A2 Level Physics

Chapter 5 – Waves and Particle Nature of Light

5.10.1 Energy Levels in Atoms (Edexcel only)

Notes

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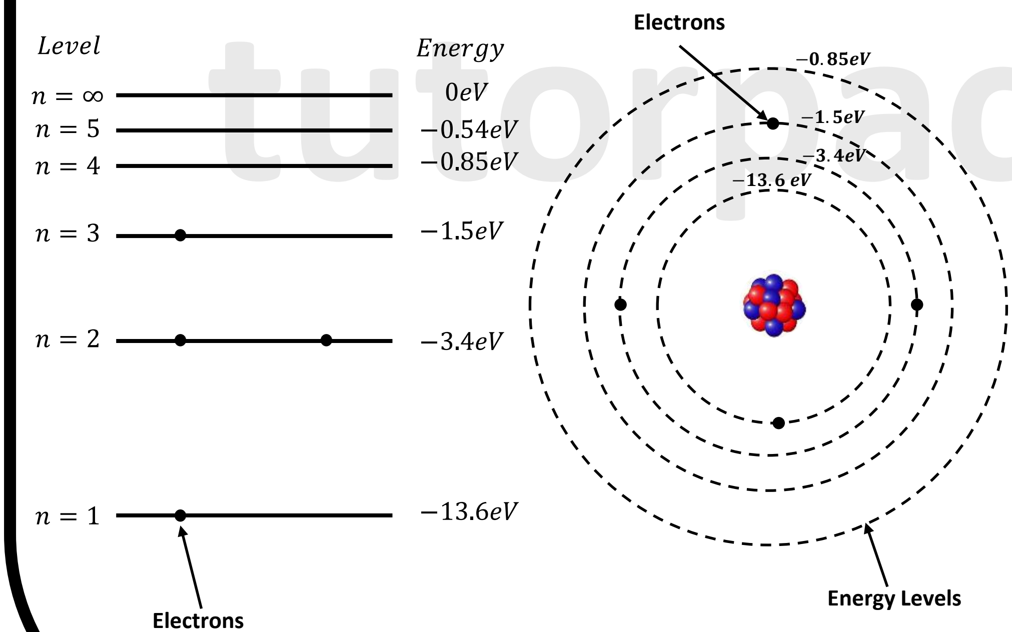
Energy Levels of Electrons

In an atom, electrons can only occupy a number of discrete energy levels. The word discrete refers to separate or distinct.

The diagram below shows a typical arrangement of energy levels and comparing it to a typical atom with energy levels. The vertical axis of the diagram on the left is a scale of electron energy, E. The horizontal lines indicate the different permitted energy levels. Each level is assigned a number, with $n = 1$ representing the ground state. The dots represent electrons occupying the different levels.

It's worth noting that the energy scale normally starts at the top with $E = 0$.

Energy can be expressed in J or eV.



Energy Levels of Electrons

When an electron gains energy or is excited, it goes from one level to another. It can only do so by absorbing a photon with an energy equal to the difference between the two levels; for example, to raise an electron from $n = 1$ to $n = 2$, you'd require $10.2\ eV$ $(13.6\ eV - 3.4\ eV)$; there cannot be any energy left over.

However when the electron is excited or gains energy it doesn't like being in an energy level above where its supposed to be usually. As a result, once it has absorbed the energy it will want to immediately return to its original energy level. When an electron loses energy or is de-excited, it falls, from one level to another. It does so by emitting a single photon whose energy corresponds exactly to the difference between the two levels and has a specific wavelength. For example, an electron excited to $n = 2$ will de-excite and return to $n = 1$, releasing a $10.2\ eV$ photon. The equation below can be used to show a transition between the two energy levels:

$$\Delta E = E_1 - E_2 = hf = \frac{hc}{\lambda}$$

Where:

E = change in energy in J

E_1 = energy of initial energy level in J

E_2 = energy of final energy level in J

hf = photon energy

h = the Planck constant $(6.63 \times 10^{-34}\ Js)$

f = frequency in Hz

c = speed of light in a vacuum $(3 \times 10^{8}\ ms^{-1})$

λ = wavelength in m

Energy Levels of Electrons

It's important to remember that electrons can't have an energy value that's halfway between two levels. The electron will be pulled back to its original level if it does not have enough energy to go to the next level.

The values of all energy levels are negative, with the ground state having the largest negative value. The energy of an electron that is completely free from an atom is 0. This negative sign is used to represent the amount of energy necessary to remove an electron from an atom.

The ground state represents the most stable configuration. When the electrons in an atom occupy the lowest possible energy level, the atom is in the most stable state.

Ionisation

When an electron is removed from an atom, it becomes **ionised**. Each energy level in an atom represents the energy required to remove an electron from that level. **Ionisation energy is the energy required to remove an electron from the ground state of an atom**.

Example:

The ionisation energy of an unexcited hydrogen atom is 13.6 eV – this is the energy required to remove an electron from the $n = 1$ energy level.

Photon emission – fluorescent tubes

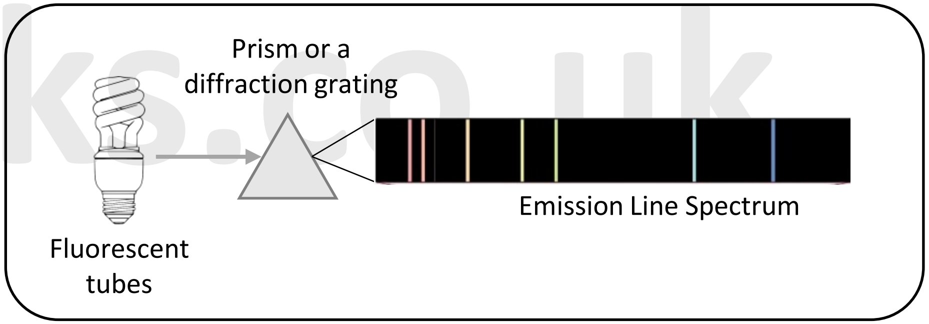
- Fluorescent tubes work by electron excitation and photon emission.
- They contain mercury vapor and are subjected to high voltage, which accelerates free electrons.
- These electrons ionise some mercury atoms, leading to the production of more free electrons.
- When free electrons collide with mercury atoms, they excite the atomic mercury electrons to a higher energy level.
- Excited electrons returning to their ground state release energy as high-energy photons within the UV range.
- The emitted photons have various energies and wavelengths, matching the different electron transitions.
- The tube's inner surface has a phosphorus coating that absorbs these UV photons.
- The phosphorus electrons are excited to higher energy levels and then cascade down to lower levels.
- As phosphorus electrons return to lower energy levels, they emit visible light photons with lower energy.

Low-energy bulbs made of fluorescent tubes.

Line emission spectra

When you use a prism or a diffraction grating* to disperse the light from a fluorescent tube, you observe a line spectrum. These devices separate light into its different wavelengths by bending each wavelength at a distinct angle. A diffraction grating, in comparison to a prism, yields sharper and more distinct spectral lines. This results in a line emission spectrum, where each bright line on a dark background represents a specific wavelength emitted by the source.

Line spectra serve as proof that electrons within atoms exist in distinct energy levels. Atoms emit photons whose energies equal to the difference between two energy levels. Since only specific photon energies are allowed, the line spectrum exclusively displays these associated wavelengths.

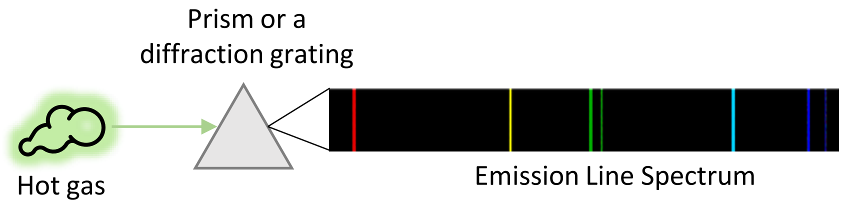


Line emission spectrum of fluorescent tube light.

*A CD functions as a diffraction grating due to its etched lines of dots, creating a rainbow when reflecting white light by splitting it into its component wavelengths.

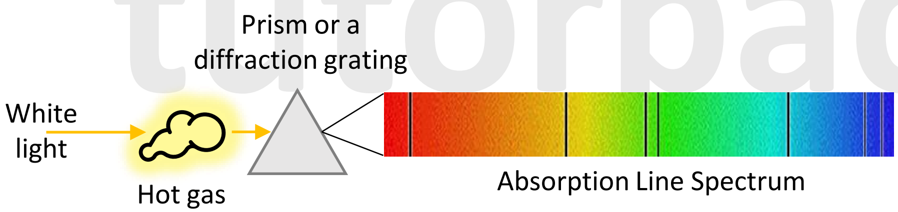
Spectra

When a gas is heated, it begins to glow. The gas produces light as it glows. If you take this light and pass it through a diffraction grating or prism, a spectrum is formed. The spectrum is made up of several discrete lines, indicating the presence of only certain wavelengths.

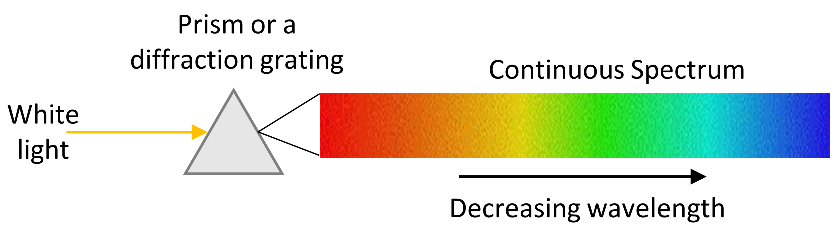


This is an emission line spectrum, rightfully called because it arises from the light emitted by a material.

When white light passes through a heated gas, it produces black lines in its spectrum. This absorption line spectrum shows that certain wavelengths have been absorbed.



If you split white light up with a prism, the colours all blend together, leaving no gaps in the spectrum. This simply demonstrates that white light has a continuous spectrum.



Spectra

Explaining Spectra

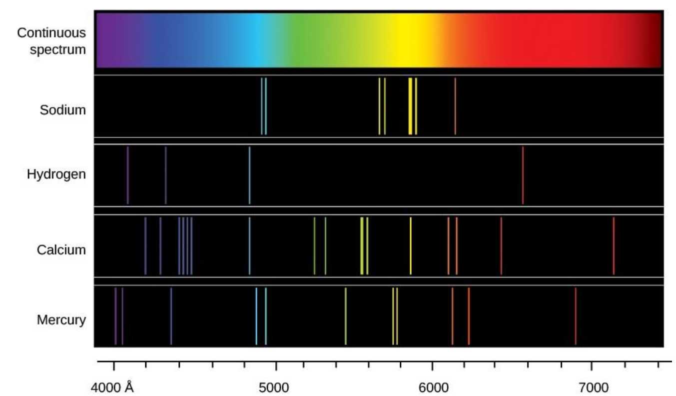
When an electron is de-excited, energy is released in the form of a photon of a specified wavelength. The photon's energy is calculated using $E = hc/\lambda$, where h is the Planck constant, c is the speed of light, and λ is the wavelength of the photon. The energy released is the difference between the initial energy level of the electron, and the final energy level of the photon. The transitions between the different energy levels emit photons of different wavelengths.

The wavelengths of light produced by the de-exciting of electrons varies for each element since each has its own set of discrete energy levels. When atoms in a gas are excited, spectroscopy is used to identify elements based on the wavelength of light emitted.

In an absorption spectrum, electrons jump up from lower levels to higher levels, absorbing photons in the process.

After measuring the wavelengths in a line spectrum, it is possible to work out the energy level diagram of the atoms producing the spectrum.

Different atoms have different spectral lines, which can be used to identify elements within stars.



Please see '5.10.2 Energy Levels in Atoms worked examples' pack for exam style questions.

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