Episode 501: Spectra and energy levels

Summary

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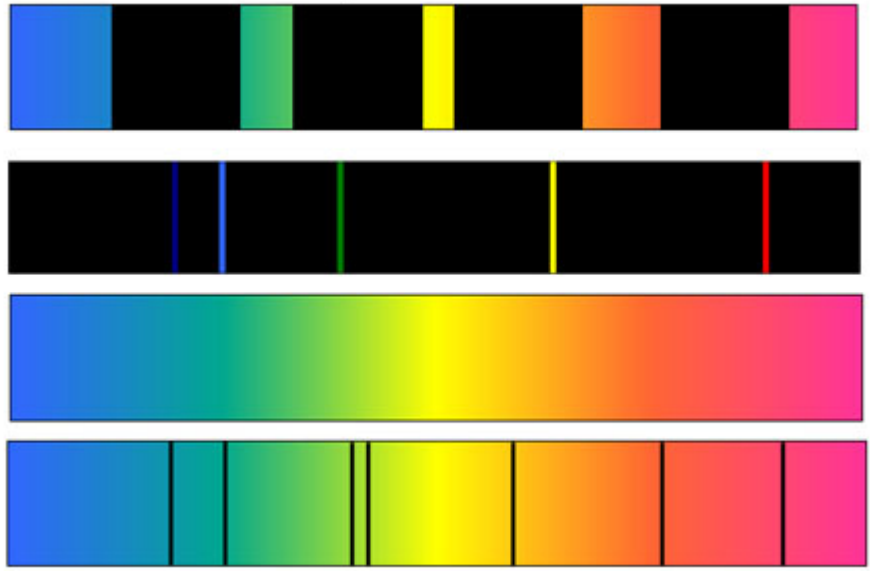
Demonstration: Looking at emission spectra
Show a white light and a set of standard discharge lamps: sodium, neon, hydrogen and helium. Allow students to look at the spectrum of each gas. They can do this using a direct vision spectroscope or a bench spectroscope, or simply by holding a diffraction grating up to their eye.

What is the difference? (The white light shows a continuous spectrum; the gas discharge lamps show line spectra.)



(Diagram: resourcefulphysics.org)

Emission and absorption spectra

The spectrum of a gas gives a kind of 'finger print' of an atom. You could relate this to the simple flame tests that students will have used at pre-16 level. Astronomers examine the light of distant stars and galaxies to discover their composition (and a lot else).

Discussion: The meaning of quantisation

Relate the appearance of the spectra to the energy levels within the atoms of the gas. Students will already have a picture of the atom with negatively charged electrons in orbit round a central positively charged nucleus. Explain that, in the classical model, an orbiting electron would radiate energy and spiral in towards the nucleus, resulting in the catastrophic collapse of the atom.

This must be replaced by the Bohr atomic structure – orbits are quantised. The electron's energy levels are discrete. An electron can only move directly between such levels, emitting or absorbing individual photons as it does so. The ground state is the condition of lowest energy – most electrons are in this state.

Think about a bookcase with adjustable shelves. The bookshelves are quantised – only certain positions are allowed. Different arrangement of the shelves represents different energy level structures for different atoms. The books represent the electrons, added to the lowest shelf first etc

Demonstration: Illustrating quantisation

Throw a handful of polystyrene balls round the lab and see where they settle. The different levels on which they end up – the floor, on a desk, on a shelf – gives a very simple idea of energy levels.

Some useful clipart can be found below:

Episode 501-1: The emission of light from an atom (Word, 35 KB)

Resourceful Physics >Teachers>OHT>Emission of Light

An energy input raises the electrons to higher energy levels. This energy input can be by either electrical, heat, radiation or particle collision.

When the electrons fall back to a lower level there is an energy output. This occurs by the emission of a quantum of radiation.

Discussion: Energy levels in a hydrogen atom
Show a scale diagram of energy levels. It is most important that this diagram is to scale to emphasise the large energy drops between certain levels.

The students may well ask the question, “Why do the states have negative energy?” This is because the zero of energy is considered to be that of a free electron 'just outside' the atom. All energy states 'below' this – i.e. within the atom are therefore negative. Energy must be put into the atom to raise the electron to the 'surface' of the atom and allow it to escape.

Worked example and student questions:

Calculating frequencies
Calculate the frequency and wavelength of the quantum of radiation (photon) emitted due to a transition between two energy levels. (Use two levels from the diagram for the hydrogen atom.)

$E_2 - E_1 = hf$

Point out that this equation links a particle property (energy) with a wave property (frequency).

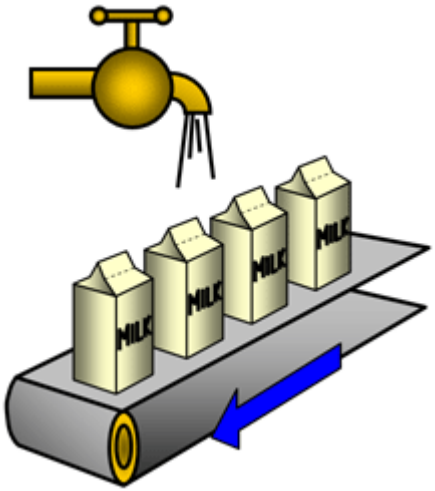
Ask your students to calculate the photon energy and frequency for one or two other transitions. Can they identify the colour or region of the spectrum of this light?

Emphasise the need to work in SI units. The wavelength is expressed in metres, the frequency in hertz, and the energy difference in joules. You may wish to show how to convert between joules and electronvolts.

Discussion: Distinguishing quantisation and continuity
The difference between the quantum theory and the classical theory is similar to the difference between using bottles of water (quantum) or water from a Episode (classical). The bottles represent the quantum idea and the continuous flow from the Episode represents the classical theory.

The quantisation of energy is also rather like the kangaroo motion of a car when you first learn to drive – it jumps from one energy state to another, there is no smooth acceleration.

It is all a question of scale. We do not 'see' quantum effects generally in everyday life because of the very small value of Planck's constant. Think about a person and an ant walking across a gravelled path. The size of the individual pieces of gravel may seem small to us but they are giant boulders to the ant.



We know that the photons emitted by a light bulb, for example, travel at the speed of light

(3×10^8 m s^{-1}) so why don't we feel them as they hit us? (Although all energy is quantised we are not aware of this in everyday life because of the very small value of Planck's constant.)

Students may worry about the exact nature of photons. It may help if you give them this quotation from Einstein:

'All the fifty years of conscious brooding have brought me no closer to the answer to the question, "What are light quanta?". Of course, today every rascal thinks he knows the answer, but he is deluding himself.'

Worked example: Photon flux
Calculate the number of quanta of radiation being emitted by a light source.

Consider a green 100 W light. For green light the wavelength is about 6×10^{-7} m and so:

Energy of a photon = $E = hf = hc / l = 3.3 \times 10^{-19}$ J

The number of quanta emitted per second by the light $N = 100 \times l / hc = 3 \times 10^{20}$ s^{-1}.

Student calculations: Photon flux
Episode 501-2: Photons streaming from a lamp (Word, 22 KB)

Episode 501-3: Quanta (Word, 27 KB)

Student experiment: Relating photon energy to frequency
Episode 501-4: Relating photon energy to frequency (Word, 70 KB)

Students can use LEDs of different colours to investigate the relationship between frequency and photon energy for light.

Download this episode
Episode 501: Spectra and energy levels (Word, 138 KB)

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