



Premium IB tuition — Physics, Chemistry, Maths

IB CHEMISTRY · STRUCTURE 1.4.1 – 1.4.4

Counting Particles by Mass: The Mole Concept

Notes · Formulae · Solved Examples · Practice & Mark Schemes

STANDARD LEVEL

HIGHER LEVEL

HL STRETCH INCLUDED

First assessment 2025 syllabus | Paper 1A (MCQ), Paper 1B (data-based) and Paper 2 practice
20 multiple-choice questions · 15 fully solved examples · 5 exam-style questions with full mark
schemes

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AN IMPORTANT NOTE ON SL AND HL FOR THIS SUB-TOPIC

In the current IB Chemistry syllabus (first assessment 2025), **the whole of Structure 1.4 is common core content** — it is studied and examined at both Standard Level and Higher Level. There is **no AHL (HL-only) material in 1.4.1–1.4.4**. Anyone who tells you otherwise is working from the old (2016) syllabus.

So that this booklet is still genuinely useful to HL students, every section is labelled as follows:

LABEL	WHAT IT MEANS
SL + HL	Examinable syllabus content for both levels. This is the great majority of the booklet — all notes, all formulae, all four sub-topics.
HL STRETCH	Not extra syllabus content. These are harder question styles, multi-step problems and forward links to AHL topics where HL candidates are expected to apply mole calculations with greater sophistication. SL students may safely skip these, but strong SL candidates will benefit from attempting them.

HOW TO GET THE MOST OUT OF THIS BOOKLET

- ▶ **Read Part A actively.** Cover the worked lines with a card and try each one before reading on.
- ▶ **Memorise Part B.** Only the Avogadro constant and the A_r values are given in the IB data booklet — the relationships themselves are not.
- ▶ **Do Part E under timed conditions.** Allow about 1.5 minutes per MCQ and roughly 1.5 minutes per mark for Papers 1B and 2. Under the 2025 syllabus a calculator *and* the data booklet are permitted in every paper, including Paper 1A — so the challenge is speed and accuracy, not mental arithmetic.
- ▶ **Mark yourself honestly against Part F,** then re-do every question you lost a mark on *from scratch* two days later.

Data used throughout: A_r values are quoted to 2 decimal places as given in the IB Chemistry data booklet periodic table; the Avogadro constant is taken as $N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$. Final answers are given to 3 significant figures unless stated otherwise.

PART A — NOTES & KEY POINTS

STRUCTURE 1.4.1

The mole and the Avogadro constant

SL + HL

Syllabus statement: the mole (mol) is the SI unit of amount of substance (n). One mole contains exactly the number of elementary entities given by the Avogadro constant.

Why chemists need the mole

Atoms and molecules are far too small and far too numerous to count individually. A chemist can, however, **weigh** a sample very accurately. The mole is the bridge: it is a counting unit that lets us convert a mass we *can* measure into a number of particles we *cannot*. In exactly the same way that "a dozen" always means 12 regardless of what is being counted, "a mole" always means 6.02×10^{23} regardless of what is being counted.

DEFINITIONS YOU MUST KNOW

Amount of substance (n) — the physical quantity that measures the number of specified elementary entities in a sample. Its SI unit is the **mole**, symbol **mol**.

The mole — the amount of substance that contains exactly $6.022\,140\,76 \times 10^{23}$ elementary entities.

The Avogadro constant (N_A , sometimes written L) — the number of elementary entities per mole: $N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$. Note the units — it is a *constant*, not a pure number.

THE RELATIONSHIP

$$N = n \times N_A \quad \Leftrightarrow \quad n = N / N_A$$

N = number of particles (a pure number, no units) · n = amount in mol · $N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$

"Elementary entities" — you must always say *what you are counting*

A mole is meaningless until you state the entity. The entity may be an **atom**, a **molecule**, an **ion**, a **formula unit**, an **electron**, or even a specified group of particles. "One mole of oxygen" is ambiguous and would not be credited: one mole of oxygen *atoms* (O) has half the mass of one mole of oxygen *molecules* (O₂).

SAMPLE	ENTITIES PRESENT IN 1 MOL	NUMBER OF ATOMS
He	1 mol He atoms	6.02×10^{23}
O ₂	1 mol O ₂ molecules = 2 mol O atoms	1.20×10^{24}
H ₂ O	1 mol H ₂ O molecules = 2 mol H + 1 mol O	1.81×10^{24}
NaCl	1 mol formula units = 1 mol Na ⁺ + 1 mol Cl ⁻ (2 mol ions)	1.20×10^{24}
MgCl ₂	1 mol formula units = 1 mol Mg ²⁺ + 2 mol Cl ⁻ (3 mol ions)	1.81×10^{24}
Al ₂ (SO ₄) ₃	1 mol formula units = 2 mol Al ³⁺ + 3 mol SO ₄ ²⁻ (5 mol ions)	1.02×10^{25} (17 atoms per unit)

THREE TRAPS THAT COST MARKS EVERY SINGLE SESSION

1. **Molecules ≠ atoms.** 0.500 mol of CO₂ contains 0.500 mol of molecules but 1.50 mol of atoms (0.500 mol C + 1.00 mol O).
2. **Ionic compounds contain no molecules.** Never write "1 mol of NaCl molecules". Use *formula units*. This is a common Paper 2 penalty.
3. **Ions ≠ formula units.** "How many ions in 0.10 mol of Fe₂(SO₄)₃?" needs the factor 5, not 1.

NATURE OF SCIENCE — THE 2019 REDEFINITION

Until May 2019 the mole was *defined* as the number of atoms in exactly 12 g of carbon-12, and N_A was a measured quantity with an experimental uncertainty. Since the redefinition of the SI base units, the position is reversed: N_A is now a **fixed, exact, defined number** ($6.022\,140\,76 \times 10^{23} \text{ mol}^{-1}$) and the mass of 12 g of carbon-12 is the quantity carrying a tiny experimental uncertainty. This is a good illustration for TOK and for "improvements in instrumentation drive changes in scientific definitions". For all IB calculations the difference is far below 3 s.f. and can be ignored.

HL STRETCH WHERE THIS REAPPEARS AT HL

The entity-counting skill is assumed — not re-taught — in AHL work: counting **moles of electrons** in electrolysis calculations (Reactivity 3.2, where $Q = I t$ and one mole of electrons carries 96 500 C), counting **moles of photons** when relating energy to wavelength, and counting **moles of hydrogen ions** in pH and buffer work (Reactivity 3.1). If you are hazy on "which entity am I counting?" now, it will cost you marks two years from now.

STRUCTURE 1.4.2

Relative atomic mass and relative formula mass

SL + HL

Syllabus statement: masses of atoms are compared on a scale relative to ¹²C and are expressed as relative atomic mass (A_r) and relative formula mass (M_r).

The carbon-12 scale

The actual mass of a single atom is inconveniently small (of the order of 10^{-23} g). Chemists therefore compare atomic masses to a single agreed reference: one atom of carbon-12, which is assigned a mass of exactly 12 units. Everything is quoted as a **ratio** against one-twelfth of that mass.

DEFINITIONS YOU MUST KNOW

Relative atomic mass (A_r) — the weighted mean mass of the atoms in a naturally occurring sample of the element, relative to $1/12$ of the mass of an atom of carbon-12.

Relative formula mass (M_r) — the sum of the relative atomic masses of all the atoms shown in the formula of the substance. For substances made of discrete molecules this is also called the *relative molecular mass*.

Both are ratios, so both are dimensionless — they have NO UNITS. Writing " $A_r(\text{Cl}) = 35.45 \text{ g}$ " or " $M_r = 18.02 \text{ g mol}^{-1}$ " loses the mark.

Why A_r values are not whole numbers

Most elements exist naturally as a mixture of **isotopes** — atoms of the same element with different numbers of neutrons and therefore different masses (Structure 1.2). A_r is the *weighted average* of the isotopic masses, weighted by natural abundance. Chlorine's A_r of 35.45 does not mean any chlorine atom weighs 35.45 units; it means a natural sample behaves, on average, as if each atom did.

CALCULATING A_r FROM ISOTOPIC ABUNDANCES

$$A_r = \Sigma (\text{isotopic mass} \times \text{fractional abundance})$$

Equivalently, using percentage abundances: $A_r = \frac{(\text{mass}_1 \times \%_1) + (\text{mass}_2 \times \%_2) + \dots}{100}$

Building up M_r — the mechanical part

Add one A_r for every atom in the formula. The only skill being tested is careful reading of the formula, so slow down on brackets and dots.

FORMULA	ATOM COUNT	M_r
H_2O	$2 \text{ H} + 1 \text{ O} = 2(1.01) + 16.00$	18.02
$\text{Ca}(\text{NO}_3)_2$	$1 \text{ Ca} + 2 \text{ N} + 6 \text{ O} = 40.08 + 2(14.01) + 6(16.00)$	164.10
$(\text{NH}_4)_2\text{SO}_4$	$2 \text{ N} + 8 \text{ H} + 1 \text{ S} + 4 \text{ O}$	132.17
$\text{Al}_2(\text{SO}_4)_3$	$2 \text{ Al} + 3 \text{ S} + 12 \text{ O}$	342.17
$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	$1 \text{ Cu} + 1 \text{ S} + 4 \text{ O} + 5(2 \text{ H} + 1 \text{ O}) = 159.62 + 5(18.02)$	249.72

THE BRACKET-AND-DOT CHECKLIST

- ▶ A subscript outside a bracket multiplies *everything* inside it. In $\text{Ca}(\text{NO}_3)_2$ there are 6 oxygens, not 3.
- ▶ The dot in a hydrate such as $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ means "plus" — the water **is** part of the formula and **is** included in M_r .
- ▶ Use the A_r from the periodic table, **not** the mass number. $A_r(\text{Cl}) = 35.45$, not 35.
- ▶ Do not round part-way through. Carry full precision and round only the final answer.
- ▶ For ionic compounds, the mass of the transferred electrons is negligible, so $A_r(\text{Na}^+) \approx A_r(\text{Na})$. Use ordinary A_r values for ions.

HL STRETCH REVERSE ABUNDANCE PROBLEMS

A favourite harder question gives you A_r and the two isotopic masses and asks for the abundances. Let the abundance of the lighter isotope be x (as a fraction); the heavier is then $(1 - x)$. Solve the single linear equation

$$\text{mass}_{\text{light}} \cdot x + \text{mass}_{\text{heavy}} \cdot (1 - x) = A_r$$

This is fully within SL scope too, but appears more often on HL papers and in multi-part data questions. See Solved Example 5.

STRUCTURE 1.4.3

Molar mass, M

SL + HL

Syllabus statement: molar mass (M) has the units g mol^{-1} .

DEFINITION

Molar mass (M) — the mass of one mole of a substance. Units: g mol^{-1} . Numerically, M is equal to A_r (for an element made of single atoms) or to M_r (for a compound), but M carries units and A_r/M_r do not.

QUANTITY	SYMBOL	UNITS	EXAMPLE: WATER
Relative formula mass	M_r	none	$M_r(\text{H}_2\text{O}) = 18.02$
Molar mass	M	g mol^{-1}	$M(\text{H}_2\text{O}) = 18.02 \text{ g mol}^{-1}$
Mass of sample	m	g	a 9.01 g sample
Amount of substance	n	mol	$n = 0.500 \text{ mol}$

THE CENTRAL RELATIONSHIP

$$n = \frac{m}{M} \quad m = n M \quad M = \frac{m}{n}$$

Chain this with 1.4.1 to move all the way from mass to particle count: $N = (m / M) \times N_A$

The mass of a single particle

Because M is the mass of N_A particles, the mass of one particle follows immediately:

$$\text{mass of one particle} = M / N_A \quad (\text{in g})$$

For magnesium, $M = 24.31 \text{ g mol}^{-1}$, so one atom has mass $24.31 / (6.02 \times 10^{23}) = 4.04 \times 10^{-23} \text{ g}$. A sensible sanity check: any single-atom answer *must* come out somewhere around 10^{-22} to 10^{-25} g . If your answer is 10^{+23} , you have divided the wrong way round.

Percentage composition by mass

This asks what fraction of a compound's mass is contributed by one particular element.

$$\% \text{ by mass of X} = \frac{(\text{number of X atoms in formula}) \times A_r(\text{X})}{M_r \text{ of the compound}} \times 100$$

The same expression handles the **percentage of water of crystallisation** in a hydrate — simply put the total mass of the water into the numerator:

$$\% \text{ H}_2\text{O in MgSO}_4 \cdot 7\text{H}_2\text{O} = (7 \times 18.02) / 246.52 \times 100 = 126.14 / 246.52 \times 100 = \mathbf{51.2 \%}$$

UNITS AND MAGNITUDE — THE SILENT MARK-KILLERS

- ▶ **Always convert to grams first.** $1 \text{ kg} = 10^3 \text{ g}$; $1 \text{ tonne} = 10^6 \text{ g}$; $1 \text{ mg} = 10^{-3} \text{ g}$; $1 \text{ }\mu\text{g} = 10^{-6} \text{ g}$. A mass in kg divided by M in g mol^{-1} gives an answer $1000\times$ too big.
- ▶ **Molar mass is never in kg mol^{-1} in IB Chemistry**, even though that is the strict SI form.
- ▶ **Sanity-check the size.** If $m < M$ then n must be less than 1. A 5 g sample of a compound of $M = 250 \text{ g mol}^{-1}$ cannot possibly be 50 mol.
- ▶ **Give the unit with the answer.** "0.0200" is not the same answer as "0.0200 mol"; on Paper 2 the unit is often the mark.

HL STRETCH WHERE MOLAR MASS IS DOING THE HEAVY LIFTING LATER

Molar mass is the entry point to almost every quantitative topic that follows: $c = n/V$ for solutions (1.4.5), $pV = nRT$ for gases (1.4.6), limiting reagent, theoretical yield and atom economy (Reactivity 2.1), $q = mc\Delta T$ converted to ΔH in kJ mol^{-1} (Reactivity 1.2), and titration analysis. A very common HL Paper 2 structure is: mass $\rightarrow$ moles $\rightarrow$ stoichiometric ratio $\rightarrow$ moles $\rightarrow$ the quantity actually asked for. Every one of those chains begins with $n = m/M$. Practise until it is automatic.

STRUCTURE 1.4.4

Empirical and molecular formulae

SL + HL

Syllabus statement: the empirical formula of a compound gives the simplest ratio of atoms of each element present in that compound. The molecular formula gives the actual number of atoms of each element present in a molecule.

DEFINITIONS

Empirical formula — the formula showing the *simplest whole-number ratio* of the atoms of each element present.

Molecular formula — the formula showing the *actual number* of atoms of each element present in one molecule. It is always a whole-number multiple of the empirical formula.

COMPOUND	MOLECULAR FORMULA	EMPIRICAL FORMULA	COMMENT
Water	H ₂ O	H ₂ O	Already in simplest ratio
Ethane	C ₂ H ₆	CH ₃	Ratio 2:6 simplifies to 1:3
Ethene	C ₂ H ₄	CH ₂	Different compounds, same empirical formula
Propene	C ₃ H ₆	CH ₂	
Glucose	C ₆ H ₁₂ O ₆	CH ₂ O	Multiplying factor $n = 6$
Benzene	C ₆ H ₆	CH	Multiplying factor $n = 6$
Sodium chloride	—	NaCl	Giant structures have no molecular formula
Silicon dioxide	—	SiO ₂	

METHOD 1 — EMPIRICAL FORMULA FROM MASSES OR PERCENTAGES (THE FIVE-STEP ROUTINE)

- 1 **Write down the mass of each element.** If given percentages, simply treat them as masses in a 100 g sample — 40.0 % C becomes 40.0 g of C. This is legitimate because the ratio is unaffected by sample size.
- 2 **Divide each mass by that element's A_r** to get the amount in mol.
- 3 **Divide every answer by the smallest of them.** This forces the smallest value to exactly 1.
- 4 **Multiply all values by the smallest integer that makes them all whole numbers** (see the ratio table below).
- 5 **Write the empirical formula** using those whole numbers as subscripts.

STEP 4 — RECOGNISING THE COMMON AWKWARD RATIOS

IF A VALUE ENDS IN...	...IT IS REALLY	MULTIPLY EVERYTHING BY	WORKED ILLUSTRATION
.50	a half	2	1 : 1.50 → 2 : 3
.33 or .67	a third	3	1 : 1.33 → 3 : 4
.25 or .75	a quarter	4	1 : 2.25 → 4 : 9
.20, .40, .60, .80	a fifth	5	1 : 1.40 → 5 : 7

Anything within roughly ± 0.1 of a whole number (e.g. 2.98 or 3.04) should simply be rounded — that scatter is experimental error, not a fraction. A value such as 2.50 is *never* rounded to 3.

METHOD 2 — MOLECULAR FORMULA FROM EMPIRICAL FORMULA AND M_r

$$n = \frac{M_r \text{ of the compound}}{\text{mass of the empirical formula unit}} \quad \text{then} \quad \text{molecular formula} = (\text{empirical formula})_n$$

n must come out as a whole number — if it does not, re-check your empirical formula. Note that $n = 1$ is a perfectly acceptable answer (as for water or ammonia).

METHOD 3 — EMPIRICAL FORMULA FROM COMBUSTION ANALYSIS

A known mass of an organic compound containing C, H and possibly O is burned completely in excess oxygen. All the carbon ends up in CO_2 and all the hydrogen in H_2O , both of which are absorbed and weighed.

- 1 $n(\text{CO}_2) = m(\text{CO}_2) / 44.01 \rightarrow n(\text{C}) = n(\text{CO}_2)$ (one C per CO_2)
- 2 $n(\text{H}_2\text{O}) = m(\text{H}_2\text{O}) / 18.02 \rightarrow n(\text{H}) = 2 \times n(\text{H}_2\text{O})$ (two H per H_2O)
- 3 Convert back to masses: $m(\text{C}) = n(\text{C}) \times 12.01$ and $m(\text{H}) = n(\text{H}) \times 1.01$
- 4 **Oxygen is found by difference:** $m(\text{O}) = \text{mass of sample} - m(\text{C}) - m(\text{H})$, then $n(\text{O}) = m(\text{O}) / 16.00$
- 5 Divide the three amounts by the smallest and proceed as in Method 1.

EMPIRICAL FORMULA TRAPS

- ▶ **Never take oxygen from the CO_2 or H_2O .** That oxygen came from the air, not the compound. Oxygen in the sample is *only* ever found by mass difference.
- ▶ **Do not forget the factor of 2 for hydrogen.** Each H_2O molecule contains two H atoms.
- ▶ **If the percentages given add to less than 100 %, the remainder is usually the oxygen.** Check the total before you start.
- ▶ **Divide by A_r , not by the atomic number.** Surprisingly common under time pressure.
- ▶ **Do not round mid-calculation.** Rounding 1.33 to 1.3 can turn a clean $\times 3$ into a mess.
- ▶ **Answer the question asked.** "Empirical formula" and "molecular formula" are different marks; if the question gives you M_r , it almost certainly wants both.

HL STRETCH HARDER EMPIRICAL-FORMULA CONTEXTS

HL data-based questions dress the same arithmetic in less familiar clothing: determining x in a hydrate $\text{M} \cdot x\text{H}_2\text{O}$ from mass loss on heating; finding the formula of a metal oxide from the mass gain of a metal heated in air, or the mass loss when the oxide is reduced by hydrogen; deducing a formula from a mass spectrum's molecular ion peak (Structure 3.2) combined with combustion data; and using a formula together with the degree of unsaturation. In every case the underlying five steps are identical — only the route to the individual masses changes. See Solved Examples 13–15.

PART B — FORMULA & DATA SHEET

Everything on this page applies to **both SL and HL**. Only the items in the final column marked "given" appear in the IB data booklet — the rest must be memorised.

RELATIONSHIP	WHAT IT DOES	UNITS OF ANSWER	IN DATA BOOKLET?
$n = N / N_A$	Number of particles → amount in mol	mol	No
$N = n \times N_A$	Amount in mol → number of particles	none (a count)	No
$N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$	The Avogadro constant	mol^{-1}	Given
$n = m / M$	Mass → amount in mol	mol	No
$m = n \times M$	Amount → mass	g	No
$M = m / n$	Determining an unknown molar mass	g mol^{-1}	No
$N = (m / M) \times N_A$	Mass → number of particles in one step	none (a count)	No
mass of 1 particle = M / N_A	Molar mass → mass of a single atom/molecule	g	No
$A_r = \Sigma(\text{mass} \times \text{fractional abundance})$	Isotopic data → relative atomic mass	none	No
$M_r = \Sigma A_r$	Formula → relative formula mass	none	A_r values given
$\% X = (a \times A_r(X)) / M_r \times 100$	Percentage composition by mass	%	No
$n = M_r \div (\text{empirical formula mass})$	Multiplying factor for molecular formula	none	No

SYMBOLS, MEANINGS AND UNITS — LEARN THIS TABLE VERBATIM

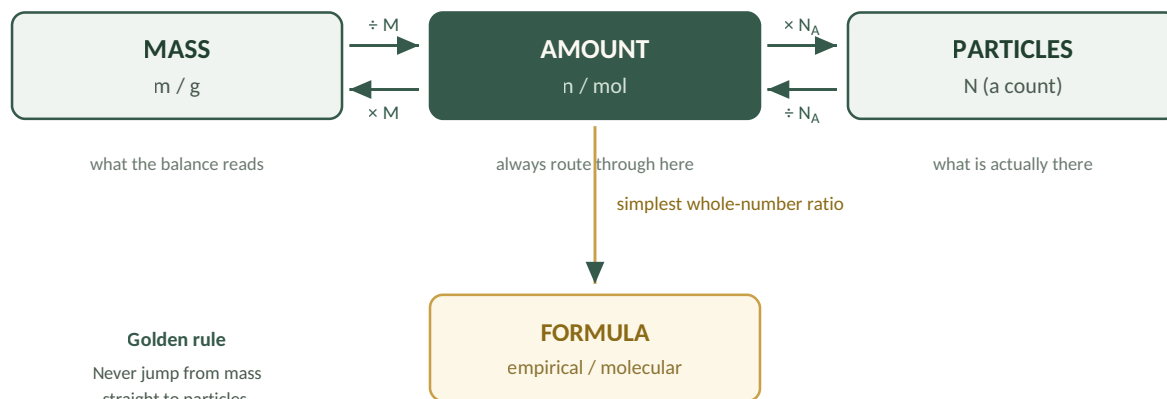
SYMBOL	QUANTITY	UNIT	WATCH OUT FOR
n	Amount of substance	mol	Never call this "number of moles of particles"; it is the amount.
N	Number of entities	none	Must state the entity: atoms, molecules, ions, formula units.
N_A or L	Avogadro constant	mol^{-1}	The unit mol^{-1} is part of the answer.
m	Mass	g	Convert kg, mg, tonnes to grams first.
M	Molar mass	g mol^{-1}	Has units — unlike M_r .
A_r	Relative atomic mass	none	Weighted mean, so rarely a whole number.
M_r	Relative formula mass	none	Include water of crystallisation.

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PART C — EXAM TECHNIQUE

The mole road map

Almost every question in this sub-topic is a journey between four boxes. Identify where the question starts and where it must finish, then read off the arrows.



Significant figures — how IB actually marks them

- ▶ Give the final answer to **the same number of significant figures as the least precise piece of data** in the question. In practice this is almost always **3 s.f.**
- ▶ **Carry full precision through the working** and round only at the very end. Premature rounding is the single most common source of "close but wrong" answers.
- ▶ An answer given to **too many** figures is usually accepted; an answer to **one** significant figure normally is not.
- ▶ Data-booklet A_r values are quoted to 2 d.p. — use them as given (35.45, not 35.5, and certainly not 35).
- ▶ Trailing zeros count: writing 0.0200 mol shows 3 s.f.; writing 0.02 mol shows 1.

Twelve habits that protect marks

#	HABIT	WHY IT MATTERS
1	Write the relationship before the numbers	Method marks are awarded for the correct relationship even when the arithmetic fails.
2	Label every line with the substance	" $n(\text{CO}_2) = 0.0150 \text{ mol}$ " is markable; a bare "0.0150" may not be.
3	Convert to grams on line one	Removes the most common factor-of-1000 error before it can propagate.
4	Underline what is actually asked	Atoms or molecules? Empirical or molecular? Mass or amount?
5	Show every step, however trivial	Error carried forward (ECF) means a slip in part (a) need not cost you parts (b) and (c).
6	Never round mid-calculation	Use the calculator's answer memory rather than re-typing rounded values.
7	Attach the unit to the final answer	mol, g, g mol^{-1} , % — and <i>no</i> unit for A_r and M_r .
8	Sanity-check the order of magnitude	Single atoms $\approx 10^{-23} \text{ g}$; laboratory samples $\approx 10^{-3}$ to 10 mol.
9	Estimate first, calculate second, on Paper 1A	At roughly 1.5 minutes per mark, a 5-second estimate often eliminates three options before you touch the calculator.
10	Count atoms out loud through brackets	$\text{Ca}(\text{NO}_3)_2 \rightarrow$ "one calcium, two nitrogen, six oxygen".
11	Read "state" and "deduce" carefully	"State" needs no working; "determine" and "deduce" need working shown.
12	Re-read the stem after answering	Confirms you answered for the right substance and the right entity.

KNOW YOUR STANDARD VALUES BY SIGHT — IT IS WORTH MINUTES ON PAPER 1A

Under the 2025 syllabus a calculator **and** the data booklet are permitted in *every* paper, including Paper 1A. That is a change from the old syllabus, and it shifts the challenge from mental arithmetic to **speed**: with about 1.5 minutes per mark, reaching for the periodic table for every A_r will cost you the paper.

Learn these by sight so you can estimate instantly and use the calculator only to confirm: $\text{H} = 1.01$, $\text{C} = 12.01$, $\text{N} = 14.01$, $\text{O} = 16.00$, $\text{Na} = 22.99$, $\text{S} = 32.07$, $\text{Cl} = 35.45$, $\text{Ca} = 40.08$, $\text{Fe} = 55.85$, $\text{Cu} = 63.55$. And these M_r values: $\text{H}_2\text{O} = 18.02$, $\text{CO}_2 = 44.01$, $\text{O}_2 = 32.00$, $\text{N}_2 = 28.02$, $\text{NH}_3 = 17.04$, $\text{CH}_4 = 16.05$, $\text{NaCl} = 58.44$, $\text{CaCO}_3 = 100.09$, $\text{H}_2\text{SO}_4 = 98.09$, $\text{C}_6\text{H}_{12}\text{O}_6 = 180.18$. Also recognise $6.02 \times 10^{23} \div 2 = 3.01 \times 10^{23}$ and $\div 4 = 1.505 \times 10^{23}$ on sight — they appear constantly as distractors.

PART D — FULLY SOLVED EXAMPLES

Cover the solution, attempt the question, then compare. Examples 1–14 are **SL + HL** core; Examples 5 and 15 are marked **HL Stretch** because of their extra step, but the underlying content is common to both levels.

Sub-topic 1.4.1 — Counting entities

SL + HL

Example 1. A sample of iron contains 3.01×10^{23} atoms. Calculate (a) the amount of iron, in mol, and (b) the mass of the sample.

- 1 Choose the relationship: $n = N / N_A$
- 2 $n(\text{Fe}) = (3.01 \times 10^{23}) \div (6.02 \times 10^{23} \text{ mol}^{-1}) = 0.500 \text{ mol}$
- 3 $M(\text{Fe}) = 55.85 \text{ g mol}^{-1}$, so $m = n M = 0.500 \times 55.85 = 27.925 \text{ g}$

(a) **0.500 mol** (b) **27.9 g (3 s.f.)**

Example 2. Calculate the total number of *atoms* in 0.250 mol of aluminium sulfate, $\text{Al}_2(\text{SO}_4)_3$.

- 1 Count atoms per formula unit — read carefully through the bracket: $2 \text{ Al} + 3 \text{ S} + 12 \text{ O} = \mathbf{17 \text{ atoms}}$.
- 2 Amount of atoms = $0.250 \text{ mol} \times 17 = 4.25 \text{ mol of atoms}$.
- 3 $N = 4.25 \times 6.02 \times 10^{23} = 2.5585 \times 10^{24}$

2.56×10^{24} atoms

If the question had asked for *ions*, the count would be $2 \text{ Al}^{3+} + 3 \text{ SO}_4^{2-} = 5$ ions per formula unit, giving 1.25 mol of ions = 7.53×10^{23} ions. Always re-read which entity is wanted.

Example 3. A sample contains 0.200 mol of iron(III) sulfate, $\text{Fe}_2(\text{SO}_4)_3$. Determine (a) the number of ions present and (b) the number of oxygen atoms present.

- 1 Ions per formula unit: 2 Fe^{3+} and 3 SO_4^{2-} — the sulfate ion is a single polyatomic ion, so that is **5 ions**, not $5 + 12$.
- 2 $n(\text{ions}) = 0.200 \times 5 = 1.00 \text{ mol} \rightarrow N = 1.00 \times 6.02 \times 10^{23} = 6.02 \times 10^{23} \text{ ions}$
- 3 Oxygen atoms per formula unit = $3 \times 4 = 12$, so $n(\text{O}) = 0.200 \times 12 = 2.40 \text{ mol}$
- 4 $N(\text{O}) = 2.40 \times 6.02 \times 10^{23} = 1.4448 \times 10^{24}$

(a) **6.02×10^{23} ions** (b) **1.44×10^{24} oxygen atoms**

Example 4. Naturally occurring chlorine consists of 75.77 % ^{35}Cl (isotopic mass 34.97) and 24.23 % ^{37}Cl (isotopic mass 36.97). Calculate the relative atomic mass of chlorine.

- 1 $A_r = \Sigma(\text{isotopic mass} \times \text{fractional abundance})$
- 2 $= (34.97 \times 0.7577) + (36.97 \times 0.2423)$
- 3 $= 26.4967 + 8.9578 = 35.4545$

$A_r(\text{Cl}) = 35.45$ — **no units**. Note that the answer lies much closer to 35 than to 37, as it must, because the lighter isotope is three times more abundant. Use that as your check.

Example 5. HL STRETCH Gallium has two naturally occurring isotopes, ^{69}Ga (isotopic mass 68.93) and ^{71}Ga (isotopic mass 70.92). The relative atomic mass of gallium is 69.72. Calculate the percentage abundance of each isotope.

- 1 Let the fractional abundance of ^{69}Ga be x . Then the fractional abundance of ^{71}Ga is $(1 - x)$.
- 2 $68.93x + 70.92(1 - x) = 69.72$
- 3 $68.93x + 70.92 - 70.92x = 69.72 \rightarrow -1.99x = -1.20$
- 4 $x = 1.20 \div 1.99 = 0.6030$

60.3 % ^{69}Ga and 39.7 % ^{71}Ga

Check: the two abundances must sum to 100 %, and the A_r of 69.72 sits nearer 68.93 than 70.92, so the lighter isotope must be the more abundant. Both conditions hold.

Example 6. Determine the relative formula mass of (a) calcium nitrate, $\text{Ca}(\text{NO}_3)_2$, and (b) hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

- 1 (a) Count through the bracket: 1 Ca, 2 N, 6 O.
- 2 $M_r = 40.08 + 2(14.01) + 6(16.00) = 40.08 + 28.02 + 96.00 = 164.10$
- 3 (b) Treat the dot as a plus sign. Anhydrous part: $63.55 + 32.07 + 4(16.00) = 159.62$
- 4 Water part: $5 \times M_r(\text{H}_2\text{O}) = 5 \times 18.02 = 90.10$
- 5 $M_r = 159.62 + 90.10 = 249.72$

(a) 164.10 (b) 249.72 — neither answer carries units.

Sub-topic 1.4.3 — Molar mass calculations

SL + HL

Example 7. Calculate the mass of 0.150 mol of anhydrous sodium carbonate, Na_2CO_3 .

- 1 $M = 2(22.99) + 12.01 + 3(16.00) = 45.98 + 12.01 + 48.00 = 105.99 \text{ g mol}^{-1}$
- 2 $m = n M = 0.150 \times 105.99 = 15.8985 \text{ g}$

15.9 g — the unit is essential.**Example 8.** A beaker contains 4.50 g of water. Calculate (a) the amount of water in mol, (b) the number of water molecules, and (c) the total number of atoms present.

- 1 $M(\text{H}_2\text{O}) = 2(1.01) + 16.00 = 18.02 \text{ g mol}^{-1}$
- 2 $n = m / M = 4.50 \div 18.02 = 0.24972 \text{ mol}$
- 3 $N(\text{molecules}) = 0.24972 \times 6.02 \times 10^{23} = 1.5033 \times 10^{23}$
- 4 Each molecule contains 3 atoms, so $N(\text{atoms}) = 3 \times 1.5033 \times 10^{23} = 4.510 \times 10^{23}$

(a) 0.250 mol (b) 1.50×10^{23} molecules (c) 4.51×10^{23} atoms**Example 9.** Calculate the mass, in grams, of one molecule of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$.

- 1 $M = 6(12.01) + 12(1.01) + 6(16.00) = 72.06 + 12.12 + 96.00 = 180.18 \text{ g mol}^{-1}$
- 2 mass of one molecule = $M / N_A = 180.18 \div (6.02 \times 10^{23})$
- 3 $= 2.9930 \times 10^{-22} \text{ g}$

 $2.99 \times 10^{-22} \text{ g}$ The negative exponent is the check. A single molecule must be vanishingly light; an answer of 10^{+25} means you multiplied instead of dividing.**Example 10.** Determine the percentage composition by mass of ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$.

- 1 Atom count through the bracket: 2 N, 8 H, 1 S, 4 O.
- 2 $M_r = 2(14.01) + 8(1.01) + 32.07 + 4(16.00) = 28.02 + 8.08 + 32.07 + 64.00 = 132.17$
- 3 % N = $(28.02 / 132.17) \times 100 = 21.20 \%$ · % H = $(8.08 / 132.17) \times 100 = 6.11 \%$
- 4 % S = $(32.07 / 132.17) \times 100 = 24.26 \%$ · % O = $(64.00 / 132.17) \times 100 = 48.42 \%$

21.2 % N, 6.11 % H, 24.3 % S, 48.4 % OAlways add your percentages: $21.20 + 6.11 + 24.26 + 48.42 = 99.99 \%$. Anything other than 100 % (± 0.1 from rounding) means an arithmetic slip.

Sub-topic 1.4.4 — Empirical and molecular formulae

SL + HL

Example 11. A compound contains 26.6 % potassium, 35.4 % chromium and 38.1 % oxygen by mass. Determine its empirical formula.

	K	CR	O
Mass in 100 g / g	26.6	35.4	38.1
$\div A_r$	$26.6 \div 39.10$	$35.4 \div 52.00$	$38.1 \div 16.00$
Amount / mol	0.6803	0.6808	2.381
$\div$ smallest (0.6803)	1.000	1.001	3.500
$\times 2$ (to clear the half)	2	2	7

Empirical formula = $K_2Cr_2O_7$ (potassium dichromate)

The critical judgement is at 3.500 — that is a genuine half, not experimental scatter, so it must be cleared by multiplying by 2. Rounding it to 4 would give the wrong compound.

Example 12. A hydrocarbon has the empirical formula C_2H_5 and a relative molecular mass of 58.14. Determine its molecular formula.

- 1 Mass of the empirical formula unit: $2(12.01) + 5(1.01) = 24.02 + 5.05 = 29.07$
- 2 $n = M_r \div \text{empirical formula mass} = 58.14 \div 29.07 = 2.000$
- 3 Molecular formula = $(C_2H_5)_2$

C_4H_{10} (butane)

Write out the final formula properly — $(C_2H_5)_2$ alone will not usually be credited as the molecular formula.

Example 13. Complete combustion of 0.0850 g of a hydrocarbon produced 0.257 g of carbon dioxide and 0.132 g of water. The relative molecular mass of the hydrocarbon is 58.1. Determine (a) its empirical formula and (b) its molecular formula.

- 1 $n(CO_2) = 0.257 \div 44.01 = 5.840 \times 10^{-3} \text{ mol} \rightarrow n(C) = 5.840 \times 10^{-3} \text{ mol}$
- 2 $n(H_2O) = 0.132 \div 18.02 = 7.325 \times 10^{-3} \text{ mol} \rightarrow n(H) = 2 \times 7.325 \times 10^{-3} = 1.465 \times 10^{-2} \text{ mol}$
- 3 Ratio C : H = $5.840 \times 10^{-3} : 1.465 \times 10^{-2}$. Divide by the smaller: 1 : 2.508
- 4 2.508 is a half, so multiply both by 2: C : H = 2 : 5 $\rightarrow$ empirical formula C_2H_5
- 5 Empirical formula mass = 29.07; $n = 58.1 \div 29.07 = 2.00$

(a) C_2H_5 (b) C_4H_{10}

Verification worth doing: mass of C = $5.840 \times 10^{-3} \times 12.01 = 0.0701 \text{ g}$; mass of H = $1.465 \times 10^{-2} \times 1.01 = 0.0148 \text{ g}$. Total = 0.0849 g $\approx$ 0.0850 g, the mass of the sample. Because the two masses account for the whole sample, the compound contains *no* oxygen — consistent with "hydrocarbon".

Example 14. 5.00 g of hydrated barium chloride, $\text{BaCl}_2 \cdot x\text{H}_2\text{O}$, is heated to constant mass. The residue of anhydrous BaCl_2 has a mass of 4.26 g. Determine the value of x .

- 1 Mass of water driven off = $5.00 - 4.26 = 0.74 \text{ g}$
- 2 $M(\text{BaCl}_2) = 137.33 + 2(35.45) = 208.23 \text{ g mol}^{-1}$
- 3 $n(\text{BaCl}_2) = 4.26 \div 208.23 = 2.046 \times 10^{-2} \text{ mol}$
- 4 $n(\text{H}_2\text{O}) = 0.74 \div 18.02 = 4.107 \times 10^{-2} \text{ mol}$
- 5 Ratio $\text{H}_2\text{O} : \text{BaCl}_2 = 4.107 \times 10^{-2} \div 2.046 \times 10^{-2} = 2.007$

$x = 2$, so the formula is $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$

"Heated to constant mass" is the examiner's signal that *all* the water has been driven off. x must be a whole number, so 2.007 rounds to 2 — that 0.007 is experimental scatter, not chemistry.

Example 15. **HL STRETCH** (a) 1.12 g of iron is heated in excess oxygen until it is completely converted to an oxide. The oxide produced has a mass of 1.60 g. Determine the empirical formula of the oxide. (b) A quarry produces limestone that is 92.0 % calcium carbonate by mass. Calculate the amount of CaCO_3 , in mol, in 1.00 tonne of this limestone.

- 1 (a) The extra mass must all be oxygen: $m(\text{O}) = 1.60 - 1.12 = 0.48 \text{ g}$
- 2 $n(\text{Fe}) = 1.12 \div 55.85 = 2.005 \times 10^{-2} \text{ mol}$ · $n(\text{O}) = 0.48 \div 16.00 = 3.00 \times 10^{-2} \text{ mol}$
- 3 Divide by the smaller: $\text{Fe} : \text{O} = 1.000 : 1.496 \rightarrow \text{a half, so } \times 2 \rightarrow 2 : 3$
- 4 Empirical formula = Fe_2O_3 (iron(III) oxide — check the oxidation state is sensible for iron)
- 5 (b) Convert to grams *first*: $1.00 \text{ tonne} = 1.00 \times 10^6 \text{ g}$
- 6 Mass of $\text{CaCO}_3 = 0.920 \times 1.00 \times 10^6 = 9.20 \times 10^5 \text{ g}$
- 7 $M(\text{CaCO}_3) = 40.08 + 12.01 + 3(16.00) = 100.09 \text{ g mol}^{-1}$
- 8 $n = 9.20 \times 10^5 \div 100.09 = 9192 \text{ mol}$

(a) Fe_2O_3 (b) $9.19 \times 10^3 \text{ mol}$

Part (b) rehearses two habits that HL Paper 2 punishes hard: converting tonnes to grams before anything else, and applying the purity factor to the *mass* before dividing by M , never afterwards.

SECTION E1

Paper 1A style — multiple choice

SL + HL

20 questions. Target time: 30 minutes. Choose the one best answer. A calculator and the data booklet are permitted, but train yourself to eliminate options by estimation first. Mark scheme with full explanations begins on the page after Section E3.

1. Which sample contains the greatest number of atoms?

- A. 1.0 mol of He
- B. 1.0 mol of H_2O
- C. 1.0 mol of CH_4
- D. 1.0 mol of CO_2

2. What amount of oxygen atoms, in mol, is present in 0.50 mol of $\text{Al}_2(\text{SO}_4)_3$?

- A. 2.0
- B. 4.0
- C. 6.0
- D. 12

3. Which sample has the greatest mass?

- A. 1.00 mol of H_2O
- B. 0.500 mol of CO_2
- C. 0.200 mol of H_2SO_4
- D. 0.100 mol of $\text{C}_6\text{H}_{12}\text{O}_6$

4. Which statement about relative atomic mass is correct?

- A. It has the units g mol^{-1} .
- B. It is the mass, in grams, of one atom of the element.
- C. It is the weighted mean mass of the atoms in a naturally occurring sample of the element, relative to $1/12$ of the mass of an atom of carbon-12.
- D. It is always a whole number because atoms contain whole numbers of protons and neutrons.

5. An element has two isotopes, of isotopic mass 10 (abundance 20.0 %) and isotopic mass 11 (abundance 80.0 %). What is its relative atomic mass?

- A. 10.2
- B. 10.5
- C. 10.8
- D. 11.0

6. What is the relative formula mass of ammonium carbonate, $(\text{NH}_4)_2\text{CO}_3$?

- A. 78.06
- B. 92.09
- C. 96.11
- D. 112.12

7. What is the mass of 0.250 mol of calcium carbonate, CaCO_3 ?

- A. 12.5 g
- B. 25.0 g
- C. 40.0 g
- D. 400 g

8. How many molecules are present in 8.0 g of oxygen gas, O_2 ?

- A. 1.5×10^{23}
- B. 3.0×10^{23}
- C. 6.0×10^{23}
- D. 1.2×10^{24}

9. A compound contains 40.0 % carbon, 6.7 % hydrogen and 53.3 % oxygen by mass. What is its empirical formula?

- A. CHO
- B. CH_2O
- C. $\text{C}_2\text{H}_4\text{O}$
- D. $\text{C}_2\text{H}_6\text{O}$

10. A compound has the empirical formula CH_2 and a relative molecular mass of 56. What is its molecular formula?

- A. C_2H_4
- B. C_3H_6
- C. C_4H_8
- D. C_5H_{10}

11. Which pair of compounds has the same empirical formula?

- A. C_2H_4 and C_3H_6
- B. C_2H_6 and C_3H_6
- C. CH_4 and C_2H_4
- D. C_6H_6 and C_6H_{12}

12. What is the percentage by mass of nitrogen in ammonium nitrate, NH_4NO_3 ?

- A. 17.5 %
- B. 35.0 %
- C. 43.7 %
- D. 70.0 %

13. Which sample contains the greatest amount, in mol, of chloride ions?

- A. 0.10 mol of NaCl
- B. 0.10 mol of MgCl_2
- C. 0.10 mol of AlCl_3
- D. 0.10 mol of CCl_4

14. Which sample contains 6.02×10^{23} atoms?

- A. 1.0 mol of H_2
- B. 0.50 mol of O_2
- C. 1.0 mol of H_2O
- D. 2.0 mol of He

15. The relative atomic mass of magnesium is 24.31. What is the mass, in grams, of one atom of magnesium?

- A. 4.04×10^{-23}
- B. 2.43×10^{-23}
- C. 1.46×10^{25}
- D. 4.04×10^{-26}

16. What is the percentage by mass of water in $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$? ($M_r = 246.52$)

- A. 7.31 %
- B. 48.8 %
- C. 51.2 %
- D. 68.6 %

17. 0.100 mol of a hydrocarbon has a mass of 5.60 g. What is the hydrocarbon?

- A. C_3H_4
- B. C_4H_8
- C. C_4H_{10}
- D. C_5H_{12}

18. Which statement is correct?

- A. One mole of any substance has a mass of 12 g.
- B. One mole of any gas contains 6.02×10^{23} atoms.
- C. The Avogadro constant has the units mol^{-1} .
- D. Relative formula mass has the units g mol^{-1} .

19. Complete combustion of 0.100 mol of a hydrocarbon produces 0.300 mol of CO_2 and 0.400 mol of H_2O . What is the molecular formula of the hydrocarbon?

- A. C_3H_4
- B. C_3H_6
- C. C_3H_8
- D. C_4H_8

20. Which 1.0 g sample contains the greatest number of atoms?

- A. H_2
- B. He
- C. Li
- D. C

iBLaurel 

SECTION E2

Paper 1B style — data-based questions

SL + HL

2 questions, 20 marks. Target time: 30 minutes. Calculator and data booklet permitted. Show all working — method marks are available even where the final answer is wrong.

Question 1. A student investigates the isotopic composition of magnesium using a mass spectrometer.^[10 marks] The instrument produces the following data.

ISOTOPE	ISOTOPIC MASS	RELATIVE ABUNDANCE / %
^{24}Mg	23.985	78.99
^{25}Mg	24.986	10.00
^{26}Mg	25.982	11.01

(a) State what is meant by the term *relative atomic mass*. [2]

(b) Calculate the relative atomic mass of magnesium from the data above. Give your answer to four significant figures. [2]

(c) The student weighs out a strip of magnesium ribbon of mass 0.486 g. Determine the amount of magnesium, in mol, in the ribbon. [2]

(d) Calculate the number of magnesium atoms in the ribbon. [1]

(e) The ribbon is completely converted to magnesium oxide, MgO. Calculate the mass of magnesium oxide formed. [2]

(f) Explain why the relative atomic mass of magnesium is not a whole number, even though every individual magnesium atom has a mass number that is a whole number. [1]

Question 2. A compound **X** is known to contain only carbon, hydrogen and oxygen. A 0.450 g sample of **X** is burned completely in an excess of oxygen. The carbon dioxide and water produced are absorbed separately and weighed. [10 marks]

MEASUREMENT	VALUE / G
Mass of sample of X	0.450
Mass of CO ₂ collected	0.660
Mass of H ₂ O collected	0.270

(a) Calculate the amount, in mol, of carbon dioxide produced, and hence the mass of carbon in the sample. [2]

(b) Calculate the amount, in mol, of water produced, and hence the mass of hydrogen in the sample. [2]

(c) Deduce the mass of oxygen present in the 0.450 g sample of **X**. Explain why the oxygen cannot be determined directly from the mass of CO₂ collected. [2]

(d) Determine the empirical formula of **X**. [2]

(e) The relative molecular mass of **X** is 90.08. Deduce the molecular formula of **X**. [2]

SECTION E3

Paper 2 style — structured questions

SL + HL

3 questions, 30 marks. Target time: 45 minutes. Calculator and data booklet permitted.

Question 1. Hydrated copper(II) sulfate has the formula $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$. A student determines x by heating a sample in a crucible until all the water of crystallisation has been driven off. The following masses are recorded.

MEASUREMENT	MASS / G
Empty crucible	24.30
Crucible + hydrated copper(II) sulfate	27.80
Crucible + residue after heating to constant mass	26.54

- (a) Calculate the mass of hydrated copper(II) sulfate used and the mass of anhydrous copper(II) sulfate produced. [2]
- (b) Determine the amount, in mol, of anhydrous CuSO_4 and the amount, in mol, of water driven off. [3]
- (c) Deduce the value of x . [2]
- (d) State what is meant by *heating to constant mass* and explain why this procedure is necessary. [2]
- (e) Suggest one reason why a student might obtain a value of x that is lower than the true value. [1]

Question 2. Nitrogen fertilisers are compared by their percentage of nitrogen by mass. Three common fertilisers are ammonium nitrate, NH_4NO_3 ; urea, $\text{CO}(\text{NH}_2)_2$; and ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$. [10 marks]

- (a) Calculate the relative formula mass of each of the three fertilisers. [3]
- (b) Determine the percentage by mass of nitrogen in each fertiliser, and identify which delivers the most nitrogen per kilogram. [3]
- (c) A farmer needs to supply 1.00 kg of nitrogen to a field. Calculate the mass of urea, in kg, required. [2]
- (d) Calculate the number of nitrogen atoms in a 10.0 g sample of urea. [2]

Question 3. Vitamin C (ascorbic acid) contains only carbon, hydrogen and oxygen. Analysis of a pure sample gives the composition 40.91 % carbon and 4.58 % hydrogen by mass. The relative molecular mass of vitamin C is 176.14. [10 marks]

- (a) Deduce the percentage by mass of oxygen in vitamin C. [1]
- (b) Determine the empirical formula of vitamin C, showing all your working. [4]
- (c) Deduce the molecular formula of vitamin C. [2]
- (d) A vitamin C tablet contains 500 mg of vitamin C. Calculate the amount, in mol, of vitamin C in one tablet, and the number of vitamin C molecules it contains. [3]

HL STRETCH EXTENSION PART (E)

- (e) A second organic compound has the same empirical formula as vitamin C but a relative molecular mass of 88.07. State its molecular formula, and explain why two compounds with the same empirical formula may have very different chemical properties. [2]

PART F — MARK SCHEMES

MARK SCHEME E1

Paper 1A style — answer key

SL + HL

Q	1	2	3	4	5	6	7	8	9	10
Ans	C	C	B	C	C	C	B	A	B	C
Q	11	12	13	14	15	16	17	18	19	20
Ans	A	B	C	B	A	C	B	C	C	A

EXPLANATIONS — INCLUDING WHY THE DISTRACTORS ARE WRONG

1 — C. Count atoms per formula unit and multiply by 1.0 mol: He = 1, H₂O = 3, CH₄ = 5, CO₂ = 3. Equal amounts of substance contain equal numbers of *molecules*, never equal numbers of *atoms*.

2 — C. The subscript 4 inside the bracket and the subscript 3 outside multiply: $4 \times 3 = 12$ oxygen atoms per formula unit. $n(\text{O}) = 0.50 \times 12 = 6.0$ mol. Option D is the atom count per formula unit, not the amount; option A forgets the bracket entirely.

3 — B. A 18.02 g; B 22.0 g; C 19.6 g; D 18.0 g. Note how nearly identical A and D are — a reminder that a large M_r does not guarantee a large mass when the amount is small.

4 — C. This is the definition verbatim. A is wrong because A_r is a ratio and therefore has *no* units (that description fits molar mass). B describes the absolute mass of a single atom. D is wrong because A_r is a *weighted mean* across isotopes, so it is rarely a whole number.

5 — C. $A_r = (10 \times 0.200) + (11 \times 0.800) = 2.00 + 8.80 = 10.8$. Option B (10.5) is the trap: it is the simple, unweighted average of 10 and 11, which ignores the abundances.

6 — C. (NH₄)₂CO₃ contains 2 N, 8 H, 1 C and 3 O: $28.02 + 8.08 + 12.01 + 48.00 = 96.11$. Option A (78.06) is exactly what you get by ignoring the subscript 2 outside the bracket — by far the most common error in this question type.

7 — B. $M_r(\text{CaCO}_3) = 100.09$, so $m = 0.250 \times 100.09 = 25.0$ g. Option C uses $A_r(\text{Ca})$ alone; option D divides instead of multiplying.

8 — A. $n(\text{O}_2) = 8.0 \div 32.00 = 0.25$ mol, so $N = 0.25 \times 6.02 \times 10^{23} = 1.5 \times 10^{23}$ molecules. Option B (3.0×10^{23}) is the number of oxygen *atoms* — read the question. Option C would follow from using $A_r = 16.00$ instead of $M_r(\text{O}_2) = 32.00$.

9 — B. C: $40.0 \div 12.01 = 3.33$; H: $6.7 \div 1.01 = 6.63$; O: $53.3 \div 16.00 = 3.33$. Dividing by 3.33 gives 1 : 1.99 : 1, so CH₂O. Option D is ethanol's molecular formula, which has a completely different composition.

10 — C. Empirical formula mass of CH₂ = 14.03; $n = 56 \div 14.03 = 4.00$; molecular formula = (CH₂)₄ = C₄H₈.

11 — A. C₂H₄ simplifies to CH₂ and C₃H₆ also simplifies to CH₂. In D, C₆H₆ gives CH but C₆H₁₂ gives CH₂. In B, C₂H₆ gives CH₃ and C₃H₆ gives CH₂.

12 — B. NH₄NO₃ contains **two** nitrogen atoms. $M_r = 80.06$, so % N = $(2 \times 14.01) \div 80.06 \times 100 = 35.0$ %. Option A (17.5 %) counts only one nitrogen — a classic slip when the same element appears twice in a formula.

13 — C. AlCl_3 gives $3 \times 0.10 = 0.30 \text{ mol Cl}^-$. Option D is the discriminator: CCl_4 is a *covalent molecular* compound, so it contains chlorine *atoms* covalently bonded to carbon — no chloride ions at all.

14 — B. 0.50 mol of O_2 contains $0.50 \times 2 = 1.0 \text{ mol of O atoms} = 6.02 \times 10^{23} \text{ atoms}$. A gives 1.2×10^{24} , C gives 1.8×10^{24} , D gives 1.2×10^{24} .

15 — A. mass of one atom = $M \div N_A = 24.31 \div (6.02 \times 10^{23}) = 4.04 \times 10^{-23} \text{ g}$. Option D is the same answer expressed in *kilograms* — correct physics, wrong unit for the question. Option C comes from multiplying instead of dividing.

16 — C. $\% \text{H}_2\text{O} = (7 \times 18.02) \div 246.52 \times 100 = 126.14 \div 246.52 \times 100 = 51.2 \%$. Option A forgets the factor of 7; option B is the percentage of anhydrous MgSO_4 , which is the complement ($100 - 51.2 = 48.8$).

17 — B. $M = m \div n = 5.60 \div 0.100 = 56.0 \text{ g mol}^{-1}$. $\text{C}_4\text{H}_8 = 4(12.01) + 8(1.01) = 56.12$. $\text{C}_4\text{H}_{10} = 58.14$ and $\text{C}_3\text{H}_4 = 40.07$, neither of which matches.

18 — C. The Avogadro constant is a number *per mole*, so its unit is mol^{-1} . A confuses the mole with carbon-12's molar mass; B fails for any molecular gas (1 mol of O_2 contains 1.2×10^{24} atoms); D is wrong because M_r is a dimensionless ratio — it is *molar mass* that has the unit g mol^{-1} .

19 — C. Every CO_2 contains one C, so $n(\text{C})$ per mole of hydrocarbon = $0.300 \div 0.100 = 3$. Every H_2O contains two H, so $n(\text{H})$ per mole = $(2 \times 0.400) \div 0.100 = 8$. The molecular formula is C_3H_8 . Option B is what you get by forgetting to double the water.

20 — A. Convert each 1.0 g to an amount of *atoms*: $\text{H}_2 \rightarrow (1.0 \div 2.02) \times 2 = 0.99 \text{ mol}$; $\text{He} \rightarrow 0.25 \text{ mol}$; $\text{Li} \rightarrow 0.14 \text{ mol}$; $\text{C} \rightarrow 0.083 \text{ mol}$. The lightest element gives the most atoms per gram, and the factor of 2 for the diatomic molecule makes hydrogen the clear winner.

MARK SCHEME E2

Paper 1B style

SL + HL

"ECF" = error carried forward: a correct method applied to an earlier incorrect answer still earns the later marks. Accept answers within reasonable rounding. "OWTTE" = or words to that effect.

Question 1 [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	The weighted mean mass of the atoms in a naturally occurring sample of the element [1] relative to $^{1/12}$ of the mass of an atom of carbon-12 [1] <i>Award the second mark only if the carbon-12 reference is explicit. Do not award if units are given.</i>	2
(b)	$A_r = (23.985 \times 0.7899) + (24.986 \times 0.1000) + (25.982 \times 0.1101)$ [1] $= 18.9457 + 2.4986 + 2.8606 = 24.30$ [1] <i>Accept 24.305. Do not accept a unit. Award only 1 mark for a simple mean of 24.98.</i>	2
(c)	$n = m / M = 0.486 / 24.31$ [1] = 0.0200 mol [1] <i>Accept 2.00×10^{-2} mol. Accept use of 24.30 from (b) (ECF). Unit required for the second mark.</i>	2
(d)	$N = 0.0200 \times 6.02 \times 10^{23} = 1.20 \times 10^{22}$ atoms [1] (ECF from (c))	1
(e)	Mg : MgO is 1 : 1, so $n(\text{MgO}) = 0.0200$ mol; $M(\text{MgO}) = 24.31 + 16.00 = 40.31 \text{ g mol}^{-1}$ [1] $m = 0.0200 \times 40.31 = 0.806 \text{ g}$ [1] <i>Accept 0.81 g. ECF from (c).</i>	2
(f)	Magnesium exists as a mixture of isotopes with different masses, and A_r is the weighted average of these, so it need not be a whole number [1] OWTTE <i>Do not accept "because of the electrons" or "because of rounding".</i>	1

Question 2 [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	$n(\text{CO}_2) = 0.660 / 44.01 = 0.0150 \text{ mol}$ [1] $n(\text{C}) = 0.0150 \text{ mol}$, so $m(\text{C}) = 0.0150 \times 12.01 = 0.180 \text{ g}$ [1]	2
(b)	$n(\text{H}_2\text{O}) = 0.270 / 18.02 = 0.0150 \text{ mol}$ [1] $n(\text{H}) = 2 \times 0.0150 = 0.0300 \text{ mol}$, so $m(\text{H}) = 0.0300 \times 1.01 = 0.0303 \text{ g}$ [1] <i>Do not award the second mark if the factor of 2 is omitted.</i>	2
(c)	$m(\text{O}) = 0.450 - 0.180 - 0.0303 = 0.240 \text{ g}$ [1] (ECF) The oxygen in the CO_2 and H_2O comes from the excess oxygen used for combustion as well as from X, so it cannot be attributed to the sample; oxygen must therefore be found by difference [1] OWTTE	2
(d)	$n(\text{O}) = 0.240 / 16.00 = 0.0150 \text{ mol}$; ratio C : H : O = 0.0150 : 0.0300 : 0.0150 [1] Divide by smallest $\rightarrow 1 : 2 : 1 \rightarrow \text{CH}_2\text{O}$ [1]	2
(e)	Empirical formula mass = $12.01 + 2(1.01) + 16.00 = 30.03$; $n = 90.08 / 30.03 = 3.00$ [1] Molecular formula = $\text{C}_3\text{H}_6\text{O}_3$ [1] <i>Do not accept $(\text{CH}_2\text{O})_3$ as the final answer.</i>	2

MARK SCHEME E3

Paper 2 style

SL + HL

Question 1 — Hydrated copper(II) sulfate [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	Hydrated salt = $27.80 - 24.30 = 3.50$ g [1] Anhydrous $\text{CuSO}_4 = 26.54 - 24.30 = 2.24$ g [1]	2
(b)	$M(\text{CuSO}_4) = 63.55 + 32.07 + 64.00 = 159.62$ g mol ⁻¹ [1] $n(\text{CuSO}_4) = 2.24 / 159.62 = 0.0140$ mol [1] Mass of water = $3.50 - 2.24 = 1.26$ g, so $n(\text{H}_2\text{O}) = 1.26 / 18.02 = 0.0699$ mol [1]	3
(c)	Ratio $\text{H}_2\text{O} : \text{CuSO}_4 = 0.0699 / 0.0140 = 4.98$ [1] Therefore $x = 5$, formula $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ [1] ECF from (b). Award the second mark only if a whole number is stated.	2
(d)	The sample is heated, cooled and reweighed repeatedly until two successive weighings agree [1] This confirms that <i>all</i> the water of crystallisation has been driven off, so the mass loss corresponds to the full amount of water [1] OWTTE	2
(e)	Any one of: incomplete dehydration / insufficient heating time, so not all the water is driven off; the cooled residue reabsorbs moisture from the air before weighing; spitting of the sample is <i>not</i> accepted here because it would lower both masses. [1] The answer must explain a lower apparent amount of water. Accept "some hydrate remained undecomposed" OWTTE.	1

Question 2 — Nitrogen fertilisers [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	NH_4NO_3 : $2(14.01) + 4(1.01) + 3(16.00) = 80.06$ [1] $\text{CO}(\text{NH}_2)_2$: $12.01 + 16.00 + 2(14.01) + 4(1.01) = 60.07$ [1] $(\text{NH}_4)_2\text{SO}_4$: $2(14.01) + 8(1.01) + 32.07 + 4(16.00) = 132.17$ [1] No units. Note that all three contain exactly 2 N, so the numerator is 28.02 in every case.	3
(b)	NH_4NO_3 : $(28.02 / 80.06) \times 100 = 35.0$ % [1] Urea: $(28.02 / 60.07) \times 100 = 46.6$ % · $(\text{NH}_4)_2\text{SO}_4$: $(28.02 / 132.17) \times 100 = 21.2$ % [1] Urea delivers the most nitrogen per kilogram [1]	3
(c)	Mass of urea = 1.00 kg $\div 0.4665$ [1] = 2.14 kg [1] Alternative route, equally credited: $n(\text{N}) = 1000 / 14.01 = 71.4$ mol $\rightarrow n(\text{urea}) = 35.7$ mol $\rightarrow m = 35.7 \times 60.07 = 2144$ g = 2.14 kg .	2
(d)	$n(\text{urea}) = 10.0 / 60.07 = 0.166$ mol, so $n(\text{N}) = 2 \times 0.166 = 0.333$ mol [1] $\text{N} = 0.333 \times 6.02 \times 10^{23} = 2.00 \times 10^{23}$ nitrogen atoms [1] Award only 1 mark for 1.00×10^{23} (factor of 2 omitted).	2

Question 3 — Vitamin C [10 marks + 2 HL Stretch]

PART	EXPECTED ANSWER	MARKS
(a)	% O = $100 - 40.91 - 4.58 = 54.51\%$ [1]	1
(b)	Divide each percentage by A_r : C = $40.91 / 12.01 = 3.406$; H = $4.58 / 1.01 = 4.535$; O = $54.51 / 16.00 = 3.407$ [1] Divide by the smallest (3.406): 1.000 : 1.331 : 1.000 [1] Recognise 1.331 as a third and multiply all by 3: 3 : 4 : 3 [1] Empirical formula = $C_3H_4O_3$ [1] <i>Do not credit $CH_{1.33}O$. Rounding 1.331 to 1 loses the last two marks.</i>	4
(c)	Empirical formula mass = $3(12.01) + 4(1.01) + 3(16.00) = 88.07$; $n = 176.14 / 88.07 = 2.00$ [1] Molecular formula = $C_6H_8O_6$ [1]	2
(d)	500 mg = 0.500 g [1] $n = 0.500 / 176.14 = 2.84 \times 10^{-3} \text{ mol}$ [1] $N = 2.84 \times 10^{-3} \times 6.02 \times 10^{23} = 1.71 \times 10^{21} \text{ molecules}$ [1] <i>Award zero for the first mark if 500 is used without converting from mg to g.</i>	3
(e) HL	The empirical formula mass is 88.07, and the compound's M_r is also 88.07, so $n = 88.07 / 88.07 = 1$ and the molecular formula is $C_3H_4O_3$ [1] The empirical formula gives only the <i>ratio</i> of atoms, not how many atoms are present or how they are bonded together; molecular size and functional groups — and therefore chemical properties — can differ completely [1] OWTTE	2

SELF-ASSESSMENT: WHAT YOUR SCORE IS TELLING YOU

SECTION E1 (MCQ)	DIAGNOSIS AND NEXT STEP
17–20	Secure. Move on to Structure 1.4.5 (concentration) and 1.4.6 (Avogadro's law), then Reactivity 2.1.
13–16	The chemistry is sound but entity-counting or bracket-reading is slipping. Re-do Solved Examples 2, 3 and 10 from a blank page.
9–12	Re-read Part A in full and rebuild the mole road map from memory before attempting Sections E2 and E3.
0–8	Go back to first principles: work through $n = m/M$ and $N = n N_A$ on ten substances of your own choosing before touching exam questions.