

Light and Atomic Spectra

How does light reveal the energy levels inside atoms?

●●○ Intermediate

Atomic Structure & Periodicity

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Before you start: it helps to understand [Inside the Atom: Protons, Neutrons and Electrons](#).

What is it?

Light is **electromagnetic radiation**: a wave of electric and magnetic fields that travels through space at the **speed of light**, $c = 2.998 \times 10^8$ m/s. A wave is described by its:

- **wavelength**, λ (lambda): the distance from one crest to the next, in m or nm;
- **frequency**, ν (nu): the number of waves passing a point each second, in s^{-1} or hertz (Hz).

They are linked by

$$c = \lambda\nu$$

Light also behaves as a stream of particles called **photons**. Each photon carries a fixed amount (a **quantum**) of energy:

$$E = h\nu = \frac{hc}{\lambda}$$

where $h = 6.626 \times 10^{-34}$ J s is **Planck's constant**. Short wavelength means high frequency and **high energy**.

KEY IDEA

When atoms are heated or excited by electricity, they emit light of only certain **exact wavelengths**: a **line spectrum**, different for every element. This shows that electrons in atoms can have only certain **energies** (energy levels). Each line is a photon released when an electron drops from a higher level to a lower one.

Why does it matter?

- **Identifying elements.** Every element has its own line spectrum, like a fingerprint. Flame tests, street lamps and the analysis of starlight all rely on it; helium was discovered in the Sun's spectrum before it was found on Earth.

- **Energy levels lead to electron configurations.** The idea of quantized energy levels is the starting point for shells, subshells and orbitals.
- **Light and chemistry.** Photons with enough energy break bonds (UV light causes sunburn), and the colours of solutions are the basis of spectrophotometry.

How does it work?

1. The electromagnetic spectrum

From low to high energy: radio waves, microwaves, infrared, **visible light** (about 400 nm violet to 700 nm red), ultraviolet, X-rays and gamma rays. Only a narrow band is visible to us.

2. Line spectra

A hot solid gives a **continuous spectrum** (all colours, like a rainbow). A gas of one element, excited by heat or electricity, gives an **emission line spectrum**: bright lines at fixed wavelengths on a dark background. Hydrogen's visible lines are at 656 nm (red), 486 nm (blue-green), 434 nm and 410 nm (violet). Cool gas in front of a bright source **absorbs** the same wavelengths, giving dark lines (an absorption spectrum).

3. The Bohr model of hydrogen

In 1913 Niels Bohr proposed that the electron in a hydrogen atom can only occupy orbits with certain energies:

$$E_n = -\frac{2.179 \times 10^{-18} \text{ J}}{n^2}$$

where $n = 1, 2, 3, \dots$

- $n = 1$ is the **ground state**, the lowest energy. Higher n are **excited states**.
- Energies are **negative** because the electron is bound: zero energy means the electron has been removed ($n = \infty$).
- When the electron drops from a higher level to a lower one, it emits **one photon** whose energy equals the gap:

$$\Delta E = 2.179 \times 10^{-18} \text{ J} \times \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

(written as a positive energy for the photon emitted, with n_f the lower level). Drops to $n = 2$ give the **visible** lines of hydrogen; drops to $n = 1$ give ultraviolet lines.

∞ THINK OF IT LIKE THIS

Energy levels are like the rungs of a ladder, not a ramp. You can stand on a rung, but not between rungs. Jumping down from one rung to another always releases the same exact amount of energy, which is why each jump gives light of one exact colour.

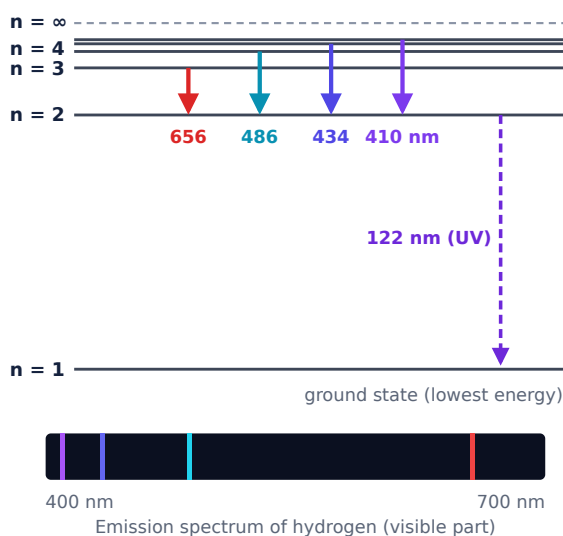
🔍 MORE PRECISELY

Bohr's model works perfectly only for hydrogen and other one-electron ions. It was replaced by **quantum mechanics**, in which electrons are described by **orbitals** (regions where the electron is likely to be found) rather than fixed orbits. The energy levels of hydrogen predicted by quantum mechanics are the same as Bohr's, which is why his formula is still used. The wave-particle nature of light was confirmed by the photoelectric effect, explained by Einstein in 1905.

Visualise it

Hydrogen energy levels

Drops to $n = 2$ give the visible lines.



Each visible line of hydrogen comes from an electron dropping to the $n = 2$ level.

Worked example

WORKED EXAMPLE: Frequency from wavelength

Question: Hydrogen's red line has a wavelength of 656 nm. What is its frequency?

Step 1 Convert to metres: $656 \text{ nm} = 656 \times 10^{-9} \text{ m}$

Step 2 Rearrange $c = \lambda\nu$:

$$\begin{aligned}\nu &= \frac{c}{\lambda} \\ &= \frac{2.998 \times 10^8 \text{ m/s}}{656 \times 10^{-9} \text{ m}} \\ &= 4.57 \times 10^{14} \text{ s}^{-1}\end{aligned}$$

WORKED EXAMPLE: Energy of a photon and of a mole of photons

Question: What is the energy of one photon of the 656 nm light, and of one mole of these photons?

Step 1 One photon:

$$\begin{aligned}E &= h\nu \\ &= 6.626 \times 10^{-34} \text{ J s} \\ &\quad \times 4.570 \times 10^{14} \text{ s}^{-1} \\ &= 3.03 \times 10^{-19} \text{ J}\end{aligned}$$

Step 2 One mole:

$$\begin{aligned}&3.028 \times 10^{-19} \text{ J} \\ &\quad \times 6.022 \times 10^{23} \text{ mol}^{-1} \\ &= 1.824 \times 10^5 \text{ J/mol} \\ &= 182 \text{ kJ/mol}\end{aligned}$$

 WORKED EXAMPLE: Predicting a line with the Bohr model

Question: Calculate the wavelength of the photon emitted when a hydrogen electron drops from $n = 3$ to $n = 2$.

Step 1 Energy gap:

$$\begin{aligned}\Delta E &= 2.179 \times 10^{-18} \text{ J} \\ &\times \left(\frac{1}{2^2} - \frac{1}{3^2} \right) \\ &= 2.179 \times 10^{-18} \text{ J} \\ &\times 0.1389 \\ &= 3.026 \times 10^{-19} \text{ J}\end{aligned}$$

Step 2 Wavelength, from $E = hc/\lambda$:

$$\begin{aligned}\lambda &= \frac{hc}{\Delta E} \\ &= \frac{6.626 \times 10^{-34} \text{ J s}}{3.026 \times 10^{-19} \text{ J}} \\ &\quad \times 2.998 \times 10^8 \text{ m/s} \\ &= 6.56 \times 10^{-7} \text{ m} \\ &= 656 \text{ nm}\end{aligned}$$

Step 3 This is exactly hydrogen's red line: the Bohr model explains the observed spectrum.

 WORKED EXAMPLE: Ionization energy of hydrogen

Question: How much energy is needed to remove the electron from a hydrogen atom in its ground state? Give the answer per mole.

Step 1 From $n = 1$ ($E_1 = -2.179 \times 10^{-18} \text{ J}$) to $n = \infty$ ($E = 0$): $\Delta E = +2.179 \times 10^{-18} \text{ J}$ per atom.

Step 2 Per mole:

$$\begin{aligned}&2.179 \times 10^{-18} \text{ J} \\ &\quad \times 6.022 \times 10^{23} \text{ mol}^{-1} \\ &= 1.312 \times 10^6 \text{ J/mol} \\ &= 1312 \text{ kJ/mol}\end{aligned}$$

Step 3 This matches the measured first ionization energy of hydrogen, 1312 kJ/mol.

Common mistake

COMMON MISTAKE: Forgetting to convert nm to m

In $c = \lambda\nu$ and $E = hc/\lambda$, the wavelength must be in **metres** to match c in m/s. 656 nm is 6.56×10^{-7} m, not 656 m.

COMMON MISTAKE: Thinking longer wavelength means more energy

Energy is **inversely** proportional to wavelength. Red light (700 nm) has **less** energy per photon than violet light (400 nm), and UV has more than both.

COMMON MISTAKE: Mixing per-photon and per-mole energies

$E = h\nu$ gives the energy of **one** photon (about 10^{-19} J). To compare with bond energies in kJ/mol, multiply by Avogadro's number and divide by 1000.

NOTATION NOTE

- ν is the Greek letter nu (frequency), not the letter v (velocity).
- $1 \text{ Hz} = 1 \text{ s}^{-1}$; $1 \text{ nm} = 10^{-9} \text{ m}$.

Remember this

REMEMBER THIS

- $c = \lambda\nu$ and $E = h\nu = hc/\lambda$; $c = 2.998 \times 10^8 \text{ m/s}$, $h = 6.626 \times 10^{-34} \text{ J s}$.
- Short wavelength = high frequency = high energy. Visible light: about 400–700 nm.
- Atoms emit and absorb only certain wavelengths (line spectra) because electron energy levels are quantized.
- Bohr: $E_n = -2.179 \times 10^{-18} \text{ J}/n^2$; a photon carries the energy difference between two levels.
- Per mole: multiply the photon energy by N_A .

References

1. Bureau International des Poids et Mesures (BIPM). The International System of Units (SI), 9th ed.; BIPM, 2019.
[Link](#)
2. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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