A-LEVEL PHYSICS • EXERCISE SHEET

22.4 Energy levels in atoms and line spectra

Name: ______________ Class: __________ Date: ____________ **Total: 25 marks**

KEY WORDS emission spectrum 发射光谱 • energy level 能级 • discrete 分立 • absorption spectrum 吸收光谱 • electron 电子

Objective

Build the skills to answer exam questions on **energy levels** and line spectra.

You must be able to:

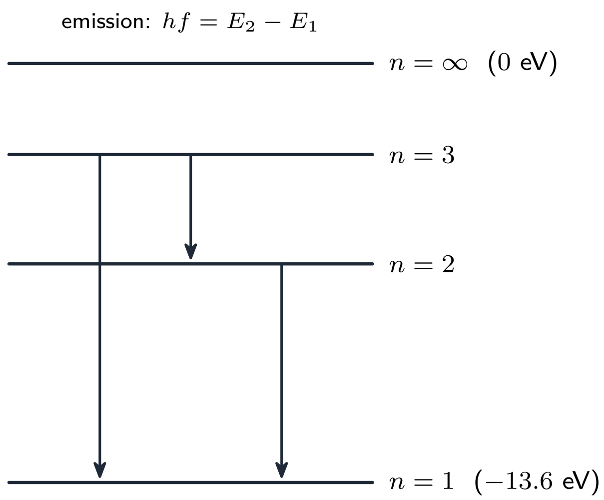
- explain **discrete energy levels** and line spectra
- use $hf = E_2 - E_1$

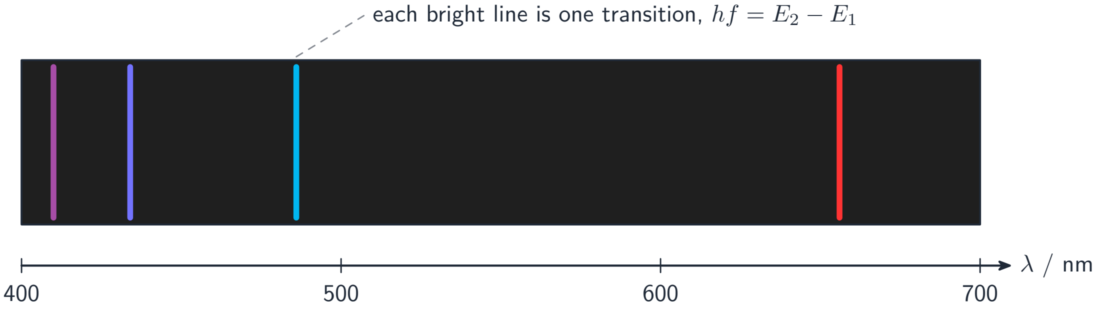
1 Worked examples

Study these first. Each one shows the method for a question type used later — follow the steps and you can do the Practice and Exam-style questions yourself.

■ Energy levels

An atom's electrons can only sit at **discrete** energy levels (written **negative**, with 0 for a free electron):





Line spectra from atomic energy levels

When an electron **drops** from E_2 to a lower E_1, it emits a photon:

$$hf = E_2 - E_1.$$

Because the levels are **discrete**, only certain photon energies (wavelengths) appear — an **emission spectrum** is a set of **bright lines** on a dark background (a fingerprint of the element). An **absorption spectrum** is dark lines on a bright background.

2 Practice

Now apply the methods above.

2.1 State what is meant by **discrete** energy levels. [1]

2.2 Write the equation linking photon frequency to two energy levels. [1]

2.3 State why the energy levels are written as **negative** values. [1]

2.4 State what an **emission spectrum** looks like. [1]

3 Exam-style questions

3.1 An **emission** line spectrum consists of: [1]

- **A** a continuous range of colours
- **B** dark lines on a bright background
- **C** bright lines on a dark background
- **D** a single colour only

3.2 An electron drops from an energy level of -3.4 eV to -13.6 eV. Calculate the **energy** of the emitted photon, in eV and in joules. [3]

3.3 Calculate the **wavelength** of the photon emitted in question 3.2. [3]

3.4 Explain why only **certain** wavelengths appear in an emission spectrum. [2]

3.5 The lowest four energy levels of a hydrogen atom are at -13.6 eV, -3.40 eV, -1.51 eV and -0.85 eV. Take $h = 6.63 \times 10^{-34}$ J s, $c = 3.00 \times 10^{8}$ m s^{-1} and 1 eV $= 1.60 \times 10^{-19}$ J.

(a) Use the **photon model** to explain how the existence of sharp **spectral lines** shows that the atom has discrete energy levels. [3]

(b) **Calculate** the energy of the ground state in **joules**. [1]

(c) **Show that** the energy difference between the $n = 1$ and $n = 2$ levels is 10.2 eV. [1]

(d) **Complete** the table for transitions **down to the ground state**. [3]

transition	**energy difference / eV**	**wavelength of photon / nm**
$n = 2 \rightarrow n = 1$	10.2	____________
$n = 3 \rightarrow n = 1$	____________	____________
$n = 4 \rightarrow n = 1$	____________	____________

(e) **State and explain** whether these three photons are visible to the eye. [2]

(f) **Explain** why the energies of the levels are given as **negative** numbers. [2]

Solutions

2.1 The electron can only have **certain fixed energies** — never values in between.

2.2 $hf = E_2 - E_1$.

2.3 Energy is measured from 0 for a **free** electron, and a bound electron has **less** energy than that — so it is negative.

2.4 A series of **sharp bright lines** on a dark background.

3.1 **C** — bright lines on a dark background.

3.2 $\Delta E = -3.4 - (-13.6) = 10.2$ eV; in joules $= 10.2 \times 1.6 \times 10^{-19} = 1.6 \times 10^{-18}$ J.

3.3 $\lambda = \dfrac{hc}{E} = \dfrac{(6.63 \times 10^{-34})(3.0 \times 10^{8})}{1.6 \times 10^{-18}} = 1.2 \times 10^{-7}$ m.

3.4 The energy levels are **discrete**, so the differences $E_2 - E_1$ — and hence the photon energies/wavelengths — can only take **certain values**.

3.5 (a) An atom emits a photon when an electron falls from a higher level to a lower one, and the photon energy equals the **difference** between the two levels. Since $E = hf$, one energy difference gives one frequency and so one sharp line. A continuous range of energies would give a continuous spectrum, so sharp lines show only **certain** energies are allowed, i.e. the levels are discrete.
(b) $-13.6 \times 1.60 \times 10^{-19} = -2.18 \times 10^{-18}$ J.
(c) $-3.40 - (-13.6) = 10.2$ eV.
(d) Differences: 10.2, 12.09 and 12.75 eV. Wavelengths from $\lambda = \dfrac{hc}{E}$: 122 nm, 103 nm and 97 nm.
(e) **No** — all three lie near 100 nm, in the **ultraviolet**, well below the roughly 400–700 nm visible range.
(f) The zero of energy is taken as the electron being **free**, at infinity and at rest; a bound electron has less energy than that, so energy must be **supplied** to free it and its energy is written as negative.