac{4}{3R} = \frac{1}{3}$$

Q4. Calculate the wave number of the last line of the Paschen series in the hydrogen spectrum. Given, the value of Rydberg's constant to be $10^7 m^{-1}$.

- (a) $0.0486 \times 10^7 m^{-1}$ (b) $1.11 \times 10^6 m^{-1}$ (c) $2.5 \times 10^7 m^{-1}$ (d) $0.025 \times 10^4 m^{-1}$

Answer. b

For the last line of the Paschen series,

$$n_f = 3, n_i = \infty$$

$$\text{Wave number, } \underline{\nu} = \frac{1}{\lambda_{paschen}} = R \left(\frac{1}{3^2} - \frac{1}{\infty^2} \right) = \frac{R}{9} = \frac{10^7}{9} = 1.11 \times 10^6 m^{-1}$$

FAQs

Q1. What are the characteristics of the hydrogen spectrum?

Answer. A sample of hydrogen, when excited, emits a characteristic light even in the visible region. When the light is split into its component wavelengths, violet, blue, red and green lines are observed.

Q2. What is the importance of the hydrogen spectrum?

Answer. Hydrogen spectrum is important evidence to show that quantisation of charges is necessary. It also helps to explain why the gas glows when an electric discharge occurs.

Q3. What are the 3 main regions in which the hydrogen spectrum lies?

Answer. Hydrogen spectrum mainly lies in the UV, visible or Infrared region. The Lyman series lies in the UV region, the Balmer series lies in the visible region, while Paschen, Brackett and Pfund lie in the infrared region.

Q4. What is the difference between emission and absorption spectrum?

Answer. Absorption spectrum is mainly exhibited by gases while on the other hand, emission spectrum is mainly shown by solids and liquids. Emission spectrum contains bright lines in a dark background while the absorption spectrum contains dark lines in a bright background.