Key Stage 5

Spectroscopy

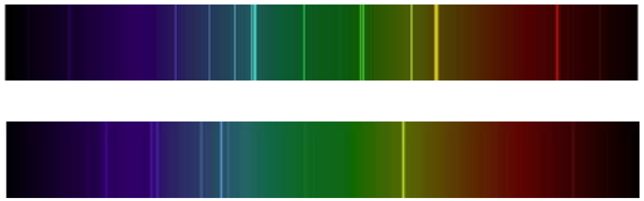
Student worksheet

Emission spectra

If you pass electricity through a tube filled with hydrogen gas, the gas gains energy and glows. This can be done with a wide range of other elements like helium, neon, vaporised sodium and carbon. The colour that you see with your eyes usually looks different for different elements.

If you look more closely at the light emitted by elements by using a spectroscope or diffraction grating you will see a characteristic pattern of lines. This set of lines is known as the ***Emission Spectrum*** for that element.

These patterns of lines at specific wavelengths act a bit like a fingerprint as no two elements have exactly the same pattern of lines.



carbon

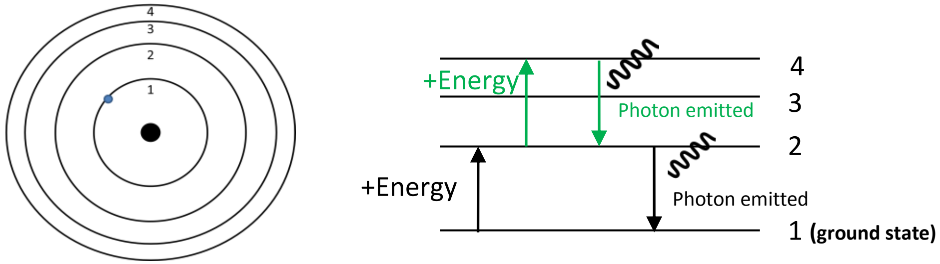
helium

Emission Spectra for carbon and helium showing distinctive lines

https://www.oxfordsparks.ox.ac.uk/content/what-are-quantum-rainbows

How are emission spectra formed?

A simple model of an atom has a central nucleus surrounded by electrons in orbits.



Simple model of a hydrogen atom with diagram just showing the electron orbits and transitions - orbits are actually much more complicated than this

Emission spectra are created when electrons gain energy and can move to a higher energy level. In this simplified model the electron moves to a further out orbit e.g., from level 1, the **ground state**, to level 2 or from level 2 to level 4.

This higher energy level is less stable and the electron ‘falls back’ to a more stable lower energy level, and in doing so gives out energy in the form of light (not necessarily in the visible spectrum).

These energy levels are different for different elements so the ‘colours’ of the light given out varies from element to element, for instance the bright orange light given out by sodium street lights, the red colour of neon tubes etc. This equates to a different pattern of lines in the emission spectra of different elements.

In the case of hydrogen, transitions from higher energy levels to the ground state (level 1) involve the largest difference in energy and the light emitted is in the ultraviolet range of the electromagnetic spectrum. Those between the higher energy levels 6,5,4,3 and the level 2 state produce lines in the visible spectrum. The lines in the visible part of the spectrum are known as ***the Balmer series.***

Hydrogen emission spectrum showing lines in visible region of the electromagnetic spectrum (Balmer series)

Calculations – Energy in Joules

The difference in energy between two different levels of a sodium atom is 3.36×10^{-19}J. What wavelength of light is emitted when an electron drops back between the two levels.

Using the two relationships **E=hf** and **c=fλ** we can calculate the wavelength emitted.

h (Planck's constant)=6.6×10^{-34} ; **E**= 3.36×10^{-19}J;
c (speed of light in a vacuum)=$3 \times 10^{8}\,ms^{-1}$

Rearranging **E=hf** we get

f=E/h

c=fλ substituting for f we get **c=Eλ/h**

Rearranging gives **λ=hc/E**

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What is the wavelength?

Calculations – Energy in Volts

Sometimes energy levels are expressed in Volts.

What is the wavelength emitted if a transition occurs between levels that differ by 3V?

Use **E=qV** (q= electron charge = 1.6×10^{-19}C) and the equations from the previous example.