

Atomic energy levels and line spectra

Physics grades 11-12 (A Level · IB · Class 11-12) · Lesson

Quick answer: Atoms have discrete energy levels, so an electron changing level emits or absorbs a photon whose energy equals the gap between the levels. This gives sharp lines instead of a continuous band. The photon wavelength in nanometres is 1240 divided by its energy in electronvolts.

Level: Physics grades 11-12 (A Level · IB · Class 11-12) · **Curriculum:** IB E.1

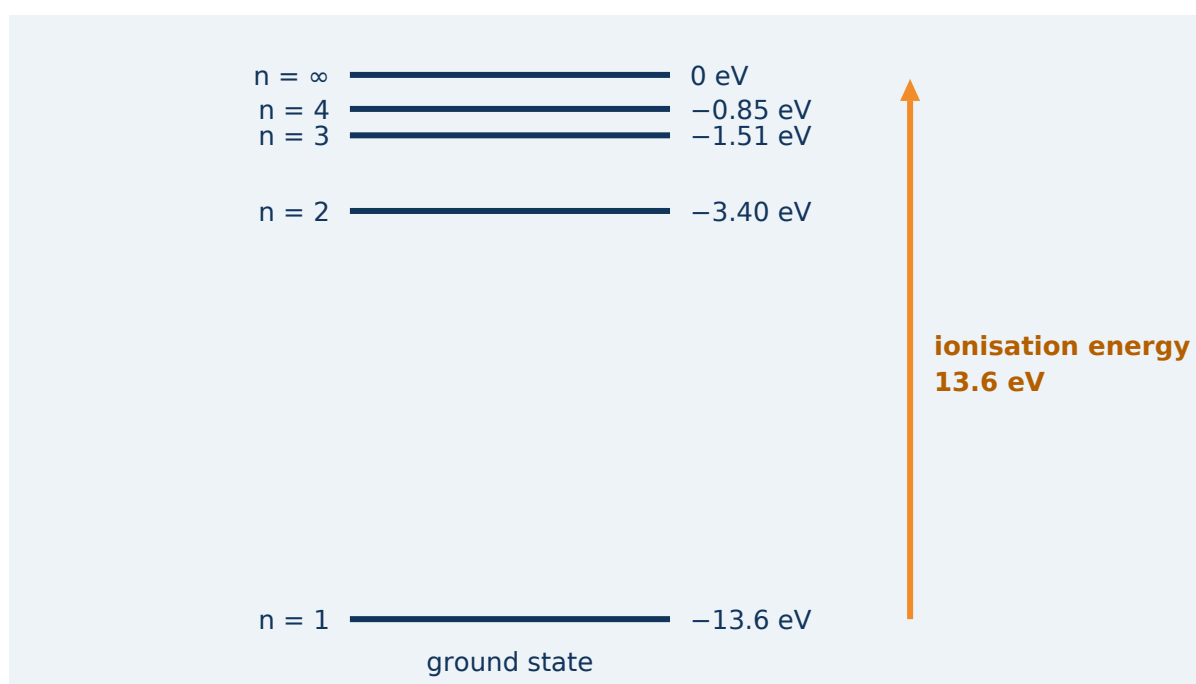
What are energy levels in an atom?

An electron bound in an atom cannot have just any energy. Only certain values are allowed, like the rungs of a ladder with no standing room between them. The values are negative, because energy must be supplied to pull the electron free; the zero of energy is placed where the electron has only just escaped.

Definition

An **energy level** is one of the discrete energies an electron bound in an atom may have. The lowest level is the **ground state**; higher levels are **excited states**. The energy needed to remove an electron from the ground state is the **ionisation energy**.

For hydrogen the levels are given by $E = -13.6 \text{ eV} / n^2$, where $n = 1, 2, 3 \dots$ is the level number. The ground state is -13.6 eV , then come -3.40 eV , -1.51 eV and -0.85 eV . The levels crowd together as n grows and converge on 0 eV .



Hydrogen energy levels: the gaps shrink as n increases, and 13.6 eV lifts the electron out of the atom.

What is an electronvolt and why is it used?

A joule is enormous compared with the energy of one electron in an atom, so atomic physics uses a smaller unit that is tied to the electron itself.

Formula

$$1 \text{ eV} = 1.60 \times 10^{-19} \text{ J}$$

One electronvolt is the energy gained by one electron accelerated through a potential difference of 1 V. To change eV into J multiply by 1.60×10^{-19} ; to change J into eV divide by it.

Worked example 1

Convert the ionisation energy of hydrogen, 13.6 eV, into joules.

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

How do transitions produce emission lines?

When an electron falls from a higher level to a lower one, the atom emits a single photon. The photon carries away exactly the energy difference between the two levels, so each possible fall gives one precise colour.

Key law

Photon energy = energy difference: $hf = E_i - E_f$, and since $f = c / \lambda$, $E = hc / \lambda$.

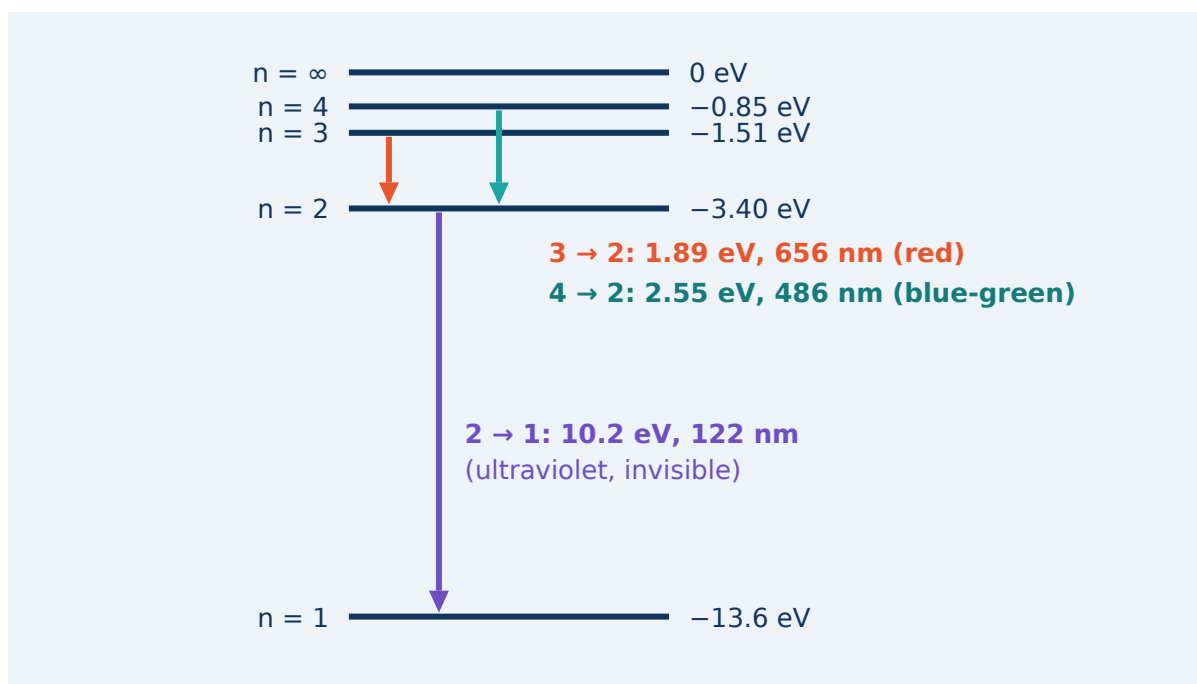
With E in eV and λ in nm, use $hc \approx 1240 \text{ eV nm}$: $\lambda \text{ (nm)} = 1240 / E \text{ (eV)}$. Here E_i and E_f are the initial and final level energies, h is the Planck constant and c is the speed of light.

Worked example 2

Find the wavelength of the photon emitted when a hydrogen electron falls from $n = 3$ to $n = 2$.

$$\text{Energy gap} = -1.51 - (-3.40) = 1.89 \text{ eV}$$

$$\lambda = 1240 / 1.89 = \mathbf{656 \text{ nm}}, \text{ which is red light.}$$

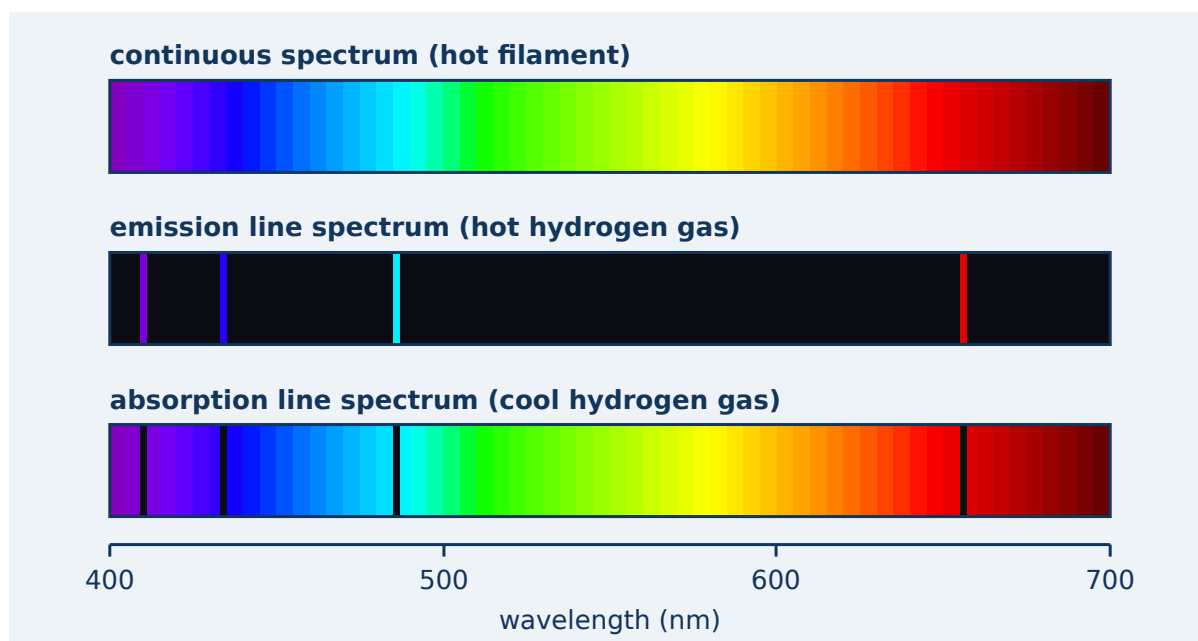


Three hydrogen transitions: bigger gaps give photons of higher energy and shorter wavelength.

Transition	Photon energy (eV)	Wavelength (nm)	Region
3 → 2	1.89	656	red
4 → 2	2.55	486	blue-green
5 → 2	2.86	434	violet
2 → 1	10.2	122	ultraviolet

How do emission and absorption spectra differ?

A hot solid glows with every colour, a continuous spectrum. A hot, low-pressure gas gives only the colours its atoms can emit: bright lines on a dark background. If white light passes through a cooler gas, atoms absorb photons whose energy matches a gap, and the same wavelengths are missing from the light that emerges, leaving dark lines.



The dark lines of an absorption spectrum sit exactly where the bright lines of the emission spectrum appear.

Common mistake

Students often say an atom absorbs *any* photon with enough energy. A photon is absorbed only if its energy equals a gap between two levels exactly (or exceeds the ionisation energy). A colliding electron, in contrast, can give up just part of its kinetic energy.

Lina's tip

Draw the arrow between the two levels, write both energies, subtract (larger minus smaller), and only then convert with $1240 / E$. The photon energy is always positive.

What can spectra tell us about stars and materials?

Every element has its own pattern of levels, so its pattern of lines is a fingerprint. Astronomers match the dark lines in starlight with laboratory patterns to find which elements are present, and chemists identify metals in a flame or an alloy the same way. Helium was first found in sunlight before anyone had isolated it on Earth.

Key points to remember

- Hydrogen levels: $E = -13.6 \text{ eV} / n^2$, so $n = 1$ is -13.6 eV
- $1 \text{ eV} = 1.60 \times 10^{-19} \text{ J}$
- Transition: photon energy $hf = E_{\text{initial}} - E_{\text{final}}$
- $\lambda \text{ (nm)} = 1240 / E \text{ (eV)}$
- Emission lines are bright; absorption lines of the same gas are dark at the same wavelengths

Atomic energy levels and line spectra: frequently asked questions

What is an energy level?

It is one of the fixed energies an electron bound in an atom may have. The electron cannot have energies between the levels, which is why atoms emit only particular colours.

Why are the energy levels of hydrogen negative?

Zero energy is defined as the electron just free from the atom. A bound electron has less energy than that, so its energy is negative, and -13.6 eV means 13.6 eV must be supplied to free it from the ground state.

How do you convert electronvolts to joules?

Multiply the energy in eV by 1.60×10^{-19} . For instance 2.0 eV is $3.2 \times 10^{-19} \text{ J}$. To go the other way, divide joules by 1.60×10^{-19} .

Why does an absorption spectrum have dark lines?

Atoms in the cool gas absorb photons whose energy matches a gap between their levels, then re-emit them in all directions. Those wavelengths are therefore weak in the transmitted beam and appear dark.

How do you find the wavelength of an emitted photon?

Subtract the final level energy from the initial level energy to get the photon energy in eV, then divide 1240 by it to get the wavelength in nm.