# Quantised Energy Levels in Atoms

- **describe the quantised states of the atom in terms of electrons forming standing waves, recognising this as evidence of the dual nature of matter;**
- **explain the production of atomic absorption and emission spectra, including those from metal vapour lamps;**
- **interpret spectra and calculate the energy of photons absorbed or emitted, $\Delta E = hf$;**
- **analyse the absorption of photons by atoms in terms of**

    **– the particle-like nature of matter**

    **– the change in energy levels of the atom due to electrons changing state**

    **– the frequency and wavelength of emitted photons, $E = hf = hc/\lambda$**

    **(not including the bombardment of atoms by electrons);**

### Electrons in atoms

Niels Bohr proposed a simple model of the atom in the early 20th Century. He suggested that the electrons were in circular orbits around the positive nucleus, but that the orbits could only be certain sizes. (often referred to as Bohring) This neatly explained the line spectrum observed for excited atoms. Light is given off from the electrons jumping from one orbit to another. If the orbits could be any size, the spectrum would be continuous, as all energies of jump are possible.

This quantisation of the orbital energies does not really fit with the concept of electrons as simple particles. The resolution of the fixed energies of atomic electrons had to wait until *quantum mechanics* was developed. It now seems that electrons in atoms are behaving more like waves than particles.

The model for the atom that Rutherford proposed in 1911, that the atom consisted of a small dense, positively charged nucleus surrounded by a cloud of electrons, has a weakness because the accelerating electrons should radiate energy and spiral into the nucleus.
In 1913, Niels Bohr, said that the electrons should not be considered to be orbiting like planets. He said that they simply existed outside the nucleus with certain amounts of energy. According to Bohr, the electrons in the atom existed in certain discrete ENERGY LEVELS.

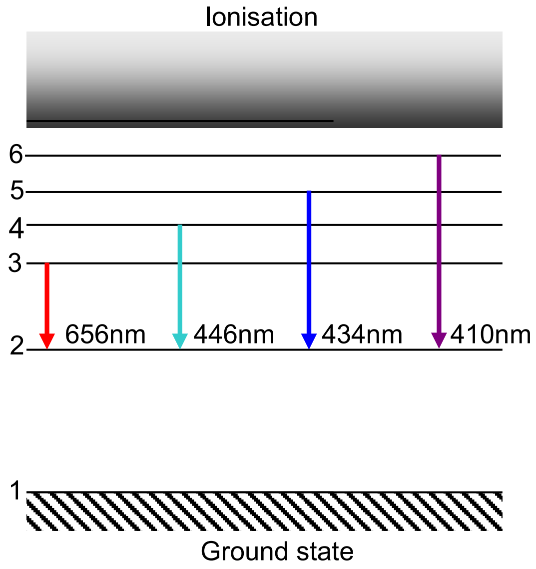
- Each element has certain allowed energy levels that are unique to that element.
- Electrons can only exist in one of these allowable energy levels, not in between. i.e. energy levels are quantised.
- If an electron is given extra energy it can move up to a higher energy level by absorbing an amount of energy equal to the difference between the energy levels.
- When an electron in a higher energy level returns to its normal (ground state) energy level, it emits the energy in the form of a photon. The energy of the photon (E = hf) is equal to the difference in energy levels the electron moves between.

## Energy levels in hydrogen

The ionisation energy for hydrogen is 13.6eV. The ground state energy, level = 0eV,
level 2 = 10.2eV, 12.1eV, 12.7eV, 13.1ev, 13.2eV ………..13.6eV.

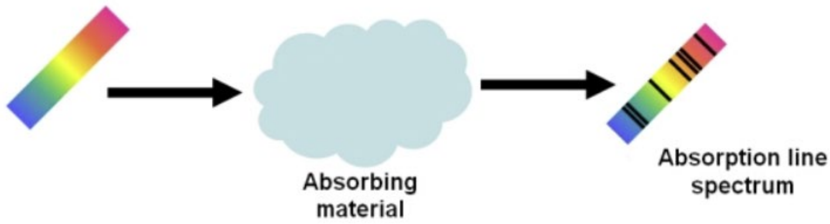
Bohr found that $E_n = E_I - \frac{E_i}{n^2}$

Where $E_n$ = the energy associated with a particular energy level of hydrogen
$E_I$ = ionisation energy, which is 13.6eV for hydrogen
n = any whole number, ie, 1, 2, 3 …



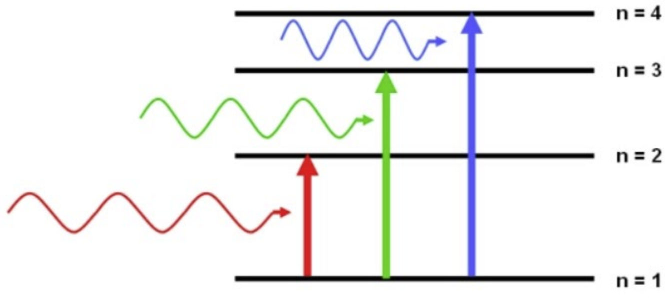

## Absorption spectra

An absorption line will appear in a spectrum if an absorbing material is placed between a source and the observer. This material could be a cloud of interstellar gas or a cloud of dust.



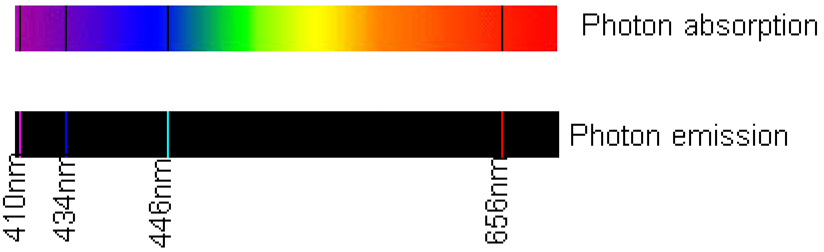

Incoming light (left) passes through a cloud of absorbing material, such as a cloud of interstellar gas. The light that leaves the cloud (right) shows absorption lines in the spectrum at discrete frequencies.

According to quantum mechanics, an atom, element or molecule can absorb photons with energies equal to the difference between two energy states.

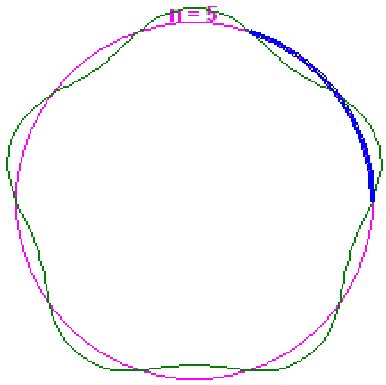



Photons with specific energies will be absorbed by an atom, ion or molecule if this energy is equal to the difference between the energy levels. In this example, three different photon energies are required to promote an electron from the ground state (n = 1) to an excited state (n = 2, 3 and 4).

Photons can be emitted or absorbed. Below are the absorption and emission lines for hydrogen.




**De Broglie's explanation of energy levels**




De Bröglie suggested that electrons have wave properties such as wavelength, and that the orbits (energy levels) that could exist were those where the wavelength of the electron set up a stable standing wave. This is consistent with the quantisation of energy levels, because standing waves have quantised wavelengths.
De Bröglie said that, in a similar way, the wavelength of the electrons orbiting the nucleus must 'fit' into the circumference of the orbit exactly. This will only happen with particular wavelengths and, therefore, energies and explains why energy levels are quantised.
Electrons with wavelengths that do not set up standing waves destructively interfere with themselves and cancel out.

Electrons orbiting a nucleus can be modelled as circular standing waves, therefore the electron is exhibiting wave like properties. The standing wave will exist only if the circumference of its orbit corresponds to a whole number of wavelengths. i.e. $2\pi r = n\lambda$, where $\lambda = \frac{h}{mv}$

This can be rewritten as $mvr = \frac{nh}{2\pi}$, where n is a whole number.
Therefore only specific values of wavelength are permitted. The momentum of the electron is related to its wavelength, and the energy of the electron is related to its momentum. Therefore the electron's energy is quantised, so only certain energy levels and orbits are possible.