



AQA CHEMISTRY 7405

AMOUNT OF SUBSTANCE II

EQUATIONS • REACTING MASSES • FORMULAE • HYDRATED SALTS

			
SPEC	CRACK IT	TEST IT	DONE WHEN

Your next step after Amount of Substance I

Use moles to connect substances in a reaction and to turn experimental masses into chemical formulae. Each section teaches a method, shows the working and gives you a different problem to solve.

AQA SECTION	THIS PACK TEACHES
3.1.2.5, selected content	Full and ionic equations; mole ratios; reacting masses; limiting reactants; linked reactions.
3.1.2.4	Empirical formulae from mass or percentage composition; molecular formulae using relative molecular mass.
3.1.2.4–3.1.2.5, practical applications	Metal oxides, combustion data and water of crystallisation; using and evaluating mass measurements.

BRING FORWARD $n = m \div M$; $m = n \times M$. Amount n is in mol, mass m in g and molar mass M in g mol^{-1} . Relative masses A_r and M_r have no unit.

PART III Concentration, titrations, the ideal gas equation, percentage yield and atom economy are taught in Amount of Substance III.

Use a calculator. Show your route and units. Unless a question states otherwise, give calculated masses to three significant figures and formula subscripts as whole numbers. Keep unrounded values during working.

Useful A_r values: H 1.0; C 12.0; N 14.0; O 16.0; Na 23.0; Mg 24.3; Al 27.0; P 31.0; S 32.1; Cl 35.5; Ca 40.1; Fe 55.8; Cu 63.5; Zn 65.4; Ba 137.3.

Specification checked 23 September 2026. Original practice questions; worked answers at the back.

1. Balance atoms without changing formulae



SPEC 3.1.2.5 Balance full equations for familiar and unfamiliar reactions when the formulae are supplied.



CRACK IT Keep every chemical formula fixed. Change only the numbers in front, then recount every element on both sides.

TERM	MEANING	EXAMPLE
Coefficient	A number before a formula; multiplies the whole formula.	3H ₂ O contains 6 H atoms and 3 O atoms.
Subscript	A small number within a formula; identifies the substance.	H ₂ O and H ₂ O ₂ are different compounds.
Balanced equation	Has the same number of atoms of each element on each side.	2H ₂ + O ₂ → 2H ₂ O
State symbol	Shows the physical state or an aqueous solution.	(s) solid; (l) liquid; (g) gas; (aq) dissolved in water.

Worked example 1 Aluminium and oxygen



- O occurs in groups of 2 and 3. Use their lowest common multiple, 6: put 3 before O₂ and 2 before Al₂O₃.
- The product now contains 4 Al atoms. Put 4 before Al.
- Check: Al 4 → 4; O 6 → 6. The balanced equation is 4Al(s) + 3O₂(g) → 2Al₂O₃(s).

Worked example 2 Count atoms inside brackets

Ca(OH)₂ + HCl → CaCl₂ + H₂O becomes Ca(OH)₂ + 2HCl → CaCl₂ + 2H₂O. There are 1 Ca, 2 O, 4 H and 2 Cl on each side. The 2 outside (OH) counts both O and H.

FINAL CHECK Use the smallest whole-number coefficients. Never turn O₂ into O₃ to make an equation balance.



TEST IT 1

a) Balance Fe + O₂ → Fe₂O₃ using the smallest whole-number coefficients. [2 marks]

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b) Balance C₂H₆ + O₂ → CO₂ + H₂O using the smallest whole-number coefficients. [2 marks]

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c) Explain why changing Fe₂O₃ to FeO is not a valid way to balance the first equation. [1 mark]

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✓ **DONE WHEN** you balance each element by changing coefficients, while preserving every formula.

2. Write and check ionic equations



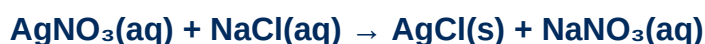
SPEC 3.1.2.5 Write full and ionic equations and conserve both atoms and total charge.



CRACK IT An ionic equation keeps the species that actually change. Remove spectator ions that appear unchanged on both sides.

TERM	LEARNABLE MEANING
Full equation	Shows the formulae of all reactants and products.
Ionic equation	Shows the reacting species after unchanged spectator ions have been removed.
Spectator ion	An ion present in the same form on both sides of the full ionic equation.
Precipitate	An insoluble solid formed when solutions react.

Worked example 1 Forming silver chloride



1. Separate the dissolved ions: $\text{Ag}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$.
2. Cancel Na^+ and NO_3^- . They have not changed.
3. Net ionic equation: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$. Total charge is 0 on both sides.

Worked example 2 Strong acid and alkali

$\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$. Cancel Na^+ and Cl^- after writing the dissolved ions. The result is $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$. Check 2 H, 1 O and total charge 0 on each side.

WHAT TO SPLIT For these examples, separate soluble ionic compounds, strong acids and strong alkalis in aqueous solution. Keep solids, liquids, gases and weak acids intact. “(aq)” alone does not mean a substance should be split into ions.



TEST IT 2

a) For $\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{NaCl}(\text{aq})$, write the full ionic equation, cancel the spectator ions and write the net ionic equation with state symbols. [5 marks]

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b) Explain why BaSO_4 must remain together in the ionic equation. [1 mark]

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✓ **DONE WHEN** your ionic equation keeps the reacting species and balances atoms, charges and states.

3. Read coefficients as mole ratios



SPEC 3.1.2.5 Use the coefficients in a balanced equation to link amounts of different substances.



CRACK IT A balanced equation gives ratios of particles and therefore moles. It does not give the same ratios of masses.

Stoichiometry means the quantitative relationships between reactants and products in a balanced equation. Name the substance beside every mole value so you can see which ratio you are using.

$$n(\text{wanted}) = n(\text{known}) \times \text{coefficient}(\text{wanted}) \div \text{coefficient}(\text{known})$$

Worked example 1 Hydrogen to ammonia



N_2	H_2	NH_3
1 mol	3 mol	2 mol
0.0800 mol	0.240 mol	0.160 mol

If 0.240 mol H_2 reacts completely with excess N_2 , $n(\text{NH}_3) = 0.240 \times 2/3 = 0.160$ mol. "Excess" means more N_2 is present than this amount of H_2 needs.

Worked example 2 Aluminium to oxygen

$4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$. For 0.400 mol Al, $n(\text{O}_2)$ needed = $0.400 \times 3/4 = 0.300$ mol. The target coefficient goes on top: oxygen 3, aluminium 4.

MASS IS DIFFERENT In $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$, the reacting masses are $28.0 \text{ g} + 6.0 \text{ g} \rightarrow 34.0 \text{ g}$, not $1 \text{ g} + 3 \text{ g} \rightarrow 2 \text{ g}$. Multiply each mole amount by its own molar mass.



TEST IT 3

a) $\text{P}_4 + 6\text{Cl}_2 \rightarrow 4\text{PCl}_3$. Calculate the amount of Cl_2 needed and the amount of PCl_3 formed when 0.120 mol P_4 reacts completely. [3 marks]

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b) A student says this equation means 1 g P_4 reacts with 6 g Cl_2 . Explain the error. [1 mark]

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✓ **DONE WHEN** you multiply by the wanted coefficient and divide by the known coefficient.

4. Calculate a reacting mass



SPEC 3.1.2.5 Use a balanced equation to calculate a mass of product or a mass of reactant needed.



CRACK IT Use a grid: known mass ↓ known moles → target moles ↑ target mass. Follow the numbered cells; the equation ratio connects only the moles row.

Put the balanced equation across the top and mass, relative mass and moles down the left. Blue cells contain givens; green cells contain the answer. Use A_r for single atoms and M_r for molecules or formula units. These relative masses have no unit; their numerical values equal molar masses in g mol^{-1} .

Mass in g → moles: $n = m \div M$ | Moles → mass in g: $m = n \times M$

Worked example 1 Follow cells 1 to 6

Calculate the mass of MgO formed from 4.86 g Mg with excess oxygen. Write 4.86 in the Mg mass cell first; mark the MgO mass as the unknown. Then fill the numbered cells in this order.

Balanced equation	2Mg	+ O ₂	→ 2MgO
Mass / g	4.86	In excess	Unknown mass
M_r / A_r	Look up A_r	32.0	Calculate M_r
Moles / mol	To calculate	Not needed	To calculate

Follow the numbered cells and green arrows: down the known column, across the moles row, then up the target column. M is the molar mass in g mol^{-1} .

Balanced equation	2Mg	+ O ₂	→ 2MgO
Mass / g	[1] 4.86	In excess	[6] 8.06
	↓		↑ × M
M_r / A_r	[2] 24.3	32.0	[5] 40.3
	↓ ÷ M		↑
Moles / mol	[3] 0.200	× 2/2 →	[4] 0.200

[1] Write the known mass: 4.86 g Mg.

[2] Look up $A_r(\text{Mg}) = 24.3$ and put it in the Mg relative-mass cell.

[3] ↓ Down the Mg column: $n = 4.86 \div 24.3 = 0.200 \text{ mol}$.

[4] → Across the moles row: $n(\text{MgO}) = 0.200 \times 2/2 = 0.200 \text{ mol}$.

[5] Calculate $M_r(\text{MgO}) = 24.3 + 16.0 = 40.3$ and fill its relative-mass cell.

[6] ↑ Up the MgO column: $m = 0.200 \times 40.3 = 8.06 \text{ g}$.

$n(\text{target}) = n(\text{known}) \times \text{target coefficient} \div \text{known coefficient}$

4 continued Apply the reacting mass grid

Worked example 2 Use a 2:1 mole ratio

Calculate the mass of O_2 released when 6.80 g H_2O_2 decomposes completely. The equation shows that 2 mol H_2O_2 produce 1 mol O_2 .

Balanced equation	$2H_2O_2$	$\rightarrow 2H_2O$	$+ O_2$
Mass / g	[1] 6.80	Not needed	[6] 3.20
	↓		↑ $\times M$
M_r / A_r	[2] 34.0	18.0	[5] 32.0
	↓ $\div M$		↑
Moles / mol	[3] 0.200	$\times 1/2$ →	[4] 0.100

Down: $6.80 \div 34.0 = 0.200$ mol H_2O_2 . Across: $0.200 \times 1/2 = 0.100$ mol O_2 . Up: $0.100 \times 32.0 = 3.20$ g O_2 . The 2:1 ratio halves the moles; it does not halve the mass.



TEST IT 4

Complete each grid using the question data. Use the lines for the mole-ratio calculation and final mass in g.

a) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$. Calculate the mass of Fe formed from 7.98 g Fe_2O_3 with excess CO. $M(Fe_2O_3) = 159.6$ and $M(Fe) = 55.8 \text{ g mol}^{-1}$. [3 marks]

Balanced equation	Fe_2O_3	$+ 3CO$	$\rightarrow 2Fe$	$+ 3CO_2$
Mass / g	7.98	In excess		Not needed
M_r / A_r				
Moles / mol		Not needed		Not needed

b) $Zn + 2HCl \rightarrow ZnCl_2 + H_2$. Calculate the mass of HCl needed to react completely with 0.654 g Zn. $M(Zn) = 65.4$ and $M(HCl) = 36.5 \text{ g mol}^{-1}$. [3 marks]

Balanced equation	Zn	$+ 2HCl$	$\rightarrow ZnCl_2$	$+ H_2$
Mass / g	0.654		Not needed	Not needed
M_r / A_r				
Moles / mol			Not needed	Not needed

✓ **DONE WHEN** you change substance at the mole-ratio step, not directly between masses.

5. Work backwards and find a sample's purity



SPEC 3.1.2.5 Use product data to calculate the starting mass and the mass of a reacting component in a sample.



CRACK IT Start with the quantity you have been given, even if it is a product. The same mass → moles → ratio → mass route works backwards.

Worked example 1 Start from a known product mass

Find the Fe_2O_3 needed to make 11.16 g Fe. Begin in the Fe column even though it is on the product side. Al is in excess.

Balanced equation	Fe_2O_3	+ 2Al	→ 2Fe	+ Al_2O_3
Mass / g	[6] 15.96 ↑ × M	In excess	[1] 11.16 ↓	Not needed
M_r / A_r	[5] 159.6 ↑	27.0	[2] 55.8 ↓ ÷ M	102.0
Moles / mol	[4] 0.1000 × 1/2 ←		[3] 0.2000	Not needed

$11.16 \div 55.8 = 0.2000 \text{ mol Fe}$; $\times 1/2 = 0.1000 \text{ mol Fe}_2\text{O}_3$; $\times 159.6 = 15.96 \text{ g}$ (16.0 g to 3 significant figures).

Worked example 2 Only part of the sample reacts

Purity is the percentage by mass of the named substance in a sample. A 5.00 g impure CaCO_3 sample releases 1.76 g CO_2 on complete decomposition. The impurities do not release gas.

Balanced equation	CaCO_3	→ CaO	+ CO_2
Mass / g	[6] 4.004 pure ↑ × M	Not needed	[1] 1.76 ↓
M_r / A_r	[5] 100.1 ↑	56.1	[2] 44.0 ↓ ÷ M
Moles / mol	[4] 0.0400 × 1/1 ←		[3] 0.0400

CO_2 : $1.76 \div 44.0 = 0.0400 \text{ mol}$. The 1:1 ratio gives 0.0400 mol CaCO_3 ; $\times 100.1$ gives 4.004 g of pure CaCO_3 .

Purity = mass of pure substance ÷ sample mass × 100

$4.004 \div 5.00 \times 100 = 80.1\%$. The 5.00 g sample mass belongs in this final calculation, not the CaCO_3 mass cell.



TEST IT 5

a) $2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$. Calculate the mass of NaHCO_3 required to form 5.30 g Na_2CO_3 . $M(\text{NaHCO}_3) = 84.0$ and $M(\text{Na}_2\text{CO}_3) = 106.0 \text{ g mol}^{-1}$. [3 marks]

b) A 4.00 g impure MgCO_3 sample releases 1.54 g CO_2 on complete decomposition: $\text{MgCO}_3 \rightarrow \text{MgO} + \text{CO}_2$. The impurities do not release gas. Calculate the percentage purity. $M(\text{MgCO}_3) = 84.3$ and $M(\text{CO}_2) = 44.0 \text{ g mol}^{-1}$. [4 marks]

✓ **DONE WHEN** you can reverse a mole ratio and distinguish the reacting mass from the total sample mass.

6. Find the limiting reactant first



SPEC 3.1.2.5 Calculate reacting masses when amounts of two reactants are supplied.



CRACK IT Fill both reactant columns before choosing a starting point. Add an $n \div$ coefficient row; the smaller value identifies the limiting reactant.

The limiting reactant is completely consumed first and limits the product formed. A reactant in excess is supplied in more than the required amount, so some remains.

Worked example Aluminium and chlorine

Mix 2.70 g Al and 7.10 g Cl_2 . Assume complete reaction. Fill mass $\rightarrow$ relative mass $\rightarrow$ available moles for both reactants before calculating product.

Balanced equation	2Al	+ 3Cl ₂	$\rightarrow$ 2AlCl ₃
Mass / g	2.70 supplied	7.10 supplied	8.90 formed
M _r / A _r	27.0	71.0	133.5
Moles / mol	0.100 available	0.100 available	0.066666... formed
$n \div$ coefficient	$0.100 \div 2$ = 0.0500	$0.100 \div 3$ = 0.033333...	Use the smaller reactant value

Product moles = smallest (reactant $n \div$ coefficient) $\times$ product coefficient

- Down both columns: $2.70 \div 27.0 = 0.100$ mol Al; $7.10 \div 71.0 = 0.100$ mol Cl_2 .
- Compare $0.100/2$ with $0.100/3$. Cl_2 is limiting: its value is smaller.
- Across: $n(\text{AlCl}_3) = (0.100 \div 3) \times 2 = 0.066666...$ mol. Up: $\times 133.5 = 8.90$ g.
- Al used = $(0.100 \div 3) \times 2 = 0.066666...$ mol. Al left = $0.100 - 0.066666... = 0.033333...$ mol; $\times 27.0 = 0.900$ g.

COMPARE MOLES PER COEFFICIENT Equal available moles do not necessarily mean both reactants are used up. Do not compare grams or moles alone. Use the limiting reactant to calculate what actually reacts.



TEST IT 6

Use mass $\rightarrow$ moles $\rightarrow$ $n \div$ coefficient for both reactants. Use the limiting reactant to calculate the product and the excess remaining.

a) $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$. A mixture contains 1.20 g H_2 and 6.40 g O_2 . Determine the limiting reactant, showing a numerical comparison. M values: H_2 2.0; O_2 32.0; H_2O 18.0 g mol⁻¹. [3 marks]

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b) Calculate the mass of H_2O formed and the mass of H_2 left when the limiting reactant is completely consumed. [3 marks]

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✓ **DONE WHEN** you identify the limiting reactant before calculating product or excess.

7. Find an empirical formula from composition



SPEC 3.1.2.4 Calculate empirical formulae from masses and percentage composition.



CRACK IT Use one column per element. Fill mass → A_r → moles → divide by the smallest → whole-number ratio. Here the ratio comes from your data, not a balanced equation.

FORMULA	DEFINITION	EXAMPLE
Empirical	The simplest whole-number ratio of atoms of each element in a compound.	Hydrogen peroxide: HO
Molecular	The actual number of atoms of each element in one molecule.	Hydrogen peroxide: H_2O_2

$$n(\text{element}) = \text{mass} \div A_r \quad | \quad \text{Ratio} = \text{each } n \div \text{smallest } n$$

Worked example 1 Mass data

A compound contains 0.486 g Mg and 0.320 g O. Fill the grid downwards. Divide every mole amount by 0.0200.

Element	Mg	O
Mass / g	0.486	0.320
A_r	24.3	16.0
Moles / mol	0.0200	0.0200
Ratio ÷ 0.0200	1	1

The simplest whole-number ratio is 1:1, so the empirical formula is MgO.

Worked example 2 Percentage data

40.00% C, 6.67% H and 53.33% O by mass: assume a 100 g sample, so the percentages become masses in g.

Element	C	H	O
Mass / g	40.00	6.67	53.33
A_r	12.0	1.0	16.0
Moles / mol	3.3333...	6.6700	3.333125
Ratio ÷ 3.333125	≈ 1	≈ 2	1

Empirical formula = CH_2O . If the ratio is 1:1.5, multiply every value by 2 to obtain 2:3; do not round 1.5 to 2.



TEST IT 7

For both parts, show an element grid, mole amounts, ratio, any whole-number scaling and the empirical formula.

a) A compound contains 1.08 g Al and 0.960 g O. Calculate its empirical formula. A_r : Al 27.0; O 16.0. [4 marks]

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b) A compound contains 69.9% Fe and 30.1% O by mass. Calculate its empirical formula. A_r : Fe 55.8; O 16.0. [4 marks]

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✓ **DONE WHEN** you turn masses or percentages into the simplest whole-number atom ratio.

8. Deal with ratios that are not whole numbers



SPEC 3.1.2.4 Convert fractional mole ratios into valid whole-number empirical formulae.



CRACK IT A value close to 1.5 is not a badly rounded 1 or 2. Multiply the entire ratio by a small whole number.

RATIO AFTER DIVIDING	MULTIPLY EVERY VALUE BY	WHOLE-NUMBER RATIO
1 : 1.5	2	2 : 3
1 : 1.333...	3	3 : 4
1 : 1.25	4	4 : 5
1 : 1.667...	3	3 : 5

Small differences such as 1.99 versus 2.00 can come from measurement precision. A clear fraction such as 1.50 requires scaling. Use all the data; do not force an arbitrary ratio just because one number is awkward.

Worked example Phosphorus oxide

The composition is 43.6% P and 56.4% O. A_r : P 31.0; O 16.0.

Element	P	O
Mass / g	43.6	56.4
A_r	31.0	16.0
Moles / mol	1.40645...	3.525
Ratio $\div 1.40645...$	1	2.506... ≈ 2.5
Whole-number ratio	2	5

Multiply the whole ratio by 2: 1 : 2.5 $\rightarrow$ 2 : 5

Empirical formula = P_2O_5 . This is the simplest ratio; it does not yet establish the molecular formula.

THE WHOLE RATIO CHANGES Writing 1:5 doubles only the oxygen value and changes the ratio. Scale every column by the same number.



TEST IT 8

Use the element-grid method for part (a), showing the scaling that makes all ratios whole numbers.

a) A compound contains 0.960 g C, 0.200 g H and 0.960 g O. Calculate its empirical formula. A_r : C 12.0; H 1.0; O 16.0. [4 marks]

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b) An experiment gives a metal:oxygen mole ratio of 1.00:1.49. State the likely whole-number ratio and explain how you obtained it. [2 marks]

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✓ **DONE WHEN** you recognise simple fractions and scale every part of the ratio together.

9. Use mass gain to find a metal oxide formula



SPEC 3.1.2.4; AT a, k Apply empirical-formula calculations to experimental metal-oxide data.



CRACK IT The mass gained during oxidation is the mass of oxygen added. Subtract the container before calculating the metal mass.

Worked example Copper heated in oxygen

MEASUREMENT	MASS / g
Empty crucible	24.800
Crucible + copper before heating	26.070
Crucible + copper oxide at constant mass	26.390

Subtract the container first: copper = $26.070 - 24.800 = 1.270$ g. Oxygen gained = $26.390 - 26.070 = 0.320$ g. Enter these masses in the element grid.

Element	Cu	O
Mass / g	1.270	0.320
A_r	63.5	16.0
Moles / mol	0.0200	0.0200
Ratio $\div 0.0200$	1	1

The 1:1 ratio gives CuO. During oxidation the metal gains oxygen from the surroundings; total mass is conserved when that oxygen is included.

A PRACTICAL END POINT Heat, cool and weigh, then repeat. Constant mass means successive readings agree within the balance resolution. The calculation assumes the final solid is the oxide and no solid has been lost.

INCOMPLETE OXIDATION The measured oxygen gain is too small. The calculated oxygen:metal ratio is too low; more heating may still increase the mass.



TEST IT 9

a) An empty crucible has mass 18.400 g. With iron before heating it has mass 20.074 g. With iron oxide at constant mass it has mass 20.794 g. Calculate the empirical formula of the oxide. A_r : Fe 55.8; O 16.0. [4 marks]

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b) Explain the source of the extra mass and predict how incomplete oxidation would affect the calculated oxygen:iron mole ratio. [3 marks]

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✓ **DONE WHEN** you isolate the metal mass and oxygen gain before calculating the formula.

10. Use M_r to find the molecular formula



SPEC 3.1.2.4 Calculate a molecular formula from an empirical formula and relative molecular mass.



CRACK IT A molecular formula is a whole-number multiple of the empirical formula. Use M_r to find that multiplier.

$$\text{multiplier} = M_r \div \text{empirical formula mass}$$

Multiply every empirical-formula subscript by the multiplier. Empirical formula mass means the sum of A_r values in the empirical formula; it has no unit. This step applies to molecular compounds.

Worked example 1 CH_2O with $M_r = 180.0$

1. Empirical formula mass = $12.0 + 2(1.0) + 16.0 = 30.0$.
2. Multiplier = $180.0 \div 30.0 = 6$.
3. $\text{CH}_2\text{O} \times 6$ gives $\text{C}_6\text{H}_{12}\text{O}_6$. Check its M_r : $72.0 + 12.0 + 96.0 = 180.0$.

Worked example 2 P_2O_5 with $M_r = 284.0$

Empirical formula mass = $2(31.0) + 5(16.0) = 142.0$. Multiplier = $284.0/142.0 = 2$. Molecular formula = P_4O_{10} . Both subscripts double.

FROM PERCENTAGES First find the empirical formula by the mass $\rightarrow$ moles $\rightarrow$ ratio method. Only then use M_r to find the molecular formula. Recheck the data if the multiplier is not close to a whole number.



TEST IT 10

a) A molecular compound has empirical formula CHCl and $M_r = 97.0$. Calculate its molecular formula. A_r : C 12.0; H 1.0; Cl 35.5. [3 marks]

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b) A molecular compound contains 54.5% C, 9.1% H and 36.4% O by mass and has $M_r = 88.0$. Determine its empirical formula and molecular formula. A_r : C 12.0; H 1.0; O 16.0. [4 marks]

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✓ **DONE WHEN** you multiply every subscript and check the final molecular mass.

11. Find a formula from combustion data



SPEC 3.1.2.4, applied data Use composition evidence from combustion products to calculate empirical and molecular formulae.



CRACK IT In complete combustion, each CO_2 contains one C atom and each H_2O contains two H atoms. Recover the amounts of C and H in the original sample.

MEASURED PRODUCT	AMOUNT IN ORIGINAL SAMPLE
CO_2	$n(\text{C}) = n(\text{CO}_2) = \text{mass of } \text{CO}_2 \div 44.0$
H_2O	$n(\text{H}) = 2 \times n(\text{H}_2\text{O}) = 2 \times \text{mass of } \text{H}_2\text{O} \div 18.0$
O in a compound containing C, H and O only	Find its mass by difference: sample mass – mass of C – mass of H.

Worked example A 0.460 g sample containing C, H and O only

Complete combustion produces 0.880 g CO_2 and 0.540 g H_2O . Find the empirical formula.

- C from CO_2 : $0.880/44.0 = 0.0200$ mol C; $\times 12.0$ gives 0.240 g C.
- H from H_2O : $0.540/18.0 = 0.0300$ mol H_2O ; $\times 2$ gives 0.0600 mol H, or 0.0600 g.
- O by difference: $0.460 - 0.240 - 0.0600 = 0.160$ g. Put the element masses in the grid.

Element	C	H	O
Mass / g	0.240	0.0600	0.160
A_r	12.0	1.0	16.0
Moles / mol	0.0200	0.0600	0.0100
Ratio $\div 0.0100$	2	6	1

Empirical formula = $\text{C}_2\text{H}_6\text{O}$.

DO NOT COUNT THE PRODUCT OXYGEN Combustion adds oxygen from outside the sample. The oxygen in CO_2 and H_2O is not all from the original compound. Use mass by difference only when the original sample contains C, H and O only.



TEST IT 11

a) Complete combustion of 0.900 g of a compound containing C, H and O only produces 1.320 g CO_2 and 0.540 g H_2O . Calculate its empirical formula. Use $M(\text{CO}_2) = 44.0$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$; A_r : C 12.0; H 1.0; O 16.0. [5 marks]

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b) The same compound has $M_r = 90.0$. Determine its molecular formula. [2 marks]

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✓ **DONE WHEN** you obtain C from CO_2 , H from twice the water amount and original O by mass difference.

12. Find the water of crystallisation



SPEC 3.1.2.5, hydrated-salt application Use a hydrated salt formula and calculate the number of waters per formula unit.



CRACK IT Separate the mass of water from the mass of anhydrous salt. Convert both to moles, then compare water with salt.

TERM	MEANING
Water of crystallisation	Water molecules present in a fixed ratio as part of a crystalline compound.
Hydrated salt	A crystalline salt containing water of crystallisation, for example $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.
Anhydrous salt	The salt without water of crystallisation, for example CuSO_4 .
The dot and x	$\text{Salt} \cdot x\text{H}_2\text{O}$ means x water molecules per salt formula unit. The dot is not a covalent bond.

Worked example Copper sulfate hydrate

A 2.496 g hydrate sample leaves 1.596 g anhydrous CuSO_4 after all its water is removed. $M(\text{CuSO}_4) = 159.6$; $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$.

Water lost = $2.496 - 1.596 = 0.900 \text{ g}$. Grid the anhydrous salt and water as two components; divide by salt moles.

Component	CuSO_4	H_2O
Mass / g	1.596	0.900
M_r	159.6	18.0
Moles / mol	0.0100	0.0500
Ratio $\div n(\text{salt})$	1	5

$$x = n(\text{water}) \div n(\text{anhydrous salt}) = 0.0500 \div 0.0100 = 5$$

Formula = $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. Its relative formula mass is $159.6 + 5(18.0) = 249.6$.

HYDRATED IS NOT THE SAME AS AQUEOUS $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$ is a solid hydrate. $\text{CuSO}_4(\text{aq})$ describes dissolved copper sulfate; “(aq)” does not specify five waters per formula unit.



TEST IT 12

a) A 2.191 g sample of $\text{CaCl}_2 \cdot x\text{H}_2\text{O}$ leaves 1.111 g anhydrous CaCl_2 . Calculate x and state the full hydrate formula. $M(\text{CaCl}_2) = 111.1$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$. [4 marks]

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b) Calculate the relative formula mass of $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$. The relative formula masses of Na_2CO_3 and H_2O are 106.0 and 18.0. [2 marks]

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✓ **DONE WHEN** you use moles of water divided by moles of anhydrous salt to find x.

13. Use crucible measurements correctly



SPEC 3.1.2.5; AT a, k Calculate a hydrate formula from heating-to-constant-mass measurements.



CRACK IT Balance readings include the crucible. First subtract it to obtain the hydrate and residue masses; then find the water lost.

Worked example Magnesium sulfate hydrate

MEASUREMENT	MASS / g
Empty crucible	19.840
Crucible + hydrate	22.304
After first heating and cooling	21.068
After second and third heating/cooling	21.044; 21.044

Use the constant mass 21.044 g. Hydrate = $22.304 - 19.840 = 2.464$ g; anhydrous salt = $21.044 - 19.840 = 1.204$ g; water = $22.304 - 21.044 = 1.260$ g.

Component	MgSO ₄	H ₂ O
Mass / g	1.204	1.260
M _r	120.4	18.0
Moles / mol	0.0100	0.0700
Ratio ÷ n(salt)	1	7

$x = 7$, giving MgSO₄·7H₂O.

THE PRACTICAL SEQUENCE Heat gently at first to reduce spitting. Cool before weighing to avoid hot-air currents and protect the balance. Repeat heat → cool → weigh to constant mass. The calculation assumes all loss is water, with no salt lost or decomposed.

Data for TEST IT 13

EMPTY CRUCIBLE / g	WITH HYDRATE / g	SUCCESSIVE COOLED READINGS / g
27.360	29.803	29.459; 29.443; 29.443



TEST IT 13

a) Use the measurements above to calculate x in BaCl₂· x H₂O. M(BaCl₂) = 208.3 and M(H₂O) = 18.0 g mol⁻¹. Show the masses and both mole amounts. [4 marks]

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b) Explain why the final reading is used and why the crucible should be cooled before each weighing. [2 marks]

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✓ **DONE WHEN** you use the final cooled constant mass and remove the container from the calculation.

14. Explain errors in a hydrate experiment



SPEC 3.1.2.5, practical evaluation Explain how incomplete heating or loss of solid changes the calculated hydration number.



CRACK IT Track the effect on both measured quantities: water mass comes from mass loss, while salt moles come from the residue.

WHAT HAPPENS	RESIDUE MASS	APPARENT WATER LOSS	CALCULATED x
Some water remains	Too high	Too low	Too low
Residue absorbs water before weighing	Too high	Too low	Too low
Some solid spits out	Too low	Too high	Too high
Salt decomposes and also loses another gas	Too low	Too high	Too high

Worked example Incomplete water removal

A 1.722 g $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ sample should leave 1.362 g CaSO_4 . Suppose 0.090 g water remains, so the measured residue is 1.452 g. $M(\text{CaSO}_4) = 136.2 \text{ g mol}^{-1}$.

1. Apparent water loss = $1.722 - 1.452 = 0.270 \text{ g}$, so apparent $n(\text{water}) = 0.0150 \text{ mol}$.
2. Treating all residue as salt gives apparent $n(\text{salt}) = 1.452/136.2 = 0.010661\dots \text{ mol}$.
3. Calculated $x = 0.0150/0.010661\dots = 1.41$. It is too low because water loss is underestimated and salt moles are overestimated.

CONSTANT MASS HAS LIMITS A stable reading alone does not prove that the residue is the intended anhydrous salt. The material must be heated under suitable conditions, and loss of solid or decomposition must be avoided.

Do not report a fractional hydration number such as 1.41 as if it were a reliable hydrate formula. Identify the experimental problem rather than forcing the answer to the expected integer.



TEST IT 14

a) A student stops heating $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$ while water still remains. Explain whether the calculated x is too high or too low. [2 marks]

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b) Another student loses some solid from the crucible by spitting. Explain the effect on the calculated x. [2 marks]

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c) Explain why reaching constant mass does not, by itself, prove that only water has been removed. [2 marks]

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✓ **DONE WHEN** you link an experimental error to mass loss, residue mass and the direction of the error in x.

15. Link two reactions through an intermediate



SPEC 3.1.2.5 Use two balanced equations to calculate a mass in a linked process.



CRACK IT Carry the moles of the intermediate from the first reaction into the second. Apply each equation ratio separately.

Worked example Limestone to calcium chloride



Find the mass of HCl needed to react with all the CaO formed from 10.01 g CaCO_3 . Assume complete decomposition and no loss of CaO. $M(\text{CaCO}_3) = 100.1$ and $M(\text{HCl}) = 36.5 \text{ g mol}^{-1}$.

Linked substances	CaCO_3	CaO intermediate	HCl target
Mass / g	[1] 10.01 ↓	Not needed	[7] 7.30 ↑ × M
M_r / A_r	[2] 100.1 ↓ ÷ M	56.1	[6] 36.5 ↑
Moles / mol	[3] 0.1000 × 1/1 →	[4] 0.1000 × 2/1 →	[5] 0.2000

$10.01 \div 100.1 \rightarrow 0.1000 \text{ mol CaCO}_3 \rightarrow \times 1/1 \rightarrow 0.1000 \text{ mol CaO} \rightarrow \times 2/1 \rightarrow 0.2000 \text{ mol HCl}$

Use the two equations above to cross the moles row in two steps. Finish: $0.2000 \times 36.5 = 7.30 \text{ g HCl}$.

NAME THE INTERMEDIATE CaO is produced in one reaction and used in the next. You do not need to convert its moles into grams and back again.

Equations for TEST IT 15



A 4.20 g sample of NaHCO_3 decomposes completely. All of the Na_2CO_3 formed is then reacted with HCl. M values: NaHCO_3 84.0; Na_2CO_3 106.0; HCl 36.5 g mol^{-1} .



TEST IT 15

a) Calculate the mass of Na_2CO_3 formed in the first reaction. [3 marks]

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b) Calculate the mass of HCl needed to react with all of this Na_2CO_3 . Show the mole ratio from the second equation. [3 marks]

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✓ **DONE WHEN** you follow the named intermediate through both mole ratios.

16. Apply formula ratios in unfamiliar problems



SPEC 3.1.2.4–3.1.2.5 Combine empirical-formula and hydrated-salt methods to solve unfamiliar mass problems.



CRACK IT Identify what is unknown. If the formula ratio is given, work backwards to a molar mass. If percentages are given, choose a 100 g sample.

Worked example 1 Identify a metal from MCl_2

A chloride MCl_2 contains 0.486 g metal M and 1.420 g Cl. Start from Cl; use the formula ratio $\text{M}:\text{Cl} = 1:2$ to work back to the metal.

Element in MCl_2	M	Cl
Mass / g	0.486	1.420
A_r	24.3	35.5
Moles / mol	0.0200	0.0400
Formula ratio	1	2

Cl: $1.420 \div 35.5 = 0.0400$ mol. M: $\times 1/2 = 0.0200$ mol. $A_r(\text{M}) = 0.486 \div 0.0200 = 24.3$, identifying Mg.

Worked example 2 Water percentage in a hydrate

$\text{CuSO}_4 \cdot x\text{H}_2\text{O}$ contains 36.06% water by mass. Assume 100 g, then grid the salt and water.

Component	CuSO_4	H_2O
Mass / g	63.94	36.06
M_r	159.6	18.0
Moles / mol	0.400626...	2.00333...
Ratio $\div n(\text{salt})$	1	5.0006... ≈ 5

CHOOSE THE RATIO SOURCE An equation links different substances in a reaction. A formula links elements or water to salt within one compound. Decide which relationship supplies your ratio.



TEST IT 16

a) A 2.160 g sample of metal M forms 4.080 g of oxide M_2O_3 . Determine the molar mass of M and identify it as Al, Fe or Cu. A_r : O 16.0; Al 27.0; Fe 55.8; Cu 63.5. Show the oxygen mass and the formula-ratio step. [5 marks]

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b) $\text{MgCl}_2 \cdot x\text{H}_2\text{O}$ contains 53.13% water by mass. Calculate x using a 100 g sample. $M(\text{MgCl}_2) = 95.3$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$. [4 marks]

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✓ **DONE WHEN** you choose the relevant ratio and work in either direction.

17. Check the method and the final answer



SPEC 3.1.2.4–3.1.2.5; MS 1.1 Select and check methods for equations, reacting masses and experimental formulae.



CRACK IT Before using a calculator, identify the target and the source of the ratio. At the end, check units, atoms, charge and numerical precision.

IF THE QUESTION ASKS FOR...	USE THIS ROUTE
A mass of another substance	Known mass → known moles → equation ratio → wanted moles → wanted mass.
A limiting reactant	Convert both masses to moles. Compare n/coefficient.
An empirical formula	Element masses ÷ A_r → divide by smallest → whole-number ratio.
A molecular formula	Find empirical formula mass; use M_r to obtain the multiplier.
Water of crystallisation	Water lost and anhydrous residue → both mole amounts → water/salt.

Worked correction Do not round in the middle

If a mole amount is 0.066666... mol and $M = 133.5 \text{ g mol}^{-1}$, keep the full value: mass = 8.90 g. Rounding moles to 0.067 first gives 8.9445 g and changes the final answer.

MASTERY CHECK	CHECK YOUR WORK
Equations	Formulae fixed; coefficients correct; atoms and net charge conserved.
Calculations	Each amount names its substance; masses in g; final mass has a unit.
Formulae	All subscripts are whole numbers; the empirical ratio is the simplest.
Measurements	Container subtracted; final cooled mass used; assumptions stated.
Precision	Round once at the end. Coefficients and subscripts are exact counts.

NEXT Use these same mole-ratio habits in Amount of Substance III, where solution concentrations, gas measurements, yield and atom economy are introduced.



TEST IT 17

a) Correct this statement: "The 2:1 ratio in a balanced equation means the reacting masses are in the ratio 2:1." [1 mark]

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b) Correct this empirical-formula step: "The mole ratio is 1:1.5, so round it to 1:2." [1 mark]

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c) A compound has empirical formula CH_2 and $M_r = 56.0$. A student writes C_4H_2 . Give the correct molecular formula and explain the correction. A_r : C 12.0; H 1.0. [2 marks]

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d) State which two mole amounts form the ratio used to find x in $\text{salt} \cdot x\text{H}_2\text{O}$, and explain why the empty crucible mass is needed. [2 marks]

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✓ **DONE WHEN** you can choose, explain and check the method without copying a worked example.

Answers for TEST IT 1 to 4

Use these worked answers to check the method as well as the final value. Mark allocations are a guide for these original questions.

TEST IT 1 | 5 marks

- a) $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$. Fe: 4 each side; O: 6 each side. [2]
 b) $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$. C: 4; H: 12; O: 14 on each side. [2]
 c) It changes the chemical identity/formula of the product instead of changing the amount of that product. [1]

TEST IT 2 | 6 marks

- a) Full ionic: $\text{Ba}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) + 2\text{Na}^{+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{Na}^{+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq})$. [2]
 Cancel Na^{+} and Cl^{-} . [1]
 Net: $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$. [2]
 b) It is an insoluble solid, not separate dissolved ions. [1]

TEST IT 3 | 4 marks

- a) $n(\text{Cl}_2) = 0.120 \times 6 = 0.720 \text{ mol}$. [1]
 $n(\text{PCl}_3) = 0.120 \times 4 = 0.480 \text{ mol}$. [1]
 Correct use of the 1:6:4 mole ratio. [1]
 b) The coefficients give a mole/particle ratio, not a mass ratio; the substances have different molar masses. [1]

TEST IT 4 | 6 marks

a) Calculation grid

Balanced equation	Fe_2O_3	+ 3CO	$\rightarrow 2\text{Fe}$	+ 3CO ₂
Mass / g	[1] 7.98	In excess	[6] 5.58	Not needed
	↓		↑ × M	
M _r / A _r	[2] 159.6	28.0	[5] 55.8	44.0
	↓ ÷ M		↑	
Moles / mol	[3] 0.0500	× 2/1 →	[4] 0.100	Not needed

- a) $n(\text{Fe}_2\text{O}_3) = 7.98/159.6 = 0.0500 \text{ mol}$. [1]
 $n(\text{Fe}) = 2 \times 0.0500 = 0.100 \text{ mol}$. [1]
 $m(\text{Fe}) = 0.100 \times 55.8 = 5.58 \text{ g}$. [1]

b) Calculation grid

Balanced equation	Zn	+ 2HCl	$\rightarrow \text{ZnCl}_2$	+ H ₂
Mass / g	[1] 0.654	[6] 0.730	Not needed	Not needed
	↓	↑ × M		
M _r / A _r	[2] 65.4	[5] 36.5	136.4	2.0
	↓ ÷ M	↑		
Moles / mol	[3] 0.0100	× 2/1 →	[4] 0.0200	Not needed

- b) $n(\text{Zn}) = 0.654/65.4 = 0.0100 \text{ mol}$. [1]
 $n(\text{HCl}) = 2 \times 0.0100 = 0.0200 \text{ mol}$. [1]
 $m(\text{HCl}) = 0.0200 \times 36.5 = 0.730 \text{ g}$. [1]

Answers for TEST IT 5 to 6

TEST IT 5 | 7 marks

a) Calculation grid

Balanced equation	2NaHCO ₃	→ Na ₂ CO ₃	+ CO ₂	+ H ₂ O
Mass / g	[6] 8.40	[1] 5.30	Not needed	Not needed
	↑ × M	↓		
M _r / A _r	[5] 84.0	[2] 106.0	44.0	18.0
	↑	↓ ÷ M		
Moles / mol	[4] 0.100	[3] 0.0500	Not needed	Not needed
		← × 2/1		

a) $n(\text{Na}_2\text{CO}_3) = 5.30/106.0 = 0.0500 \text{ mol}$. [1]

$n(\text{NaHCO}_3) = 2 \times 0.0500 = 0.100 \text{ mol}$. [1]

$m = 0.100 \times 84.0 = 8.40 \text{ g}$. [1]

b) Calculation grid

Balanced equation	MgCO ₃	→ MgO	+ CO ₂
Mass / g	[6] 2.9505 pure	Not needed	[1] 1.54
	↑ × M		↓
M _r / A _r	[5] 84.3	40.3	[2] 44.0
	↑		↓ ÷ M
Moles / mol	[4] 0.0350	[3] 0.0350	
		← × 1/1	

b) $n(\text{CO}_2) = 1.54/44.0 = 0.0350 \text{ mol}$. [1]

$n(\text{MgCO}_3) = 0.0350 \text{ mol}$. [1]

Pure MgCO₃ mass = $0.0350 \times 84.3 = 2.9505 \text{ g}$. [1]

Purity = $2.9505/4.00 \times 100 = 73.8\%$. [1]

TEST IT 6 | 6 marks

a) Calculation grid

Balanced equation	2H ₂	+ O ₂	→ 2H ₂ O
Mass / g	1.20 supplied	6.40 supplied	7.20 formed
M _r	2.0	32.0	18.0
Moles / mol	0.600 available	0.200 available	0.400 formed
n ÷ coefficient	0.300	0.200: limiting	Use 0.200 × 2

a) $n(\text{H}_2) = 1.20/2.0 = 0.600 \text{ mol}$; $n(\text{O}_2) = 6.40/32.0 = 0.200 \text{ mol}$. [1]

n/coefficient values are 0.300 and 0.200. [1]

O₂ limits the reaction. [1]

b) $n(\text{H}_2\text{O}) = 0.200 \times 2 = 0.400 \text{ mol}$, so water mass = 7.20 g. [2]

H₂ used = 0.400 mol; H₂ left = $0.600 - 0.400 = 0.200 \text{ mol} = 0.400 \text{ g}$. [1]

Answers for TEST IT 7 to 8

TEST IT 7 | 8 marks

a) Calculation grid

Element	Al	O
Mass / g	1.08	0.960
A_r	27.0	16.0
Moles / mol	0.0400	0.0600
Ratio $\div 0.0400$	1	1.5
Whole-number ratio	2	3

a) $n(\text{Al}) = 1.08/27.0 = 0.0400 \text{ mol}$; $n(\text{O}) = 0.960/16.0 = 0.0600 \text{ mol}$. [2]

Ratio 1:1.5, then multiply both by 2 $\rightarrow$ 2:3. [1]

Al_2O_3 . [1]

b) Calculation grid

Element	Fe	O
Mass / g	69.9	30.1
A_r	55.8	16.0
Moles / mol	1.25269...	1.88125
Ratio $\div 1.25269...$	1	1.502... $\approx$ 1.5
Whole-number ratio	2	3

b) For 100 g: $n(\text{Fe}) = 69.9/55.8 = 1.2527...$; $n(\text{O}) = 30.1/16.0 = 1.88125$. [2]

Ratio 1:1.502... $\approx$ 1:1.5, so 2:3. [1]

Fe_2O_3 . [1]

TEST IT 8 | 6 marks

a) Calculation grid

Element	C	H	O
Mass / g	0.960	0.200	0.960
A_r	12.0	1.0	16.0
Moles / mol	0.0800	0.200	0.0600
Ratio $\div 0.0600$	1.333...	3.333...	1
Whole-number ratio	4	10	3

a) Amounts: C 0.0800 mol; H 0.200 mol; O 0.0600 mol. [2]

Divide by 0.0600 $\rightarrow$ 1.333...:3.333...:1. Multiply all by 3 $\rightarrow$ 4:10:3. [1]

$\text{C}_4\text{H}_{10}\text{O}_3$. [1]

b) 1.00:1.49 is close to 1:1.5. [1]

Multiply both by 2 $\rightarrow$ 2:3. [1]

Answers for TEST IT 9 to 10

TEST IT 9 | 7 marks

a) Calculation grid

Element	Fe	O
Mass / g	1.674	0.720
A_r	55.8	16.0
Moles / mol	0.0300	0.0450
Ratio $\div 0.0300$	1	1.5
Whole-number ratio	2	3

a) Fe mass = 1.674 g; O mass = 0.720 g. [1]

$n(\text{Fe}) = 0.0300 \text{ mol}$; $n(\text{O}) = 0.0450 \text{ mol}$. [1]

Ratio 1:1.5 $\rightarrow$ 2:3. [1]

Fe_2O_3 . [1]

b) Oxygen from the air/oxygen supply joins the iron. [1]

Incomplete oxidation gives too little mass gain. [1]

The calculated oxygen:iron ratio is too low. [1]

TEST IT 10 | 7 marks

a) Empirical formula mass = $12.0 + 1.0 + 35.5 = 48.5$. [1]

Multiplier = $97.0/48.5 = 2$. [1]

Molecular formula $\text{C}_2\text{H}_2\text{Cl}_2$. [1]

b) Calculation grid

Element	C	H	O
Mass / g	54.5	9.1	36.4
A_r	12.0	1.0	16.0
Moles / mol	4.54166...	9.100	2.275
Ratio $\div 2.275$	≈ 2	4	1

b) For 100 g: $54.5/12.0 = 4.5417...$, $9.1/1.0 = 9.1$, $36.4/16.0 = 2.275$. [1]

Divide by 2.275 $\rightarrow$ approximately 2:4:1, so empirical formula $\text{C}_2\text{H}_4\text{O}$. [1]

Empirical formula mass = 44.0; multiplier $88.0/44.0 = 2$. [1]

Molecular formula $\text{C}_4\text{H}_8\text{O}_2$. [1]

Answers for TEST IT 11 to 12

TEST IT 11 | 7 marks

a) Calculation grid

Element	C	H	O
Mass / g	0.360	0.0600	0.480
A_r	12.0	1.0	16.0
Moles / mol	0.0300	0.0600	0.0300
Ratio $\div 0.0300$	1	2	1

a) $n(\text{C}) = 1.320/44.0 = 0.0300 \text{ mol}$. [1]

$n(\text{H}) = 2 \times 0.540/18.0 = 0.0600 \text{ mol}$. [1]

Masses C 0.360 g and H 0.0600 g; O mass = $0.900 - 0.360 - 0.0600 = 0.480 \text{ g}$. [1]

$n(\text{O}) = 0.0300 \text{ mol}$; ratio 1:2:1. [1]

CH_2O . [1]

b) Empirical formula mass = 30.0; multiplier = $90.0/30.0 = 3$. [1]

Molecular formula $\text{C}_3\text{H}_6\text{O}_3$. [1]

TEST IT 12 | 6 marks

a) Calculation grid

Component	CaCl_2	H_2O
Mass / g	1.111	1.080
M_r	111.1	18.0
Moles / mol	0.0100	0.0600
Ratio $\div n(\text{salt})$	1	6

a) Water mass = 1.080 g. [1]

$n(\text{water}) = 1.080/18.0 = 0.0600 \text{ mol}$; $n(\text{CaCl}_2) = 1.111/111.1 = 0.0100 \text{ mol}$. [1]

$x = 6$. [1]

$\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$. [1]

b) $106.0 + 10(18.0)$ [1]

= 286.0, with no unit. [1]

Answers for TEST IT 13 to 15

TEST IT 13 | 6 marks

a) Calculation grid

Component	BaCl ₂	H ₂ O
Mass / g	2.083	0.360
M _r	208.3	18.0
Moles / mol	0.0100	0.0200
Ratio ÷ n(salt)	1	2

a) Residue = 29.443 – 27.360 = 2.083 g; water = 29.803 – 29.443 = 0.360 g. [1]

$n(\text{BaCl}_2) = 2.083/208.3 = 0.0100$ mol. [1]

$n(\text{H}_2\text{O}) = 0.360/18.0 = 0.0200$ mol. [1]

$x = 2$; BaCl₂·2H₂O. [1]

b) The final two readings agree, showing constant mass within the balance resolution. [1]

Cooling prevents hot-air effects/unstable readings and protects the balance. [1]

TEST IT 14 | 6 marks

a) Too low. [1]

Apparent water loss is too small and the residue, treated as anhydrous salt, gives too many salt moles. [1]

b) Too high. [1]

Solid loss is counted as water loss and the remaining salt mass/moles are too small. [1]

c) A salt could decompose and release another gas, or some solid could have been lost. [1]

The remaining material can still reach a stable mass, so identity of the residue and the assumption about water loss must be valid. [1]

TEST IT 15 | 6 marks

a) Calculation grid

Balanced equation	2NaHCO ₃	→ Na ₂ CO ₃	+ CO ₂	+ H ₂ O
Mass / g	[1] 4.20 ↓	[6] 2.65 ↑ × M	Not needed	Not needed
M _r / A _r	[2] 84.0 ↓ ÷ M	[5] 106.0 ↑	44.0	18.0
Moles / mol	[3] 0.0500 × 1/2 →	[4] 0.0250	Not needed	Not needed

a) $n(\text{NaHCO}_3) = 4.20/84.0 = 0.0500$ mol. [1]

$n(\text{Na}_2\text{CO}_3) = 0.0500 \times 1/2 = 0.0250$ mol. [1]

Mass = 0.0250 × 106.0 = 2.65 g. [1]

b) Calculation grid

Balanced equation	Na ₂ CO ₃	+ 2HCl	→ 2NaCl	+ CO ₂	+ H ₂ O
Mass / g	[1] 2.65 ↓	[6] 1.825 ↑ × M	Not needed	Not needed	Not needed
M _r	[2] 106.0 ↓ ÷ M	[5] 36.5 ↑	58.5	44.0	18.0
Moles / mol	[3] 0.0250 × 2/1 →	[4] 0.0500	Not needed	Not needed	Not needed

b) Na₂CO₃:HCl = 1:2. [1]

$n(\text{HCl}) = 0.0250 \times 2 = 0.0500$ mol. [1]

Mass = 0.0500 × 36.5 = 1.825 g → 1.83 g. [1]

Answers for TEST IT 16 to 17

TEST IT 16 | 9 marks

a) Calculation grid

Element in M_2O_3	M	O
Mass / g	2.160	1.920
A_r	27.0	16.0
Moles / mol	0.0800	0.120
Formula ratio	2	3

a) O mass = $4.080 - 2.160 = 1.920$ g. [1]
 $n(\text{O atoms}) = 1.920/16.0 = 0.1200$ mol. [1]
 $n(\text{M}) = 0.1200 \times 2/3 = 0.0800$ mol. [1]
Molar mass = $2.160/0.0800 = 27.0$ g mol⁻¹. [1]
Metal is Al. [1]

b) Calculation grid

Component	MgCl ₂	H ₂ O
Mass / g	46.87	53.13
M_r	95.3	18.0
Moles / mol	0.491815...	2.951666...
Ratio ÷ n(salt)	1	6.0016... ≈ 6

b) For 100 g: water 53.13 g and MgCl₂ 46.87 g. [1]
 $n(\text{water}) = 2.95167...$ mol. [1]
 $n(\text{salt}) = 46.87/95.3 = 0.491815...$ mol. [1]
 $x = 6.0016... \approx 6$; MgCl₂·6H₂O. [1]

TEST IT 17 | 6 marks

a) The coefficients give a mole/particle ratio. Masses depend on the molar masses as well. [1]

b) Multiply both values by 2 to obtain 2:3. [1]

c) Empirical formula mass = 14.0; multiplier = 4. Multiply both subscripts by 4. [1]
C₄H₈. [1]

d) $x = \text{moles of water} / \text{moles of anhydrous salt}$. [1]
Subtracting the crucible gives the residue/salt mass rather than container plus salt. [1]

Specification and next steps

AQA 3.1.2.4 is covered in sections 7–11 and applied in sections 12–16. The full/ionic-equation and reacting-mass content of 3.1.2.5 is covered in sections 1–6 and 15–16. Hydrated salts and heating to constant mass are practical applications. Sections 3.1.2.2 and MS 1.1 provide the mole and numerical-precision foundations.

The remaining solution, gas, percentage-yield and atom-economy content is reserved for Amount of Substance III.

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