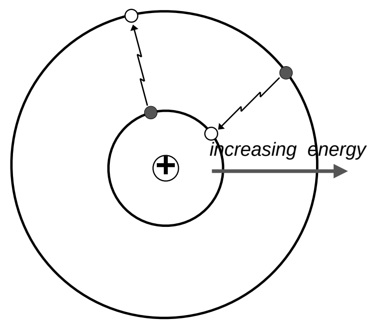
THE ATOMIC SPECTRUM OF HYDROGEN

Types of spectra

Continuous	electromagnetic spectrum	whole range of frequencies / wavelengths
Line	atomic spectra	definite frequencies producing sharp lines

Origin

When an electron changes energy levels, light of a particular frequency is emitted if the electron drops from a higher to a lower level or is adsorbed if an electron is promoted to a higher level



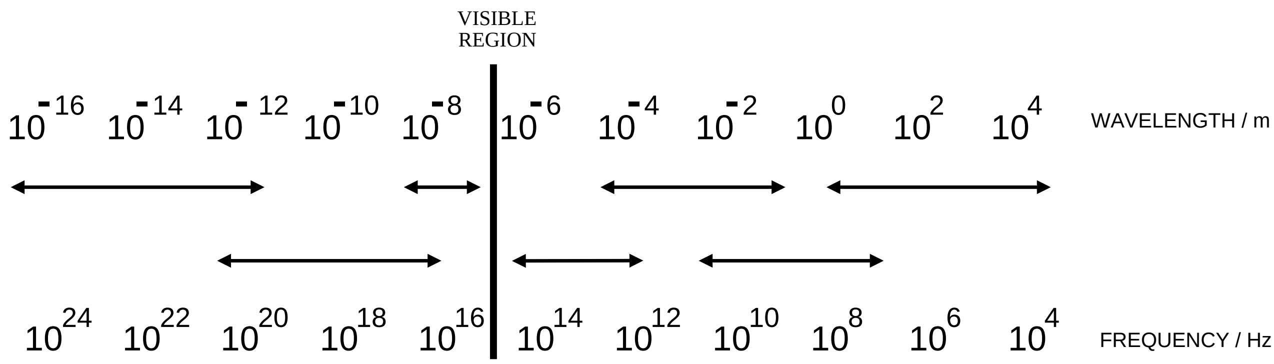
Energy, frequency and wavelength are linked

$$E = \frac{hc}{\lambda} \qquad E = h\upsilon \qquad \Delta E = E_{init} - E_{final}$$

Q.1 *What do the symbols represent?*

E	*h*
c	λ

The electromagnetic spectrum



Q.2 *Label the regions of the spectrum*

Atomic spectra of hydrogen

Absorption The absorption spectrum occurs when the electron in the lowest energy level (**ground state**) is provided with energy to lift it to higher energy levels.

Hydrogen was investigated because it had only one electron in its atom.

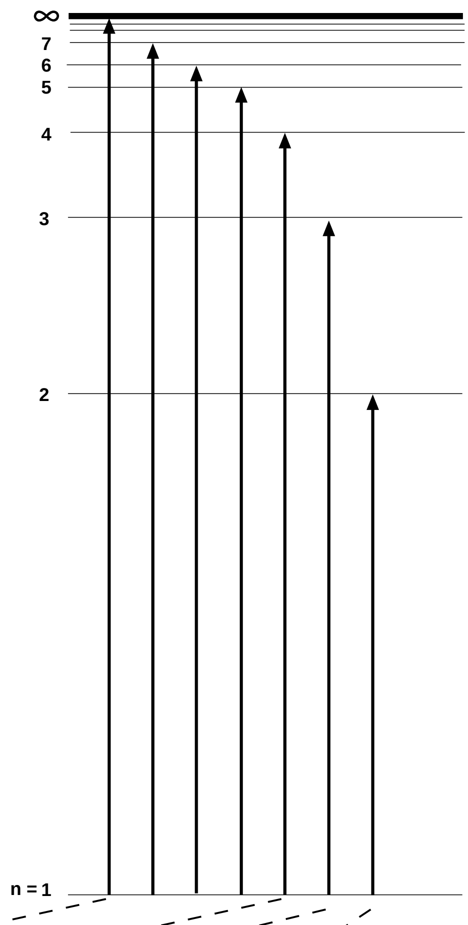
Because only selected frequencies were absorbed there must only be a selected number of possible transitions.

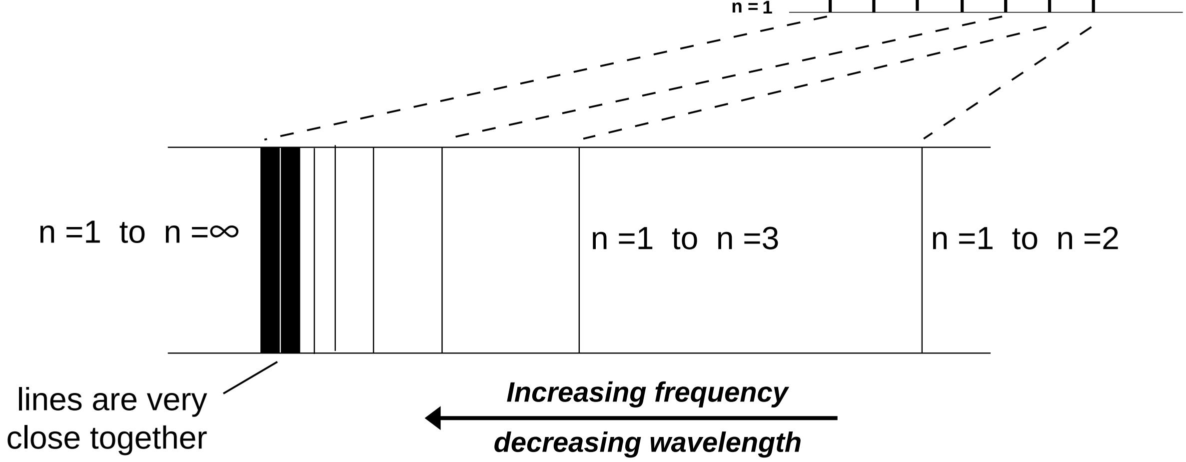
By relating the frequency absorbed one can measure the energy associated with it (ENERGY = PLANCK'S CONSTANT x FREQUENCY).

Because the frequency values converge to the higher frequency end of the spectrum it shows that the energy levels also converged. If sufficient energy is given to the electron it can escape from the atom and ionisation occurs.

Points

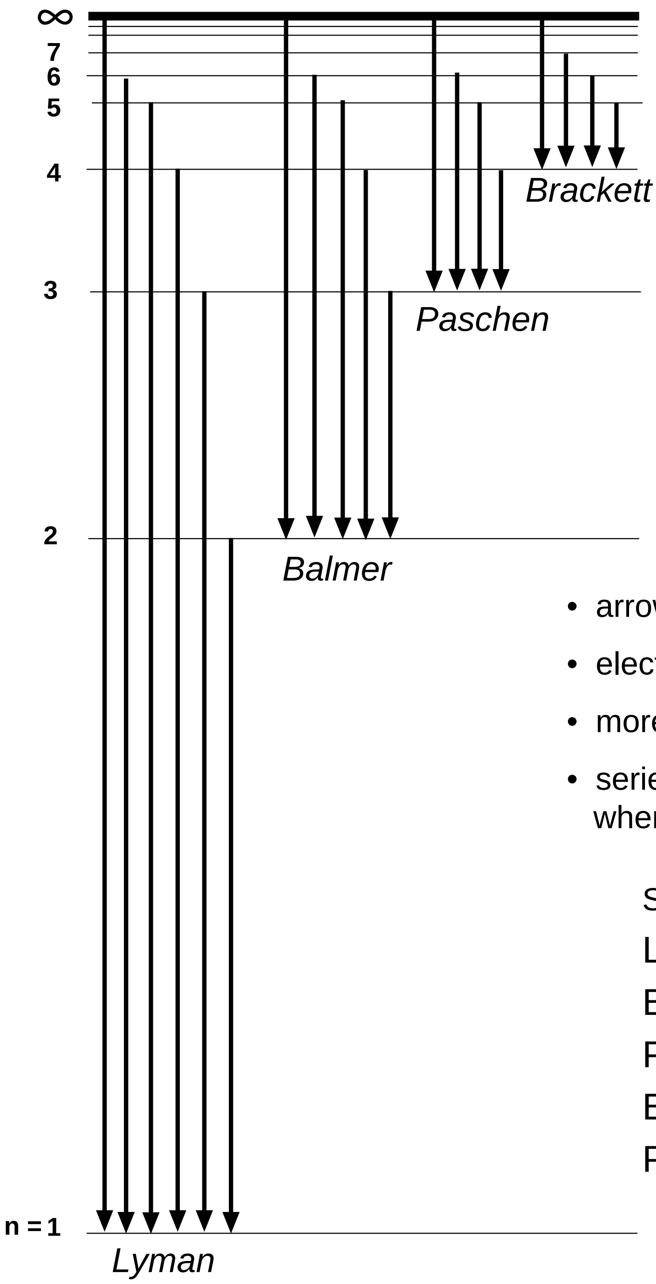
- principal energy levels labelled 1,2,3 etc
- higher energies are further from the nucleus
- energy of higher levels is greater
- hydrogen's single electron is in 1st level
- given energy it can go to any higher level
- only a fixed number of transitions are possible
- each energy change is shown in the spectrum
- a series of black lines appear in the spectrum
- the frequency corresponds to the energy
- lines converge to high frequency end
- the final transition corresponds to ionisation





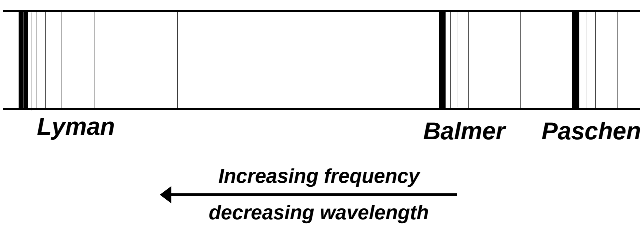
Emission Emission spectra arise when electrons, having been excited to higher energy levels, return to lower ones and give out energy.

Balmer, was the first to notice this effect and gave his name to the spectral series resulting from transitions back to the second energy level. This was because it was in the visible region.



- arrows in the diagram point downwards
- electrons are going from high to low energy
- more complicated than the absorption spectrum
- series are arranged and named according to where the transition ends.

Series	Ends on	Region
Lyman	n=1	UV
Balmer	n=2	VISIBLE
Paschen	n=3	IR
Brackett	n=4	IR
Pfund	n=5	IR



Emission spectrum