

E.1 Structure of the atom

Ahmed Gomaa · IB DP Physics SL and HL (sections marked HL are for HL only) · uses D.2 (electric force and potential energy) · original worked examples

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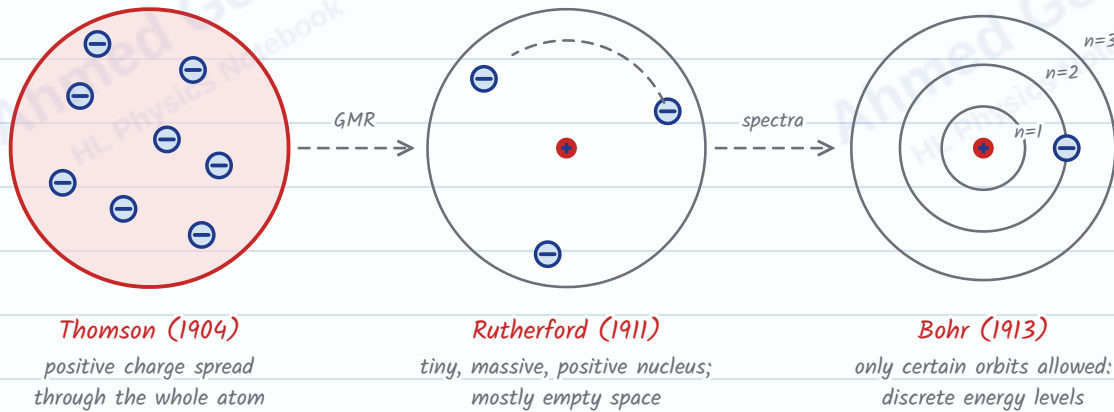
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Pen colours: blue = objects and paths, red = laws, forces and key steps, green = answers and results, pencil = construction lines. Boxes: **Why?** = the reasoning, **Mistake** = what students get wrong, yellow note = exam tip, dashed box = a derivation step by step.

The big picture

What is inside an atom? Nobody can see it, so physicists built **models** and tested them with experiments. Two pieces of evidence changed the model: the **scattering of alpha particles** showed that the positive charge sits in a tiny nucleus, and the **line spectra** of gases showed that an atom can only

have certain energies. This topic follows that story. Old models are not thrown away; each one still explains some things well.



From Thomson to Rutherford to Bohr: each new model came from an experiment the old one could not explain.

Quantity	Symbol and unit	Meaning
proton number	Z (no unit)	number of protons in the nucleus
nucleon number	A (no unit)	number of protons + neutrons
photon energy	E , J or eV	$E = hf = hc/\lambda$; $1 \text{ eV} = 1.60 \times 10^{-19} \text{ J}$
Planck constant	$h = 6.63 \times 10^{-34} \text{ J s}$	links a photon's energy to its frequency
distance of closest approach HL	d , m	nearest a head-on particle gets to a nucleus
nuclear radius HL	$R = R_0 A^{1/3}$, $R_0 = 1.2 \times 10^{-15} \text{ m}$	size of a nucleus (about 10^{-15} m)

Sizes to keep in mind: an atom is about 10^{-10} m across; a nucleus is about 10^{-15} m , so the nucleus is about **100 000 times smaller** than the atom. 1 fm (femtometre) = 10^{-15} m .

Nuclear notation

Learn it word for word.

Nucleon: a proton or a neutron.

Proton number Z : the number of protons in the nucleus.

Nucleon number A : the number of protons and neutrons in the nucleus.

Isotopes: nuclei of the same element (same Z) with different numbers of neutrons (different A).

Nuclear notation: ${}^A_ZX$ · X = chemical symbol · number of neutrons N
 $= A - Z$

A neutral atom has Z electrons. (You do not need to remember chemical symbols; they are given.)

Worked example 1

A gold nucleus is written ${}^{197}_{79}\text{Au}$. (a) How many protons, neutrons and nucleons does it have? (b) How many electrons are in a neutral gold atom and in the ion Au^{3+} ? (c) ${}^{195}_{79}\text{Au}$ is another nucleus. What is the same and what is different?

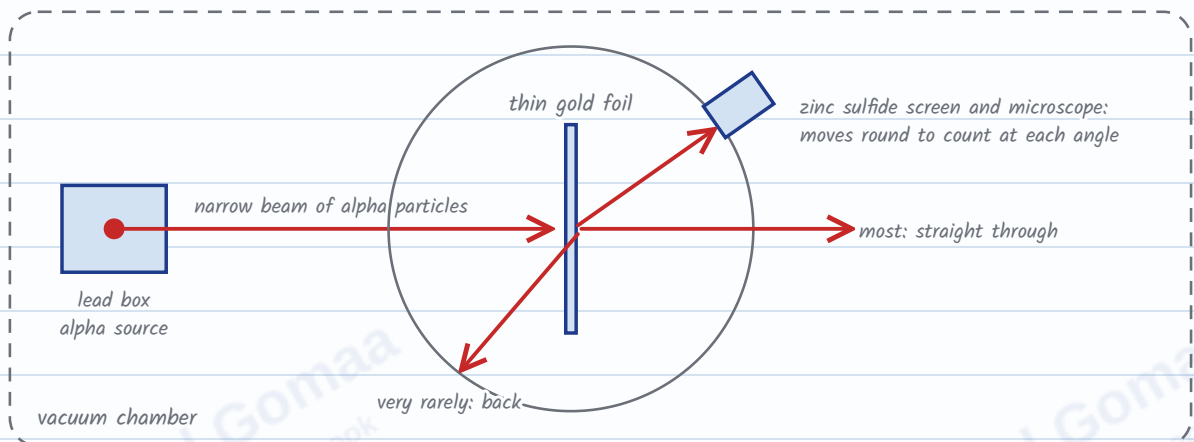
1. (a) Protons $Z = 79$; nucleons $A = 197$; neutrons $197 - 79 = 118$.

2. (b) Neutral atom: **79 electrons**. Au^{3+} has lost 3 electrons: **76 electrons** (the nucleus does not change).

3. (c) Same $Z = 79$, so it is the same element: they are **isotopes**. It has 2 fewer neutrons (116).

The Geiger-Marsden-Rutherford experiment

In 1909–1911 Geiger and Marsden, working with Rutherford, fired **alpha particles** (helium nuclei, charge $+2e$, about 7000 times the mass of an electron) at a very **thin gold foil** in a **vacuum**, and counted how many came off at each angle.

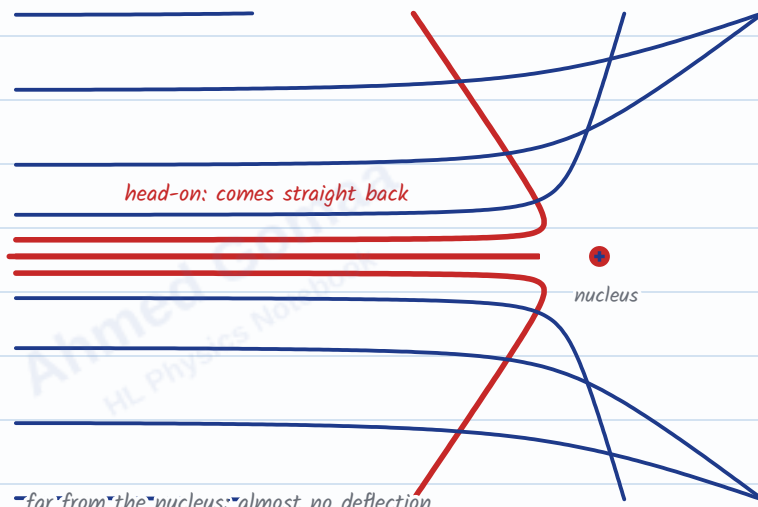


Each alpha particle that hits the screen makes a tiny flash of light (a scintillation), counted by eye.

The vacuum stops the alpha particles being slowed or scattered by air. The foil is very thin so that each alpha particle is scattered by (at most) one nucleus.

What they saw

1. **Most** alpha particles went **straight through** the foil, undeflected or with very small deflections.
2. A **few** were deflected through **large angles**.
3. A **very small number** (about 1 in 8000) came almost **straight back** (deflection greater than 90°).



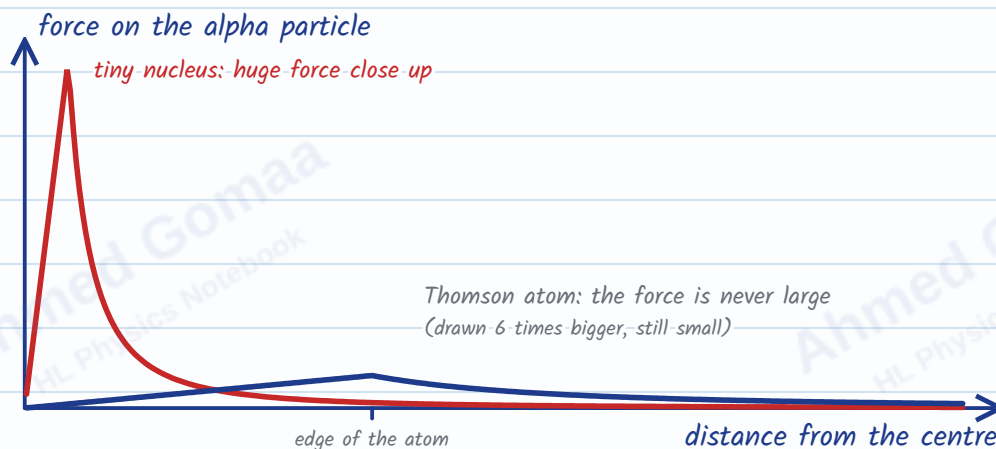
Each path is drawn from the real inverse-square repulsion. Only a particle that passes very close to the nucleus is turned through a large angle.

Why gold, and why alpha particles?

Gold can be hammered into a foil only a few hundred atoms thick, and its nucleus has a large charge ($79e$), which gives strong repulsion. Alpha particles are massive and fast: electrons in the atom are far too light to push them aside, so any large deflection must come from something heavy and positive.

Why the results mean a tiny nucleus

In the **Thomson model** the positive charge is spread through the whole atom. An alpha particle passing through feels only small, spread-out forces, which can explain the **small** deflections, but never a bounce back: the force is never large enough. To turn a fast, heavy alpha particle around, the force must be enormous, and the electric force $F = kQq/r^2$ is enormous only when r is tiny. So the positive charge must be packed into a very small region.



Why the Thomson atom fails: spread-out charge never gives a large force. A tiny nucleus gives a huge force when the alpha particle comes close.

Rutherford's conclusions

- Most of the mass of the atom and all of the positive charge are concentrated in a very small volume: the **nucleus**.
- The atom is **mostly empty space** (most alpha particles pass straight through).
- The electrons orbit the nucleus, far away from it.

Worked example 2

Compare the electric field at the surface of a gold nucleus (radius $7.0 \times 10^{-15} \text{ m}$) with the field at the surface of a Thomson gold atom of the same charge (radius $1.4 \times 10^{-10} \text{ m}$).

1. $E = kQ/r^2$ for both, with the same Q , so the ratio is $(r_{\text{atom}}/r_{\text{nucleus}})^2$.
2. Ratio = $(1.4 \times 10^{-10} / 7.0 \times 10^{-15})^2 = (2.0 \times 10^4)^2 = 4 \times 10^8$.
3. The same charge squeezed into a nucleus gives a field (and a force) hundreds of millions of times larger: enough to send an alpha particle back.

"Explain how the results led to the nuclear model" (3 marks)

- ✓ **Most** alpha particles pass straight through → the atom is mostly **empty space**.
- ✓ **A few** are deflected through large angles (some more than 90°) → a very **large repulsive force**, which needs a **small, positively charged** region.
- ✓ Only a small fraction bounce back → the nucleus is **very small** and contains **most of the mass**.

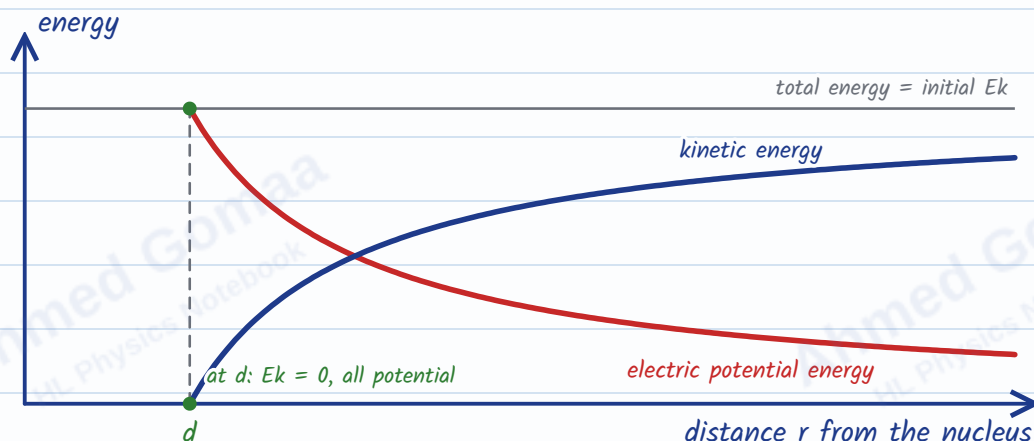
The Thomson model explains only the small deflections: its spread-out charge cannot give the large force.

✗ **Mistake**

Saying "the alpha particles bounce off the gold atoms" or "hit the nucleus". They never touch it (for these energies): they are **repelled** by the electric force. And the experiment tells us about the positive charge and the mass, not about neutrons (they were found only in 1932).

HL The distance of closest approach

Fire an alpha particle **head-on** at a nucleus. As it gets closer it slows down, because the electric force pushes back. It stops for an instant at the **distance of closest approach d** , then comes straight back. Treat the nucleus as fixed (it is much heavier) and use **conservation of energy**: the kinetic energy far away becomes electric potential energy at d .



Kinetic energy turns into electric potential energy. Their sum stays equal to the starting kinetic energy.

Distance of closest approach (head-on):

$E_k = k q_1 q_2 / d \rightarrow$ for an alpha particle on a nucleus of proton number z : $d = k(2e)(Ze) / E_k$

$k = 8.99 \times 10^9 \text{ N m}^2 \text{ C}^{-2}$ · E_k in joules · $e = 1.60 \times 10^{-19} \text{ C}$

Worked example 3

An alpha particle of kinetic energy 6.0 MeV is fired head-on at a gold nucleus ($Z = 79$). (a) Find the distance of closest approach. (b) The radius of the gold nucleus is about $7.0 \times 10^{-15} \text{ m}$. Does the alpha particle reach the nucleus?

1. Charges: alpha $q_1 = 2e$, gold $q_2 = 79e$. $E_k = 6.0 \times 10^6 \times 1.60 \times 10^{-19} \text{ J}$.

2. (a) $d = k(2e)(79e) / E_k = 158ke / (6.0 \times 10^6)$ (one e cancels with the e in "eV") $= 158 \times 8.99 \times 10^9 \times 1.60 \times 10^{-19} / 6.0 \times 10^6 = 3.8 \times 10^{-14} \text{ m}$.

3. (b) $3.8 \times 10^{-14} \text{ m}$ is about 5 times the nuclear radius: **it does not reach the nucleus**. Only the electric force acts, so Rutherford's model works.

Worked example 4

A proton and an alpha particle are each accelerated from rest through a potential difference of 4.0 MV, then fired head-on at a tin nucleus ($Z = 50$). Compare their distances of closest approach.

1. Energy gained = qV . Proton: $E_k = e \times 4.0 \text{ MV} = 4.0 \text{ MeV}$. Alpha (charge $2e$): $E_k = 8.0 \text{ MeV}$.
2. Proton: $d = k(e)(50e)/(4.0 \times 10^6 \text{ e}) = 50 \times 8.99 \times 10^9 \times 1.60 \times 10^{-19}/4.0 \times 10^6 = 1.8 \times 10^{-14} \text{ m}$.
3. Alpha: $d = k(2e)(50e)/(8.0 \times 10^6 \text{ e}) = 1.8 \times 10^{-14} \text{ m}$; the same! The alpha has twice the energy but also twice the charge, so the two effects cancel.

What changes the distance of closest approach?

From $d = kq_1q_2/E_k$:

- **more energy** $\rightarrow$ closer ($d \propto 1/E_k$);
- **bigger nuclear charge Z** $\rightarrow$ further away ($d \propto Z$);
- a different **isotope** of the same element: same Z , so the **same** d (neutrons have no charge).

✗ Mistake

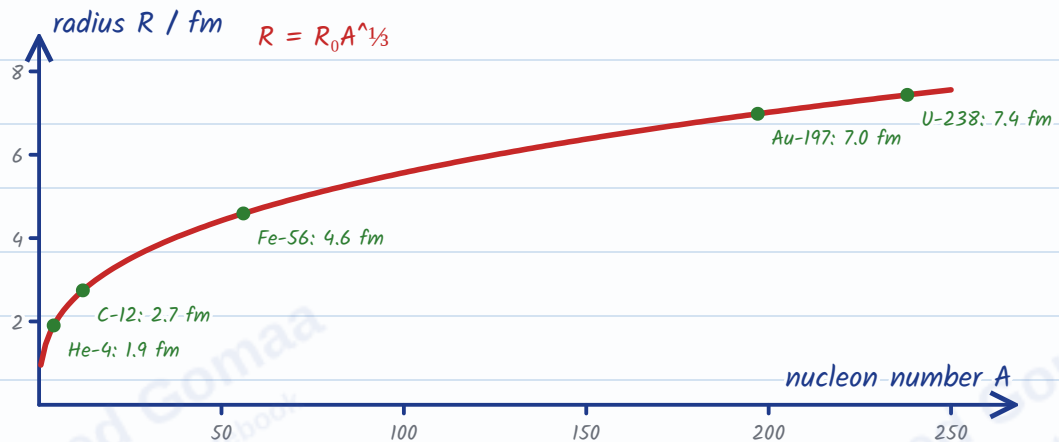
Forgetting that an alpha particle has charge $2e$, using the mass number A instead of Z for the nuclear charge, or leaving the energy in MeV. Change MeV to joules: $\times 10^6 \times 1.60 \times 10^{-19}$.

HL Nuclear radius and nuclear density

How big is a nucleus? Scattering experiments answer this. Fire faster and faster particles: the distance of closest approach shrinks. When the particle just reaches the **surface** of the nucleus, the scattering changes (next

section). The distance of closest approach at that moment is (about) the nuclear radius. Experiments with many nuclei (and with high-energy electrons) give one simple rule:

Nuclear radius: $R = R_0 A^{1/3} \cdot R_0 = 1.2 \times 10^{-15} \text{ m (data booklet)} \cdot A^{1/3}$
 $= \text{nucleon number}^{1/3}$



Doubling the number of nucleons does not double the radius: R grows as the cube root of A .

Worked example 5

(a) Find the radius of a gold-197 nucleus and of a helium-4 nucleus (an alpha particle). (b) How many times larger is the radius of uranium-238 than that of aluminium-27?

1. (a) Gold: $R = 1.2 \times 10^{-15} \times 197^{1/3} = 1.2 \times 10^{-15} \times 5.82 = 7.0 \times 10^{-15} \text{ m}$.
 Helium: $R = 1.2 \times 10^{-15} \times 4^{1/3} = 1.9 \times 10^{-15} \text{ m}$.
2. (b) $R_U/R_{Al} = (238/27)^{1/3} = 8.81^{1/3} = 2.07$. Nearly 9 times the nucleons, only about twice the radius.

Why every nucleus has the same density

Mass of a nucleus $\approx A \times u$ ($u = 1.66 \times 10^{-27}$ kg). Volume $= (4/3)\pi R^3 = (4/3)\pi(R_0 A^{1/3})^3 = (4/3)\pi R_0^3 A$. Both mass and volume are proportional to A , so A **cancels** in the density: $\rho = Au / ((4/3)\pi R_0^3 A) = 3u / (4\pi R_0^3)$. The density is the same for every nucleus. Nucleons are packed like balls in a bag: adding nucleons adds volume in proportion.

Worked example 6

Estimate the density of nuclear matter and compare it with water (1000 kg m^{-3}). What would the mass of a teaspoon (5.0 cm^3) of nuclear matter be?

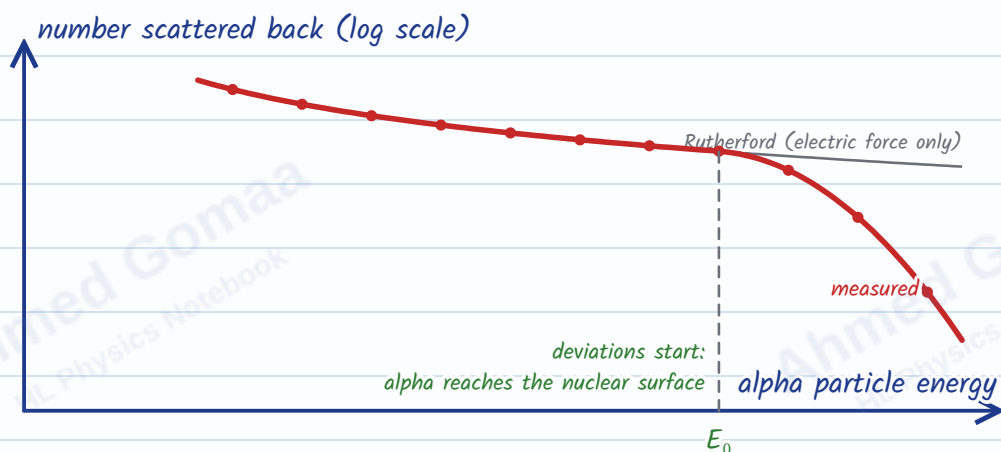
1. $\rho = 3u / (4\pi R_0^3) = 3 \times 1.66 \times 10^{-27} / (4\pi \times (1.2 \times 10^{-15})^3) = 2.3 \times 10^{17} \text{ kg m}^{-3}$.
2. That is about 2×10^{14} times the density of water.
3. Teaspoon: $m = \rho V = 2.3 \times 10^{17} \times 5.0 \times 10^{-6} = 1.1 \times 10^{12} \text{ kg}$ (about a billion tonnes). Only neutron stars are this dense.

✗ Mistake

Thinking a big nucleus is denser. All nuclei have the **same density** (about $2 \times 10^{17} \text{ kg m}^{-3}$); a big nucleus is simply bigger. Also: $R \propto A^{1/3}$, not $R \propto A$.

HL Deviations from Rutherford scattering

Rutherford worked out how many alpha particles should be scattered at each angle, assuming that **only the electric force** acts. At low energies the measurements agree perfectly. But when the alpha particles are given more and more energy, there is a point where the measurements stop agreeing.



Above an energy E_0 , fewer alpha particles are scattered back than Rutherford predicts.

Why the deviations happen

With more energy, the alpha particle gets closer to the nucleus. Above E_0 the distance of closest approach is about the size of the nucleus, so the alpha particle comes within the **range of the strong nuclear force** (about 10^{-15} m). Now there is a second force, not just the electric repulsion, so Rutherford's formula (electric force only) no longer works. The deviations are **evidence for the strong nuclear force**, and for its **short range**: it acts only when the particles are almost touching.

Using E_0 to measure the nucleus

At E_0 the alpha particle just reaches the nuclear surface, so the **nuclear radius $\approx$ the distance of closest approach** at this energy:

$$R \approx d = k(2e)(Ze)/E_0 \text{ (energy conservation, as before)}$$

Worked example 7

Alpha particles fired at gold ($Z = 79$) show deviations from Rutherford scattering when their energy is above about 26 MeV. Estimate the radius of the gold nucleus and compare it with $R_0 A^{1/3}$ ($A = 197$).

1. At 26 MeV the alpha particle just reaches the nucleus: $R \approx d = k(2e)(79e)/E_0$.
2. $d = 158 \times 8.99 \times 10^9 \times 1.60 \times 10^{-19} / 26 \times 10^6 = 8.7 \times 10^{-15} \text{ m}$.
3. $R_0 A^{1/3} = 7.0 \times 10^{-15} \text{ m}$. The estimate is a little larger because at contact the centres are $R_{\text{gold}} + R_{\text{alpha}} = 7.0 + 1.9 = 8.9 \times 10^{-15} \text{ m}$ apart: **very good agreement**.

Worked example 8

The gold foil is replaced by an aluminium foil ($Z = 13$). The alpha particle energy is slowly increased from a low value. Will the deviations from Rutherford scattering start at a lower or a higher energy than for gold? Explain.

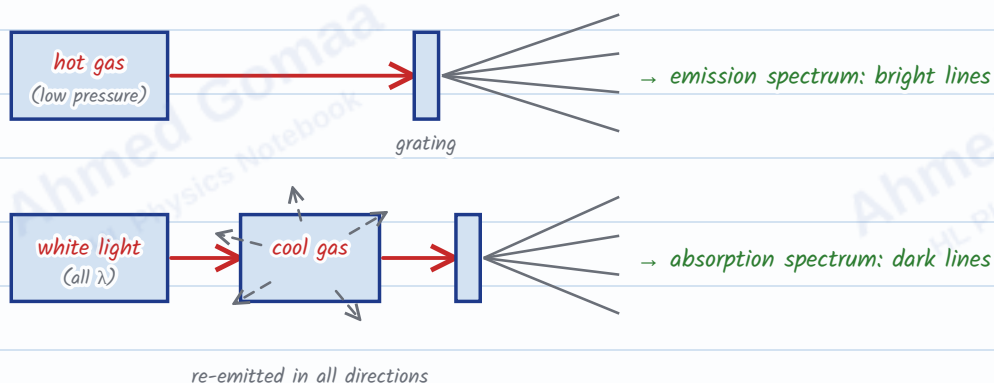
1. $d = k(2e)(Ze)/E_k$: for the same energy, the alpha particle gets closer to an aluminium nucleus (smaller Z , weaker repulsion).
2. The aluminium nucleus is also smaller, but the effect of Z is larger, so the alpha particle reaches the nuclear surface at a much lower energy.
3. **Deviations start at a lower energy for aluminium.**

“Outline why the deviation occurs” (1) · “how E_0 gives the nuclear radius” (2)

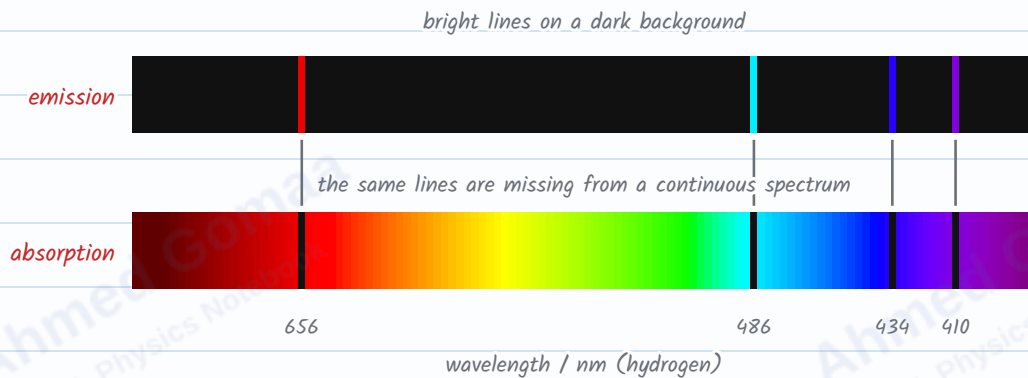
- ✓ At closest approach the alpha particle is within the range of the **strong nuclear force** of the nucleus.
- ✓ The nuclear radius is (approximately) the **distance of closest approach** at E_0 .
- ✓ Find it from **energy conservation**: the initial kinetic energy E_0 becomes electric potential energy, $E_0 = kQq/r$, so $r = kQq/E_0$. And the reverse: if the closest approach is much larger than the nucleus, the strong force does not act and Rutherford's model holds.

Emission and absorption spectra

Heat a gas at low pressure (or pass an electric current through it) and it glows. Pass that light through a diffraction grating or a prism and you do **not** get a rainbow: you get a few **bright lines** of particular wavelengths. Shine white light through a **cool** gas instead and you get a rainbow with **dark lines**, at exactly the same wavelengths.



How the two kinds of line spectrum are made.



Hydrogen: the dark absorption lines are at exactly the same wavelengths as the bright emission lines.

Learn it word for word.

Emission spectrum: the set of wavelengths emitted by an element (bright lines on a dark background).

Absorption spectrum: a continuous spectrum with dark lines at the wavelengths absorbed by an element.

Spectra are fingerprints

Every element has its own set of lines, different from every other element. So the lines tell us **which elements are present**: the chemical composition of a gas, a flame, or a star thousands of light years away. Helium was discovered in the spectrum of the Sun before it was found on Earth: its lines matched no known element.

✗ Mistake

Thinking spectral lines tell us about the nucleus, or about mass-energy. They come from **electrons moving between energy levels**, and they tell us the **chemical composition**.

Energy levels and photons

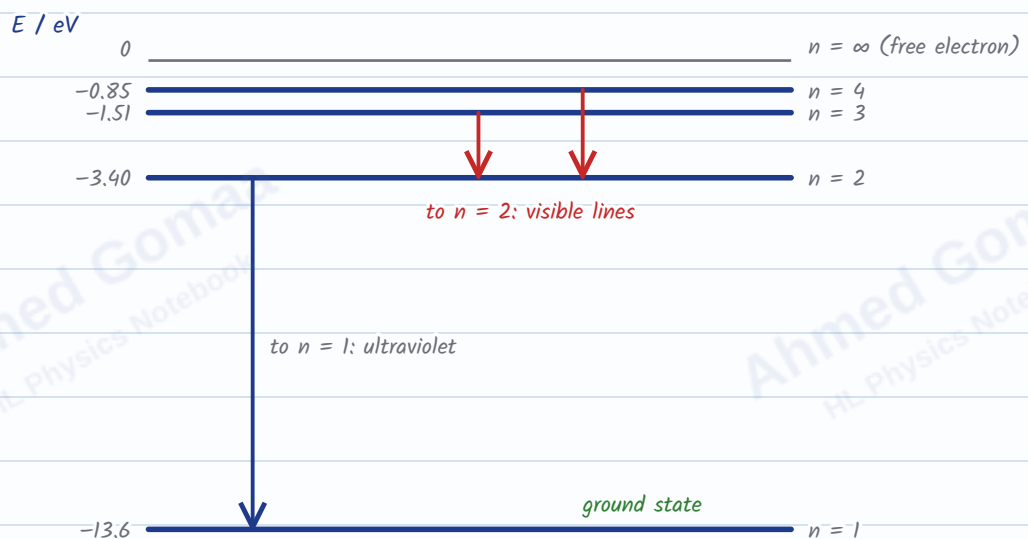
Why only certain wavelengths? Because an atom can only have certain **energies**: its energy is **discrete** (quantised). Each allowed energy is an

energy level. When an electron moves (makes a **transition**) from a higher level to a lower one, the atom loses exactly the difference in energy, carried away by one **photon**. A photon is a packet of light energy.

Photon energy: $E = hf = hc/\lambda \cdot h = 6.63 \times 10^{-34} \text{ J s} \cdot c = 3.00 \times 10^8 \text{ m s}^{-1}$

In a transition: $E_{\text{photon}} = \Delta E = E_{\text{upper}} - E_{\text{lower}}$

$1 \text{ eV} = 1.60 \times 10^{-19} \text{ J}$ (the energy an electron gains through 1 V)



The energy levels of hydrogen. The energies are negative: the electron is trapped (bound). At 0 eV it is free.

For hydrogen these levels follow $E = -13.6/n^2 \text{ eV}$ (HL: the Bohr model explains why). This formula is **only for hydrogen**. Other atoms have their own levels, which a question gives you in a diagram or a table.

Learn it word for word.

Ground state: the lowest energy level of the atom.

Excited state: any energy level above the ground state.

Worked example 9

Find the wavelength of the photon emitted when the electron in a hydrogen atom falls from $n = 3$ (-1.51 eV) to $n = 2$ (-3.40 eV). What colour is it?

1. $\Delta E = -1.51 - (-3.40) = 1.89$ eV $= 1.89 \times 1.60 \times 10^{-19} = 3.02 \times 10^{-19}$ J.

2. $\lambda = hc/\Delta E = 6.63 \times 10^{-34} \times 3.00 \times 10^8 / 3.02 \times 10^{-19} = 6.58 \times 10^{-7}$ m (658 nm).

3. This is the red line of hydrogen in the spectrum above.

Spectra are the evidence for discrete energy levels

Each line has one wavelength, so one photon energy ($E = hc/\lambda$). If an atom could have any energy, it could emit photons of any energy, and we would see a continuous spectrum. Only certain lines appear, so only certain energy differences are possible: the energy levels are discrete.

Energy in eV is your friend

$hc = 6.63 \times 10^{-34} \times 3.00 \times 10^8 = 1.99 \times 10^{-25}$ J m. A quick check: λ (in nm) $\approx 1240/E$ (in eV). For 1.89 eV: $1240/1.89 \approx 656$ nm. Use it to check answers, but show the full method in the exam.

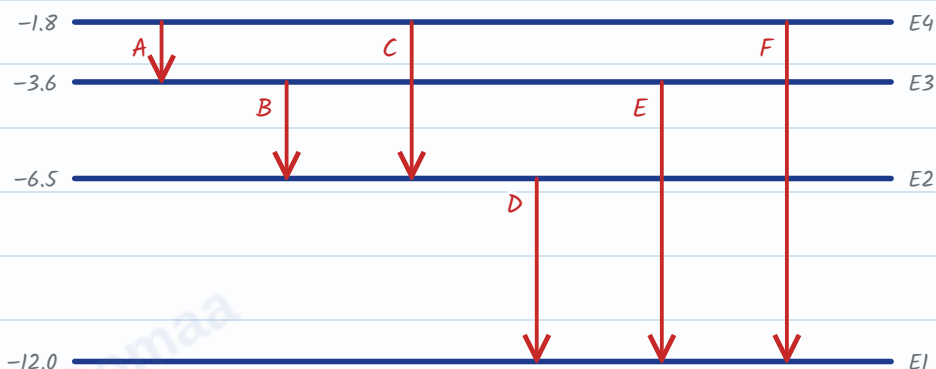
Reading an energy-level diagram

A favourite Paper 1A question gives a diagram of energy levels and some arrows, and asks which arrow gives the longest wavelength, the highest frequency, or the order of the wavelengths. One rule answers them all:

Bigger energy gap $\rightarrow$ bigger photon energy $\rightarrow$ higher frequency $\rightarrow$ SHORTER wavelength.

$E = hf$, so $f \propto \Delta E$. $E = hc/\lambda$, so $\lambda \propto 1/\Delta E$. The longest arrow gives the shortest wavelength.

E / eV



Six possible downward transitions between four levels. Measure each arrow: the longer it is, the shorter the wavelength.

Worked example 10

Use the energy levels in the diagram ($E1 = -12.0 \text{ eV}$, $E2 = -6.5 \text{ eV}$, $E3 = -3.6 \text{ eV}$, $E4 = -1.8 \text{ eV}$). (a) Which transition gives the longest wavelength, and which the shortest? (b) Put all six in order of wavelength. (c) Which lines are visible (400 nm to 700 nm)?

1. Gaps: A ($4 \rightarrow 3$) 1.8 eV · B ($3 \rightarrow 2$) 2.9 eV · C ($4 \rightarrow 2$) 4.7 eV · D ($2 \rightarrow 1$) 5.5 eV · E ($3 \rightarrow 1$) 8.4 eV · F ($4 \rightarrow 1$) 10.2 eV.

2. (a) Smallest gap A $\rightarrow$ **longest wavelength: A**. Largest gap F $\rightarrow$ **shortest wavelength: F**.

3. (b) $\lambda = hc/\Delta E$ gives: A 691 nm, B 429 nm, C 264 nm, D 226 nm, E 148 nm, F 122 nm. Order (longest first): **A, B, C, D, E, F**.

4. (c) Only **A (691 nm, red)** and **B (429 nm, violet)** are visible; the rest are ultraviolet.

Worked example 11

Photons emitted when an atom goes from level E_1 to the ground state E_0 have wavelength 600 nm, and from E_2 to E_0 they have wavelength 400 nm. Find the energy gap $E_2 - E_1$ in eV, and the wavelength of the photon emitted from E_2 to E_1 .

1. $E_1 - E_0 = hc/\lambda = 1.99 \times 10^{-25} / 600 \times 10^{-9} = 3.32 \times 10^{-19} \text{ J} = 2.07 \text{ eV}.$

2. $E_2 - E_0 = 1.99 \times 10^{-25} / 400 \times 10^{-9} = 4.97 \times 10^{-19} \text{ J} = 3.11 \text{ eV}.$

3. $E_2 - E_1 = 3.11 - 2.07 = 1.04 \text{ eV}.$ Then $\lambda = hc/\Delta E = 1200 \text{ nm}$ (infrared).

4. Energies add, wavelengths do not: $400 + 600$ is not the answer. (Check: $1/\lambda = 1/400 - 1/600$ gives 1200 nm.)

How many lines?

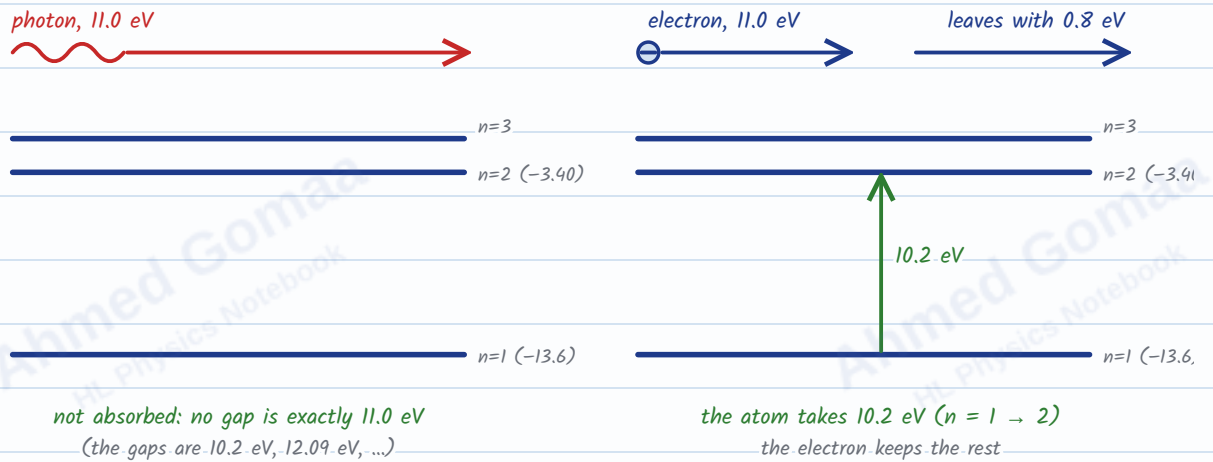
From level n (with all the levels below it) the electron can reach the ground state by many routes. The number of different lines possible is $n(n-1)/2$: 4 levels give $4 \times 3/2 = 6$ lines (as in the diagram); 5 levels give 10 lines.

✗ Mistake

Thinking the longest arrow gives the longest wavelength. It is the other way round: the longest arrow (biggest energy) gives the **shortest** wavelength. Also: subtracting wavelengths instead of energies.

Excitation, ionisation and the dark lines

An electron can move **up** to a higher level only if the atom gains **exactly** the right energy. There are two ways to give it that energy, and they behave differently.



A photon must be the exact energy of a gap. An electron can give part of its energy.

Absorbing a photon: the photon energy must **equal** an energy gap exactly; the whole photon is absorbed or not at all.

Hit by an electron: the atom can take **part** of the electron's kinetic energy (exactly one gap); the electron leaves with the rest.

Learn it word for word.

Ionisation energy: the minimum energy needed to remove an electron from an atom in its ground state (to $n = \infty$, energy 0). For hydrogen: 13.6 eV.

Worked example 12

Hydrogen atoms are in the ground state. What happens if (a) a photon of 11.0 eV, (b) an electron of 11.0 eV, (c) a photon of 14.0 eV meets an atom?

1. (a) The gaps from $n = 1$ are 10.2 eV, 12.09 eV, ...: none is 11.0 eV, so **the photon is not absorbed** (it passes through).
2. (b) The atom takes 10.2 eV ($n = 1 \rightarrow 2$); **the electron leaves with 0.8 eV**.
3. (c) 14.0 eV is more than the ionisation energy (13.6 eV): **the atom is ionised** and the freed electron leaves with $14.0 - 13.6 = 0.4$ eV of kinetic energy.

Worked example 13

Find the ionisation energy of a hydrogen atom in the state $n = 2$, and the minimum frequency of a photon that can ionise it.

1. From -3.40 eV to 0: **3.40 eV**.
2. $f = E/h = 3.40 \times 1.60 \times 10^{-19} / 6.63 \times 10^{-34} = \mathbf{8.2 \times 10^{14} \text{ Hz}}$.

Why are the absorption lines dark?

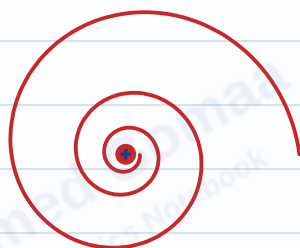
The absorbed photons excite the atoms, and the atoms quickly fall back and emit photons of the **same** wavelengths. So why are the lines missing? Because the re-emitted photons go off in **all directions**, not along the original beam. Very few of them reach the observer, so those wavelengths are much dimmer: dark lines.

"Explain, with reference to energy levels, the formation of an absorption spectrum" (3 marks)

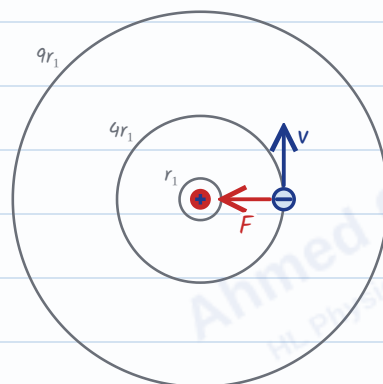
- ✓ White light (a continuous spectrum) passes through a **cool** gas.
- ✓ Photons with energy **equal to the difference between two energy levels** are absorbed, and electrons move to higher levels ($E = hf = \Delta E$).
- ✓ The photons are re-emitted **in all directions**, so these wavelengths are missing (dark lines) in the transmitted light.

HL The Bohr model of hydrogen

Rutherford's model has a serious problem. An electron in a circular orbit is **accelerating** (centripetal acceleration), and classical physics says that an accelerating charge **radiates** electromagnetic energy. It would lose energy, spiral into the nucleus in a tiny fraction of a second, and every atom would collapse. Atoms are obviously stable, so something is missing.



classical physics: an accelerating charge radiates, so it would spiral in



Bohr: only orbits with $mvr = nh/2\pi$; in these the electron does not radiate

Left: the classical prediction. Right: Bohr's allowed orbits, where the electron does not radiate.

Bohr's condition (data booklet): $mvr = nh/2\pi$, $n = 1, 2, 3, \dots$

mvr is the **angular momentum** of the electron. Only orbits in which the angular momentum is a whole-number multiple of $h/2\pi$ are allowed. In these orbits **the electron does not radiate**.

Discrete angular momentum $\rightarrow$ discrete orbits $\rightarrow$ discrete energies

Bohr's condition allows only certain values of mvr . Combined with the electric force that keeps the electron in orbit, this allows only certain **radii** ($r \propto n^2$). Each radius has its own energy, so only certain **energies** are allowed: $E = -13.6/n^2$ eV. The electron radiates only when it **jumps** from one allowed orbit to another, and the photon carries the energy difference. This explains the line spectrum of hydrogen exactly.

Show that $r \propto n^2$ and $E = -13.6/n^2$ eV

1. The electric force provides the centripetal force: $ke^2/r^2 = mv^2/r$, so $mv^2 = ke^2/r$. (1)
2. Bohr's condition: $mvr = nh/2\pi$, so $v = nh/(2\pi mr)$. (2)
3. Put (2) into (1): $m \times n^2 h^2 / (4\pi^2 m^2 r^2) = ke^2/r$. Cancel one r and one m : $r = n^2 \times h^2 / (4\pi^2 kme^2)$, so $r \propto n^2$.
4. For $n = 1$: $r_1 = h^2 / (4\pi^2 kme^2) = (6.63 \times 10^{-34})^2 / (4\pi^2 \times 8.99 \times 10^9 \times 9.11 \times 10^{-31} \times (1.60 \times 10^{-19})^2) = 5.3 \times 10^{-11}$ m (the Bohr radius).
5. Total energy = kinetic + electric potential: $E = \frac{1}{2}mv^2 - ke^2/r = \frac{1}{2}ke^2/r - ke^2/r$ (using (1)) = $-ke^2/(2r)$.
6. Put in $r = n^2 r_1$: $E = -ke^2/(2r_1) \times 1/n^2 = -13.6$ eV/ n^2 (with accurate constants). The energy is negative because the electron is bound.

Important: $E = -13.6/n^2 \text{ eV}$ is for **HYDROGEN** only

The number 13.6 eV comes from the charge of **one proton** and **one electron**, so this formula works **only for the hydrogen atom**.

- Other one-electron ions (He^+ , Li^{2+}) still have levels that go as $1/n^2$, but the number is different: the question gives it (for example $E = -K/n^2$, see worked example 15).
- For atoms with more electrons (helium, sodium, mercury ...) there is **no formula**: use the energy-level diagram or the values in the question.

Worked example 14

For the electron in a hydrogen atom in the state $n = 3$, find (a) the orbit radius, (b) its speed, (c) its energy. (Bohr radius $r_1 = 5.3 \times 10^{-11} \text{ m}$.)

1. (a) $r = n^2 r_1 = 9 \times 5.3 \times 10^{-11} = 4.8 \times 10^{-10} \text{ m}$.

2. (b) $v = nh/(2\pi mr) = 3 \times 6.63 \times 10^{-34} / (2\pi \times 9.11 \times 10^{-31} \times 4.77 \times 10^{-10}) = 7.3 \times 10^5 \text{ m s}^{-1}$ (one third of the speed in $n = 1$, $2.2 \times 10^6 \text{ m s}^{-1}$; $v \propto 1/n$).

3. (c) $E = -13.6/3^2 = -1.51 \text{ eV}$.

Worked example 15

In a hydrogen-like ion (one electron) the energy levels are $E = -K/n^2$. A photon of 40.8 eV is needed to move the electron from $n = 1$ to $n = 2$. Find K and the photon energy needed from $n = 2$ to $n = 3$.

1. $\Delta E(1 \rightarrow 2) = K(1 - 1/4) = 3K/4 = 40.8 \text{ eV}$, so $K = 54.4 \text{ eV}$ (this is helium with one electron removed, He^+).

2. $\Delta E(2 \rightarrow 3) = K(1/4 - 1/9) = 54.4 \times 5/36 = 7.56 \text{ eV}$.

Where the Bohr model works and where it fails

It works only for atoms with **one electron**: hydrogen, and ions such as He^+ and Li^{2+} (an atom with all but one electron removed). It cannot explain atoms with more electrons, the different brightness of the lines, or **why** the condition is true. Later quantum theory (the electron as a wave) explains it, but the energy levels it predicts for hydrogen are still correct.

"Explain why Bohr introduced his condition" (3 marks)

- ✓ Electrons in orbit **accelerate**, so (classically) they **radiate energy**.
- ✓ They would therefore **spiral into the nucleus**: atoms would be unstable.
- ✓ The only orbits where electrons do **not radiate** are those that satisfy $mvr = nh/2\pi$.
- ✓ So the electrons have **discrete energy levels**.

✗ Mistake

Writing the condition as $mvr = nh$ (forgetting 2π), or thinking the energy of a higher level is more negative. $E = -13.6/n^2$; $n = 1$ is the **lowest** (most negative) energy; the levels get closer together as n increases.

Writing the explanations

E.1 has short "describe", "outline" and "explain" questions. The markschemes look for the same key words every time.

"Describe the results of the Geiger-Marsden-Rutherford experiment" (2 marks)

- ✓ **Most** alpha particles passed through the foil **undeflected** (or with small deflections).
- ✓ A **very small number** were deflected through **large angles** (more than 90° , back towards the source).

"Outline the Rutherford model of the atom" (2 marks)

- ✓ (Most of) the **mass** of the atom is confined within a very small volume, the **nucleus**.
- ✓ (All) the **positive charge** is confined within this very small volume.
- ✓ Electrons orbit the nucleus (the atom is mostly empty space).

"Explain why the Thomson model cannot explain the large deflections"

- ✓ In the Thomson model the positive charge is **spread** through the whole atom.
- ✓ So the electric force on an alpha particle is always **small**: only small deflections.
- ✓ A large deflection needs a very large force: charge concentrated in a **tiny** nucleus ($F = kQq/r^2$ with r very small).

"Explain how line spectra give evidence for discrete energy levels" (2 marks)

- ✓ Only certain wavelengths (frequencies) are emitted or absorbed.
- ✓ Each corresponds to a photon of one energy ($E = hf$), equal to the difference between two levels, so only certain energy differences exist: the levels are **discrete**.

"Explain the formation of an absorption spectrum" (3 marks)

Continuous spectrum through a cool gas · photons with energy equal to a difference in energy levels absorbed, electrons excited · re-emitted in all directions → dark lines.

HL: "Explain why Bohr introduced his condition" (3 marks)

accelerating electrons radiate · they would spiral into the nucleus · only orbits with $mvr = nh/2\pi$ do not radiate · discrete energy levels.

HL: deviations from Rutherford scattering

Alpha particle within the range of the strong nuclear force · nuclear radius $\approx$ distance of closest approach at E_0 · found from energy conservation, $E_0 = kQq/r$.

Common mistakes

✗ Ten traps in E.1

1. "Most alpha particles were deflected." **Most went straight through**, only a very small number were deflected through large angles.
2. Saying the alpha particles hit or bounced off the nucleus. They are **repelled** by the electric force.
3. Mixing up Z and A : the proton number Z gives the charge; neutrons = $A - Z$.
4. HL: forgetting that the alpha particle's charge is $2e$, or leaving the energy in MeV in $d = kq_1q_2/E_k$.
5. HL: $R \propto A$ instead of $R \propto A^{1/3}$, or thinking a big nucleus is denser (all nuclei have the same density).
6. HL: explaining deviations from Rutherford scattering without the **strong nuclear force**.
7. Longest arrow = longest wavelength. No: biggest energy gap = **shortest** wavelength.
8. Subtracting wavelengths instead of energies ($E_2 - E_1 = hc/\lambda_{21}$, not $\lambda_2 - \lambda_1$).
9. A photon of "nearly the right" energy is absorbed. A photon must match a gap **exactly**; an electron can give part of its energy.
10. Dark lines "because the atoms keep the energy". The photons are **re-emitted in all directions**. And: using $E = -13.6/n^2$ eV for an atom that is **not hydrogen**; HL: $mvr = nh$ instead of $nh/2\pi$.

Practise next

You have read the notes. Now practise, in this order:

1. **Levelled practice**: first a ★ at Level 2 in every topic (the goal), then Levels 3 and 4 for a 6 or 7. Hints and worked answers help you as you go.
2. **IB-format practice** ([Paper 1A](#) · [Paper 2](#)): timed, one attempt, and your report goes to your teacher. Do these when the practice goal is reached.

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