

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:

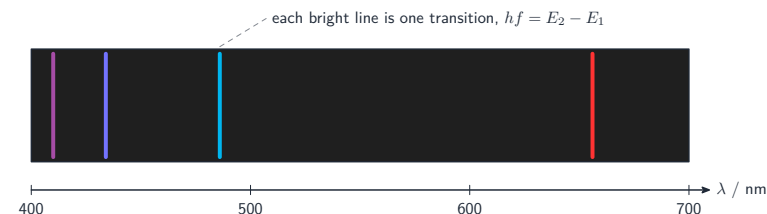
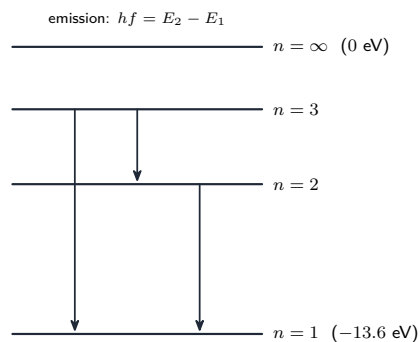
- explain **discrete energy levels** and line spectra
- use $hf = E_2 - E_1$

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 = 1.60×10^{-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

Each answer shows the steps, then the **result**.

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

- an emission spectrum is **bright lines on a dark background**

3.2

- $\Delta E = E_2 - E_1$

$$\Delta E = -3.4 \text{ eV} - (-13.6 \text{ eV}) = \mathbf{10.2 \text{ eV}}$$
$$10.2 \times 1.6 \times 10^{-19} \text{ J eV}^{-1} = \mathbf{1.6 \times 10^{-18} \text{ J}}$$

3.3

- $\lambda = hc/E$

$$\lambda = \frac{(6.63 \times 10^{-34} \text{ J s})(3.0 \times 10^8 \text{ m s}^{-1})}{1.6 \times 10^{-18} \text{ J}} = \mathbf{1.2 \times 10^{-7} \text{ m}}$$

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) $-13.6 \times 1.60 \times 10^{-19} = \mathbf{-2.18 \times 10^{-18} \text{ J}}$

(c) $-3.40 - (-13.6) = \mathbf{10.2 \text{ eV}}$

(d) differences: 10.2, 12.09 and 12.75 eV

- wavelengths from $\lambda = hc/E$: 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