Exercise walk-through

Energy levels in atoms and line spectra ▪ 22.4

eXercise

You must be able to

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

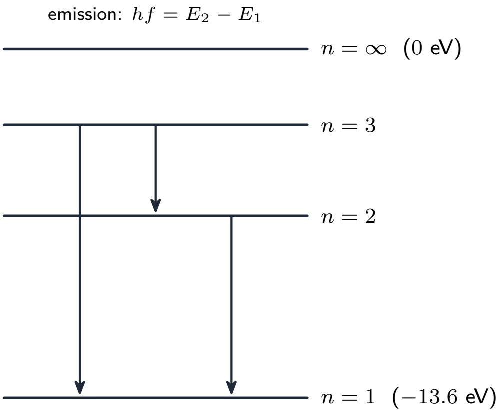
Method — Energy levels

An atom's electrons can only sit at **discrete** energy levels (written **negative**, with 0 for a free electron):
When an electron **drops** from E_2 to a lower E_1, it emits a photon:

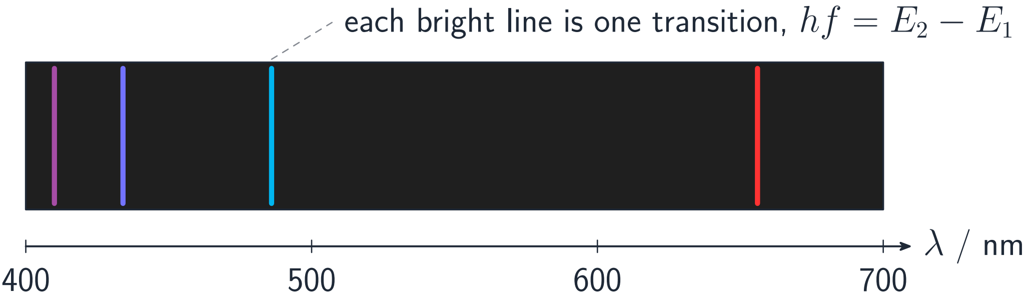
$$hf = E_2 - E_1.$$

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.

Method — Energy levels



Method — Energy levels



Line spectra from atomic energy levels

Practice 2.1 ▪ 1 mark

State what is meant by **discrete** energy levels.

Solution 2.1

Method: Energy levels

Worked steps

The electron can only have **certain fixed energies** — never values in between B1.

Practice 2.2 ▪ 1 mark

Write the equation linking photon frequency to two energy levels.

Solution 2.2

Method: Energy levels

Worked steps

$hf = E_2 - E_1$ **B1**.

Practice 2.3 ▪ 1 mark

State why the energy levels are written as **negative** values.

Solution 2.3

Method: Energy levels

Worked steps

Energy is measured from 0 for a **free** electron, and a bound electron has **less** energy than that — so it is negative **B1**.

Practice 2.4 ▪ 1 mark

State what an **emission spectrum** looks like.

Solution 2.4

Method: Energy levels

Worked steps

A series of **sharp bright lines** on a dark background B1.

Exam-style 3.1 ▪ 1 mark

An **emission** line spectrum consists of:

- **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

Solution 3.1

Method: Energy levels

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

Worked steps

C — bright lines on a dark background.

Exam-style 3.2 ▪ 3 marks

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.

Solution 3.2

Method: Energy levels

Worked steps

$\Delta E = -3.4 - (-13.6) = 10.2$ eV **M1** **A1**; in joules $= 10.2 \times 1.6 \times 10^{-19} = 1.6 \times 10^{-18}$ J **A1**.

Exam-style 3.3 ▪ 3 marks

Calculate the **wavelength** of the photon emitted in question 3.2.

Solution 3.3

Worked steps

$$\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34})(3.0 \times 10^{8})}{1.6 \times 10^{-18}}$$ **M1** **M1** $= 1.2 \times 10^{-7}$ m **A1**.

Exam-style 3.4 ▪ 2 marks

Explain why only **certain** wavelengths appear in an emission spectrum.

Solution 3.4

Method: Energy levels

Worked steps

The energy levels are **discrete**, so the differences $E_2 - E_1$ — and hence the photon energies/wavelengths — can only take **certain values** B1 B1.

Exam-style 3.5 ▪ 3 marks

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.

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

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

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

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

Exam-style 3.5 ▪ 3 marks

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

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

Solution 3.5

Method: Energy levels

Worked steps

(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 **B1**. Since $E = hf$, one energy difference gives one frequency and so one sharp line **B1**. 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 **B1**.
(b) $-13.6 \times 1.60 \times 10^{-19} = -2.18 \times 10^{-18}$ J **A1**.
(c) $-3.40 - (-13.6) = 10.2$ eV **A1**.
(d) Differences: 10.2, 12.09 and 12.75 eV **B1**. Wavelengths from $\lambda = \dfrac{hc}{E}$: 122 nm,

Solution 3.5

103 nm and 97 nm B1 B1.

(e) **No** B1 — all three lie near 100 nm, in the **ultraviolet**, well below the roughly 400–700 nm visible range B1.

(f) The zero of energy is taken as the electron being **free**, at infinity and at rest B1; a bound electron has less energy than that, so energy must be **supplied** to free it and its energy is written as negative B1.