



CrackingCHEM

GCSE | A-LEVEL

UNDERSTAND • APPLY • SUCCEED

OCR A H432

Atomic Structure I

Specification-led notes



SPEC →



CRACK IT →



TEST IT →



DONE

EVERY SPECIFICATION POINT. EXPLAINED. PRACTISED. CHECKED OFF.



SPECIFICATION COVERAGE

2.1.1(a)		Isotopes: same element, different neutrons and masses
2.1.1(b)		Particle numbers in atoms and ions; changing atomic models
2.1.1(c)		Relative isotopic mass and relative atomic mass definitions
2.1.1(d)		Using mass spectra and calculating relative atomic mass
2.1.1(e)		Relative molecular mass and relative formula mass

Written by an experienced A-level Chemistry examiner and teacher

Mapped to Cambridge OCR Chemistry A H432, Version 3.1 (2026)



The big picture

THE PROMISE This booklet turns every OCR A Atomic Structure I requirement into something you can explain, calculate or interpret in an exam.

SPEC	CRACK IT	TEST IT	DONE
This is your target	Unlock what it actually means	Prove you can do it	Mastered

OCR A BOUNDARY Electron structure and orbitals in detail belong to 2.2.1, while ionisation energies belong to 3.1.1. OCR explicitly says knowledge of the mass spectrometer is not required here: learn to use mass-spectrum data, not TOF stages or TOF calculations. This topic is limited to singly charged ions.

1. How the atomic model changed



SPEC 2.1.1(b) + HSW1/7

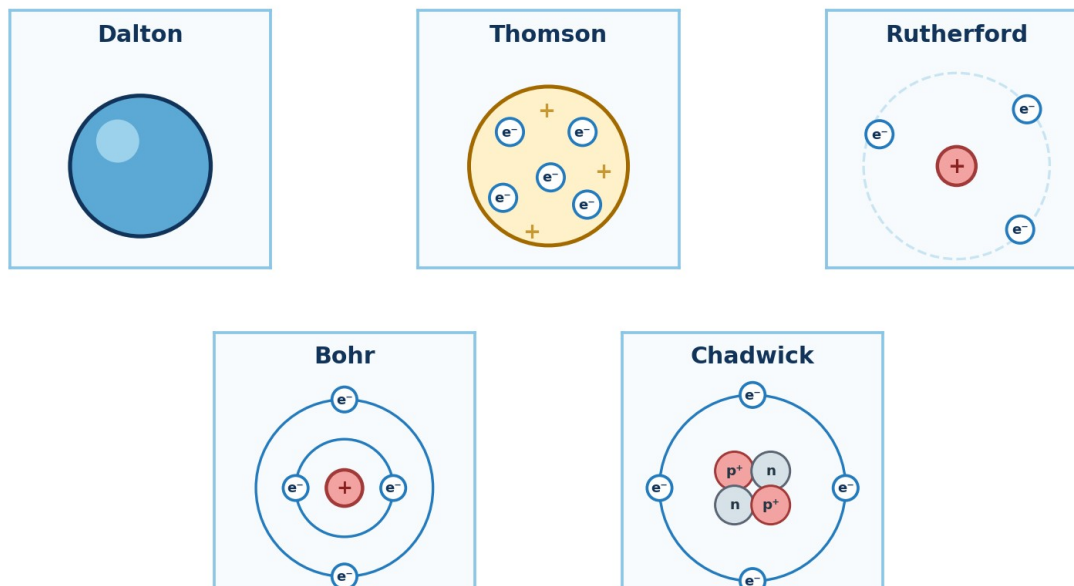
Use different models to explain atomic structure and understand how evidence caused accepted models to change. OCR highlights the Bohr model as an explanation for periodic properties.



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A scientific model changes when new evidence cannot be explained by the old model. In an exam, link each observation to the conclusion it supports.

The atomic model changed when new evidence appeared



The atomic model changed when new evidence appeared.

Dalton Atoms were pictured as tiny, solid, indivisible spheres.	Thomson The electron showed that atoms were divisible. Electrons were placed in diffuse positive charge.
Rutherford Alpha scattering showed a tiny, dense, positive nucleus and that most of the atom is empty space.	Bohr → Chadwick Bohr placed electrons in fixed energy levels, helping explain periodic properties. Chadwick's neutron completed the nuclear model and helped explain isotope masses.

EXAM TRAP Do not merely list the models. State the evidence, then state what that evidence showed.

WORKED EXAMPLE Alpha scattering: observation → inference

QUESTION During Rutherford's alpha-scattering experiment, most alpha particles passed straight through, some were deflected through large angles and a very small number rebounded. State what these observations showed about the atom.

ANSWER

1. Most alpha particles passed straight through → most of the atom is empty space.
2. Some were deflected through large angles and a very small number rebounded → positive charge and most mass are concentrated in a tiny, dense nucleus that is much smaller than the atom.



TEST IT 1 · QUICK CHECK

Explain how two observations from alpha scattering changed the model of the atom. [4]



DONE WHEN: You can connect Dalton → Thomson → Rutherford → Bohr → Chadwick using evidence, not just names.

2. Fundamental particles and the atom



SPEC 2.1.1(b)

Describe atomic structure in terms of the numbers of protons, neutrons and electrons in atoms and ions, given atomic number, mass number and any ionic charge.



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Particle	Relative charge	Relative mass	Location
Proton	+1	1	Nucleus
Neutron	0	1	Nucleus
Electron	-1	approximately 1/1840	Energy levels / orbitals outside the nucleus

- Almost all atomic mass is concentrated in the nucleus.
- The nucleus is tiny compared with the whole atom.
- A neutral atom has equal numbers of protons and electrons.

MEMORY HOOK Protons are positive. Neutrons are neutral. Electrons are negative and have negligible relative mass.



TEST IT 2 · QUICK CHECK

State the relative charge, relative mass and location of a proton, neutron and electron. [9]



DONE WHEN: You can reproduce the particle table from memory and describe the structure of an atom.

3. Particle numbers in atoms and ions



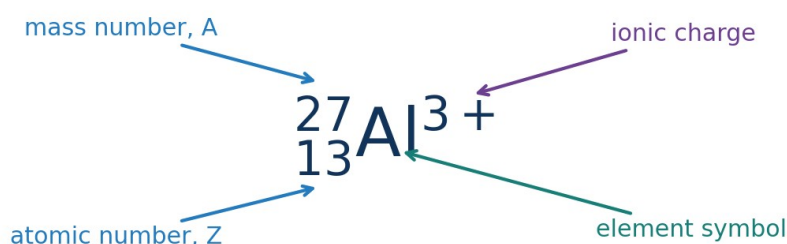
SPEC 2.1.1(b)

Determine the numbers of protons, neutrons and electrons in atoms and ions from atomic number, mass number and ionic charge.



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Atomic number, Z	The number of protons in the nucleus. It identifies the element.
Mass number, A	The total number of protons and neutrons in one atom of a particular isotope. It is a whole number.



Reading nuclide notation: mass number, atomic number, symbol and ionic charge.

$$A = Z + \text{number of neutrons}$$

A and Z are whole numbers and have no units. Therefore: neutrons = $A - Z$.

CALCULATING NUMBERS OF SUBATOMIC PARTICLES IN IONS

1. First find the neutral atom: protons = Z; electrons = Z; neutrons = $A - Z$.
2. Positive ion: electrons have been lost, so subtract the charge from Z.
3. Negative ion: electrons have been gained, so add the size of the charge to Z.
4. Only the electron number changes when an ion forms.

WORKED EXAMPLE Finding particle numbers in an ion

QUESTION State the number of protons, neutrons and electrons in $^{56}\text{Fe}^{2+}$. The atomic number of iron is 26.

1. Iron has $Z = 26$, so it contains 26 protons.
2. Neutrons = $56 - 26 = 30$.
3. Start with 26 electrons, then subtract 2 because the ion is $2+$.

ANSWER 26 protons, 30 neutrons and 24 electrons.

EXAM TRAP Work out the neutral atom first. Then subtract electrons for a positive ion or add electrons for a negative ion. This is easy to reverse under pressure.



TEST IT 3 · QUICK CHECK

State the numbers of protons, neutrons and electrons in $^{27}\text{Al}^{3+}$. The atomic number of aluminium is 13. [3]



DONE WHEN: You can move both ways: notation → particle numbers and particle numbers → notation.

4. Isotopes



SPEC 2.1.1(a)

Isotopes are atoms of the same element with different numbers of neutrons and different masses.



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Isotopes

Atoms of the same element with the same number of protons but different numbers of neutrons and different masses.

- Same element means the same atomic number and therefore the same number of protons.
- Different neutron numbers give different mass numbers.
- Neutral atoms of the isotopes have the same number of electrons and the same electron configuration, so they react in the same way.
- Their masses differ, so mass-dependent physical properties can differ.

WORKED EXAMPLE Comparing isotopes

QUESTION Compare neutral atoms of ^{35}Cl and ^{37}Cl . State their numbers of protons, neutrons and electrons, then explain why they have the same chemical properties but different masses. The atomic number of chlorine is 17.

1. Both atoms contain 17 protons and 17 electrons.
2. ^{35}Cl contains 18 neutrons; ^{37}Cl contains 20 neutrons.
3. They have the same electron configuration and therefore react in the same way; their different neutron numbers give different masses.

ANSWER Both contain 17 protons and 17 electrons; ^{35}Cl has 18 neutrons and ^{37}Cl has 20.

DO NOT MIX THESE UP Mass number describes one isotope and is a whole number. Relative atomic mass is a weighted mean for the element and is often not a whole number.



TEST IT 4 · QUICK CHECK

Explain why ^{35}Cl and ^{37}Cl have the same chemical properties but different masses. [3]



DONE WHEN: You can define isotopes using OCR's full wording and explain both their similarities and differences.

5. The carbon-12 relative mass scale



SPEC 2.1.1(c)

Explain relative isotopic mass and relative atomic mass using one-twelfth of the mass of a carbon-12 atom as the standard. Definitions are required.



CRACK IT

Relative isotopic mass

The mass of an atom of one isotope compared with one-twelfth of the mass of a carbon-12 atom. No units.

Relative atomic mass, A_r

The weighted mean mass of an atom of an element compared with one-twelfth of the mass of a carbon-12 atom. No units.

Mass number, A Counts protons + neutrons in one atom. It is always a whole number.	Relative isotopic mass A measured relative mass for one isotope. It is usually close to, but not necessarily exactly equal to, the mass number.
Relative atomic mass, A_r A weighted mean across the naturally occurring isotopes of an element. It is often not a whole number.	Units Relative masses are ratios, so they have no units.

DEFINITION PRECISION For A_r , include weighted mean, atom of an element and the carbon-12 comparison. Missing any one of these can cost a definition mark.



TEST IT 5 · QUICK CHECK

Define relative isotopic mass and relative atomic mass, then state one difference between mass number and relative isotopic mass. [5]



DONE WHEN: You can state both required definitions exactly and distinguish mass number from relative mass.

6. Reading a mass spectrum



SPEC 2.1.1(d)(i)

Use mass spectrometry to determine relative isotopic masses and relative abundances of isotopes. Knowledge of the mass spectrometer is not required; ions are limited to single charges.

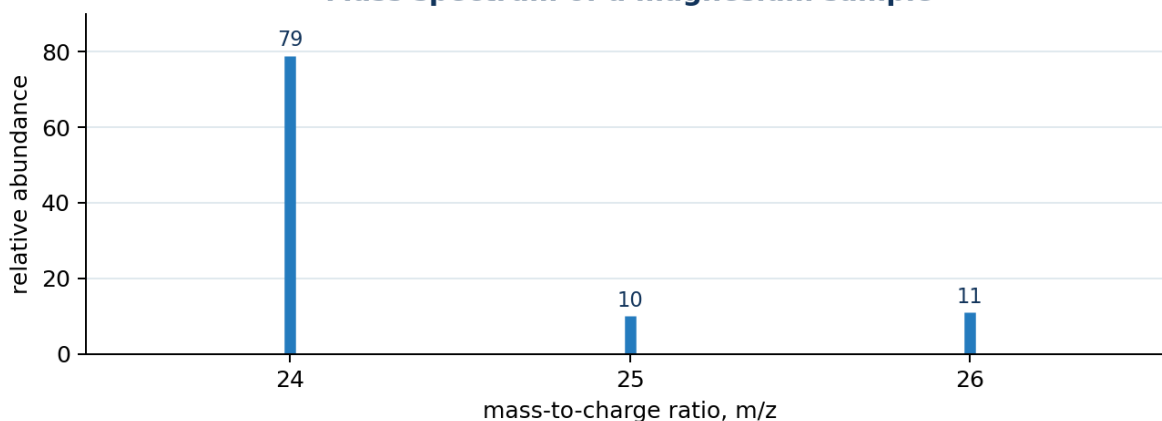


CRACK IT

OCR A FOCUS Do not waste revision time learning TOF stages or TOF equations for this specification point. OCR tests what the spectrum tells you. The single-charge limit keeps m/z interpretation straightforward.

- x-axis: mass-to-charge ratio, m/z .
- y-axis: relative abundance.
- Peak position gives the m/z of an isotope ion.
- Peak height or intensity gives its relative abundance.
- For a singly charged monatomic ion, m/z is numerically equal to its relative isotopic mass.

Mass spectrum of a magnesium sample



Mass spectrum of magnesium: peak position gives m/z ; peak height gives relative abundance.

EXAM TRAP Do not confuse peak position with peak height. Position tells you relative isotopic mass; height tells you relative abundance.

**TEST IT 6 · QUICK CHECK**

A mass spectrum has peaks at m/z 24, 25 and 26 with relative abundances 79, 10 and 11. State the relative isotopic masses and abundances of the isotopes. [3]



DONE WHEN: You can read both axes and extract the isotope masses and abundances without describing the instrument.

7. Calculating relative atomic mass



SPEC 2.1.1(d)(ii) + M0.2/M1.2

Calculate the relative atomic mass of an element from the relative abundances of its isotopes, including use of ratios, percentages and weighted means.

**CRACK IT**

$$A_r = \Sigma(\text{relative isotopic mass} \times \text{abundance}) / \Sigma(\text{abundances})$$

A_r has no units. If abundances are percentages, the denominator is 100.

WORKED EXAMPLE Type 1: every mass and abundance is known

QUESTION A magnesium sample contains isotope ions at m/z 24, 25 and 26 with relative abundances 79%, 10% and 11%. Calculate the relative atomic mass, A_r , of magnesium.

1. Multiply each relative isotopic mass by its abundance.
2. Add the products and divide by the total abundance.
3. $A_r = [24(79) + 25(10) + 26(11)]/100$.

ANSWER $A_r = 24.32$

WORKED EXAMPLE Type 2: one isotope mass is unknown

QUESTION An element has two isotopes. One has relative isotopic mass 63.0 and abundance 69.0%; the other has abundance 31.0%. The element has $A_r = 63.62$. Calculate the relative isotopic mass of the second isotope.

1. Write $63.62 = [63.0(69.0) + 31.0x]/100$.
2. Multiply by 100: $6362 = 4347 + 31.0x$.
3. Rearrange: $x = 65.0$.

ANSWER The second isotope has relative isotopic mass 65.0.

WORKED EXAMPLE Type 3: two abundances are unknown

QUESTION An element has isotopes of relative isotopic mass 24.0, 25.0 and 26.0. The 24.0 isotope has an abundance of 79.0% and the element has $A_r = 24.32$. Calculate the abundances of the 25.0 and 26.0 isotopes.

1. Let the 25.0 isotope have abundance $x\%$. The 26.0 isotope must have abundance $100 - 79 - x = 21 - x$.
2. $24.32 = [24(79) + 25x + 26(21 - x)]/100$.
3. $2432 = 2442 - x$, so $x = 10$.

ANSWER 25.0 isotope = 10.0%; 26.0 isotope = 11.0%.

SENSE CHECK A weighted mean must lie between the lightest and heaviest isotope masses and should sit closest to the most abundant isotope.

EXAM TRAP For three isotopes, the final abundance is $100 - \text{known abundance} - x$. Do not use $100 - x$ because the known abundance has already been counted.

**TEST IT 7 · QUICK CHECK**

Calculate A_r for isotopes of relative isotopic mass 35.0 and 37.0 with relative abundances 75.0% and 25.0%. [3]

**TEST IT 8 · QUICK CHECK**

An element has isotopes of relative isotopic mass 28.0, 29.0 and 30.0. The 28.0 isotope is 92.0% abundant and $A_r = 28.11$. Calculate the other two abundances. [5]



DONE WHEN: You can solve known-mass, unknown-mass and unknown-abundance weighted-mean problems and sense-check the result.

8. Relative molecular mass and relative formula mass

**SPEC 2.1.1(e)**

Use the terms relative molecular mass, M_r , and relative formula mass, and calculate them from relative atomic masses. Definitions are not required.

**CRACK IT****Simple molecules**

Use relative molecular mass, M_r . Add the A_r values for every atom in one molecule.

Examples of molecules

H_2O , CO_2 , CH_4 and NH_3 consist of separate molecules.

Compounds with giant structures

Use relative formula mass. Add the A_r values for every atom shown in the empirical formula.

Examples of giant compounds

$NaCl$ and $MgCl_2$ are giant ionic; SiO_2 is giant covalent.

relative molecular mass or relative formula mass = $\Sigma(A_r \times \text{number of each atom})$

Use the OCR term that matches the structure. Relative masses have no units.

WORKED EXAMPLE Choosing the correct term

QUESTION Calculate the relative mass of CO_2 and $MgCl_2$ using A_r values $C = 12.0$, $O = 16.0$, $Mg = 24.3$ and $Cl = 35.5$. State the correct OCR term for each substance and whether either answer has units.

- CO_2 consists of simple molecules: $M_r = 12.0 + 2(16.0) = 44.0$.
- $MgCl_2$ has a giant ionic structure: relative formula mass = $24.3 + 2(35.5) = 95.3$.
- Neither relative mass has units.

ANSWER $M_r(CO_2) = 44.0$; relative formula mass($MgCl_2$) = 95.3.

OCR WORDING OCR says definitions of relative molecular mass and relative formula mass are not required. Spend your effort choosing the correct term, using the formula correctly and leaving off units.

**TEST IT 9 · QUICK CHECK**

Calculate M_r for CH_4 and the relative formula mass of $MgCl_2$ using $H = 1.0$, $C = 12.0$, $Mg = 24.3$ and $Cl = 35.5$. Explain why different terms are used. [5]



DONE WHEN: You can choose the OCR term from the structure, calculate the relative mass and avoid adding units.



Final mastery check

THE RULE Tick a statement only when you can do it without looking back at the notes.

☐	Explain how evidence changed the accepted atomic model.
☐	State the relative charges, relative masses and locations of protons, neutrons and electrons.
☐	Use A, Z and charge to calculate particle numbers in atoms and ions.
☐	Define isotopes using OCR's full wording and explain why isotope chemistry is the same.
☐	Define relative isotopic mass and A_r precisely using the carbon-12 standard.
☐	Distinguish mass number, relative isotopic mass and relative atomic mass.
☐	Read relative isotopic masses and relative abundances from a mass spectrum.
☐	Explain the OCR boundary: single charges; no knowledge of the mass spectrometer required.
☐	Solve all three weighted-mean A_r calculation types.
☐	Choose M_r or relative formula mass and calculate it without units.



TEST IT answers

USE THESE PROPERLY Attempt each TEST IT question first. Then compare your answer with the chemistry below and return to the exact specification point if anything is missing.

TEST IT	CHECK ANSWER
1	Most alpha particles passed straight through, showing that most of the atom is empty space. Some were deflected through large angles and a very small number rebounded, showing that positive charge and most mass are concentrated in a tiny, dense nucleus.
2	Proton: charge +1, relative mass 1, nucleus. Neutron: charge 0, relative mass 1, nucleus. Electron: charge -1, relative mass approximately 1/1840, in energy levels / orbitals outside the nucleus.
3	Aluminium has 13 protons. Neutrons = $27 - 13 = 14$. A 3+ ion has lost three electrons, so electrons = $13 - 3 = 10$.
4	Both atoms contain 17 protons and 17 electrons. Their identical electron configurations mean they react in the same way. ^{35}Cl has 18 neutrons and ^{37}Cl has 20, so their masses differ.
5	Relative isotopic mass is the mass of an atom of one isotope compared with one-twelfth of the mass of a carbon-12 atom. Relative atomic mass is the weighted mean mass of an atom of an element on the same scale. Neither has units. Mass number is the whole-number total of protons and neutrons in one atom.
6	The isotope ions have m/z values 24, 25 and 26 and relative abundances 79, 10 and 11. OCR limits this requirement to singly charged ions, so each m/z value is numerically equal to the relative isotopic mass.
7	$A_r = [35.0(75.0) + 37.0(25.0)]/100 = 35.5$. The weighted mean lies between 35.0 and 37.0 and closest to 35.0, the most abundant isotope.
8	Let the abundance of relative isotopic mass 29.0 be x%. The abundance of relative isotopic mass 30.0 is $(8 - x)\%$. Solving $28.11 = [28.0(92) + 29.0x + 30.0(8 - x)]/100$ gives $29.0 = 5.0\%$ and $30.0 = 3.0\%$.
9	$M_r(\text{CH}_4) = 12.0 + 4(1.0) = 16.0$. Relative formula mass(MgCl_2) = $24.3 + 2(35.5) = 95.3$. CH_4 is a simple molecule; MgCl_2 has a giant structure. Relative masses have no units.