



Amount of Substance II

OCR A Chemistry A H432 | Specification-led student notes

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SPECIFICATION COVERAGE

OCR A REF	THIS PACK TEACHES	TAUGHT / TESTED
2.1.2(b)	Balanced chemical and ionic equations, with state symbols	Sections 1–2
2.1.3(e)(i), (g)	Mole ratios, reacting masses and limiting reactants	Sections 3–4
2.1.3(b), (c)	Empirical and molecular formulae from mass, percentage and elemental analysis data	Sections 5–7
2.1.3(d)	Hydrated salts, anhydrous salts and water of crystallisation	Sections 8–9
2.1.3(j) in part; 1.1.3(a), (c); 1.1.4(c)	Measurement of mass, interpretation of experimental results and appropriate significant figures	Sections 9–10
Combined application	Mixed exam-style practice with complete worked answers	Tests 11–14

YOUR AIM Use a balanced equation or a chemical formula to turn measured masses into amounts, then into a justified answer.

LESSON BOUNDARY Builds on Amount of Substance I. Solution concentration, titration calculations, gas calculations, percentage yield and atom economy belong to Amount of Substance III.

Mapped to OCR A H432 specification version 3.1 (May 2026)

The big picture



YOUR ROUTE Write what you know, identify what you need, then use amounts in mol to connect the two.

SYMBOL	MEANING AND UNIT
m	Mass of the named substance, in g
M	Molar mass of the named substance, in g mol^{-1}
n	Amount of the named substance, in mol
A_r	Relative atomic mass; no unit
M_r	Relative molecular mass; no unit
x	Number of waters of crystallisation per formula unit of salt; no unit

THE GRID METHOD

Known mass → **known amount** → **required amount** → **required mass**

Put substances across the top and mass, relative mass and moles down the left. Fill the knowns, calculate the first amount, apply the ratio and work up the required column.

The M_r / A_r^* row has no units. Use A_r for atoms, M_r for molecules and relative formula mass for ionic compounds. Its numerical value is the molar mass M in g mol^{-1} .

$$n = m \div M \quad | \quad m = n \times M$$

QUESTION TYPE	THE RATIO YOU NEED
Reacting masses	Coefficients in the balanced equation
Empirical formula	Amounts of atoms of each element
Molecular formula	$M_r \div \text{empirical formula mass}$
Hydrated salt	Amount of water $\div$ amount of anhydrous salt



EXAM CHECK Label each amount, for example $n(\text{CO}_2)$. Keep calculator digits until the final answer. Use the question's precision and state the final unit. Formula subscripts are whole numbers, not measured quantities.



✓ DONE WHEN You can choose the required ratio and distinguish M from M_r before calculating.

1. Balanced equations conserve atoms



SPEC 2.1.2(b) Construct balanced equations and use the appropriate state symbols.



CRACK IT Change the numbers in front of formulae. Never change a formula's subscripts to balance an equation.

A coefficient multiplies the whole formula. In $2\text{Al}_2\text{O}_3$ there are four Al atoms and six O atoms. The equation must contain the same number of atoms of each element on each side.

State symbols: (s) solid • (l) liquid • (g) gas • (aq) dissolved in water

WORKED EXAMPLE

QUESTION Aluminium reacts with oxygen to form solid aluminium oxide, Al_2O_3 . Write a balanced equation with state symbols.

START $\text{Al(s)} + \text{O}_2\text{(g)} \rightarrow \text{Al}_2\text{O}_3\text{(s)}$

OXYGEN Use 3O_2 and $2\text{Al}_2\text{O}_3$ to give six O atoms on each side.

ALUMINIUM $2\text{Al}_2\text{O}_3$ contains four Al atoms, so use 4Al .

ANSWER $4\text{Al(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)}$



EXAM CHECK Aqueous means dissolved in water. It does not mean a pure liquid. Include state symbols when the question requests them or the reaction information supplies them.



TEST IT 1

(a) Solid magnesium burns in oxygen gas to form solid magnesium oxide. Write a balanced equation, including state symbols. [2]

(b) Solid calcium carbonate decomposes on heating into solid calcium oxide and carbon dioxide gas. Write a balanced equation, including state symbols. [2]

[4 marks]



✓ DONE WHEN You can balance an equation without changing the substances involved.

2. Ionic equations conserve charge



SPEC 2.1.2(b) Construct ionic equations with state symbols and check atoms and overall charge.



CRACK IT Show the species that change. Cancel spectator ions that remain unchanged in solution.

For these reactions, split aqueous ionic compounds into their ions. Keep solids, liquids and gases together. An aqueous molecular substance is not automatically split into ions.

WORKED EXAMPLE

QUESTION Aqueous silver nitrate reacts with aqueous sodium chloride to form solid silver chloride and aqueous sodium nitrate. Write the full equation and the ionic equation.

FULL $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$

IONS $\text{Ag}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$

CANCEL Na^+ and NO_3^- are unchanged spectator ions.

ANSWER $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$

CHECK One Ag and one Cl on each side; total charge is zero on each side.

WHAT TO CHECK

- 1 Every atom balances.
- 2 The sum of the charges balances. A neutral product can form from oppositely charged ions.
- 3 Formulae, ionic charges and state symbols are shown correctly.



TEST IT 2

Aqueous barium chloride reacts with aqueous sodium sulfate. Solid barium sulfate and aqueous sodium chloride form.

(a) Write the balanced full equation using BaCl_2 , Na_2SO_4 , BaSO_4 and NaCl , with state symbols. [2]

(b) Write the ionic equation with state symbols. [2]

(c) Name both spectator ions. [1]

[5 marks]



✓ **DONE WHEN** You can remove spectator ions and conserve both atoms and charge.

3. Use a grid for reacting masses



SPEC 2.1.3(e)(i), (g) Use the balanced equation to link masses through amounts in mol.



CRACK IT Go down the known column, across using the mole ratio, then up the required column.

WORKED EXAMPLE

QUESTION What mass of Al_2O_3 forms when 5.40 g Al reacts completely with excess O_2 ? Use $M(\text{Al}) = 27.0$ and $M(\text{Al}_2\text{O}_3) = 102.0 \text{ g mol}^{-1}$.



A Record the knowns. Put ? in the cells you must calculate.

QUANTITY	Al	Al_2O_3
Mass, m / g	5.40	?
M_r / A_r^* (no unit)	27.0	102.0
Moles, n / mol	?	?

B Follow cells 1 to 6. The red arrows show the order of completion.

	Al	Al_2O_3
Mass, m / g	① 5.40	⑥ $0.100 \times 102.0 = 10.2$ $m = n \times M$
M_r / A_r^* (no unit)	② 27.0	⑤ 102.0
Moles, n / mol	③ $5.40 \div 27.0 = 0.200$ $n = m \div M$	④ $0.200 \times 2 \div 4 = 0.100$ Use the 4 : 2 mole ratio

Move between columns using the equation ratio

O_2 is in excess, so its amount is not needed. *For Al use A_r ; for ionic Al_2O_3 use relative formula mass. Use the same numerical values as M in g mol^{-1} in the highlighted equations.



TEST IT 3

Fe_2O_3 reacts completely with excess CO: $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$.
Calculate the mass of Fe from 7.98 g Fe_2O_3 . Give 3 significant figures.
 $M(\text{Fe}_2\text{O}_3) = 159.6$; $M(\text{Fe}) = 55.8 \text{ g mol}^{-1}$. [4]

[4 marks]



✓ **DONE WHEN** You can go down, across and up without treating a mole ratio as a mass ratio.

4. Use a grid for limiting reactants



SPEC 2.1.3(e)(i), (g) Compare the available amounts with the balanced equation before calculating product.



CRACK IT Fill both reactant columns first. Add an $n \div$ coefficient row. The smaller value identifies the limiting reactant.

WORKED EXAMPLE

QUESTION 2.40 g Mg reacts with 0.960 g O_2 . Identify the limiting reactant and calculate the maximum mass of MgO. $M(Mg) = 24.3$, $M(O_2) = 32.0$ and $M(MgO) = 40.3 \text{ g mol}^{-1}$.



QUANTITY	Mg available	O_2 available	MgO formed
1 Mass / g	2.40	0.960	Step 5: 0.0600×40.3 = 2.42
2 M_r / A_r^* (no unit)	24.3	32.0	40.3
3 Moles / mol	$2.40 \div 24.3$ = 0.098765...	$0.960 \div 32.0$ = 0.0300	Step 4: 0.0300×2 = 0.0600
Compare $n \div$ coefficient	$0.098765... \div 2$ = 0.049382...	$0.0300 \div 1$ = 0.0300	Not compared

FOLLOW THE GRID

DOWN BOTH COLUMNS Use $n = m \div M$ to find the available reactant amounts.

COMPARE O_2 gives the smaller $n \div$ coefficient, so O_2 is limiting.

ACROSS THEN UP $O_2 : MgO = 1 : 2$. Find $n(MgO)$, then use $m = n \times M$. Answer: 2.42 g.



EXAM CHECK The reactant columns show amounts available, not necessarily amounts used. Never apply the product ratio to all of an excess reactant.



TEST IT 4

4.00 g H_2 reacts with 28.0 g N_2 : $N_2 + 3H_2 \rightarrow 2NH_3$.

$M(H_2) = 2.0$, $M(N_2) = 28.0$ and $M(NH_3) = 17.0 \text{ g mol}^{-1}$.

(a) Identify the limiting reactant, showing your comparison. [3]

(b) Calculate the maximum mass of NH_3 formed. [2]

[5 marks]



✓ DONE WHEN You can compare $n \div$ coefficient, then calculate product from the limiting reactant.

5. Use an element grid for empirical formulae



SPEC 2.1.3(b)(i), (c) Use composition by mass to determine a compound's simplest whole-number atom ratio.



CRACK IT Put elements across the top. Go down: mass → A_r → moles → divide by the smallest → whole numbers.

An empirical formula gives the simplest whole-number ratio of atoms of each element. Here the relative-mass row uses A_r , because each column represents atoms of one element.

WORKED EXAMPLE

QUESTION A 1.116 g sample of Fe combines only with oxygen to form 1.596 g of an oxide. Determine its empirical formula. Use $A_r(\text{Fe}) = 55.8$ and $A_r(\text{O}) = 16.0$.

First find the missing mass: $m(\text{O}) = 1.596 - 1.116 = 0.480 \text{ g}$.

FILL DOWN ↓	Fe	O
1 Mass / g	1.116	0.480
2 A_r (no unit)	55.8	16.0
3 Moles / mol	$1.116 \div 55.8$ $= 0.0200$	$0.480 \div 16.0$ $= 0.0300$
4 ÷ smallest (0.0200)	$0.0200 \div 0.0200 = 1$	$0.0300 \div 0.0200 = 1.5$
5 Whole numbers (multiply both by 2)	2	3

READ THE BOTTOM ROW

Fe : O = 2 : 3, so the empirical formula is Fe_2O_3 .

The atom ratio comes from the measured amounts. Do not use coefficients from a reaction equation.



EXAM CHECK Do not round 1.5 to 2. Multiply the whole ratio: halves $\times 2$, thirds $\times 3$, quarters $\times 4$. A value such as 1.99 can reasonably represent 2 when the data are rounded.



TEST IT 5

A compound contains 1.458 g Mg and 0.560 g N. Determine its empirical formula.
 $A_r(\text{Mg}) = 24.3$; $A_r(\text{N}) = 14.0$. [4]

[4 marks]



✓ **DONE WHEN** You can use the bottom row of the grid as the formula subscripts.

6. Use a grid for percentage composition



SPEC 2.1.3(b)(ii), (c) Calculate empirical and molecular formulae from percentage composition and M_r .



CRACK IT Assume a 100 g sample. The percentages become the masses in the first row.

A molecular formula gives the actual atom counts in one molecule. First find the empirical formula from the grid, then use M_r to find the whole-number multiplier.

WORKED EXAMPLE

QUESTION A compound is 54.5% C, 9.1% H and 36.4% O by mass. Its M_r is 88.0. Determine its empirical and molecular formulae. Use $A_r(\text{C}) = 12.0$, $A_r(\text{H}) = 1.0$ and $A_r(\text{O}) = 16.0$.

FILL DOWN ↓	C	H	O
1 Mass / g (in 100 g)	54.5	9.1	36.4
2 A_r (no unit)	12.0	1.0	16.0
3 Moles / mol	$54.5 \div 12.0$ $= 4.54166\dots$	$9.1 \div 1.0$ $= 9.1$	$36.4 \div 16.0$ $= 2.275$
4 $\div$ smallest (2.275)	≈ 2	4	1
5 Simplest ratio	2	4	1

FROM THE GRID TO THE MOLECULE

EMPIRICAL FORMULA $\text{C}_2\text{H}_4\text{O}$. Empirical formula mass $= 24.0 + 4.0 + 16.0 = 44.0$.

MULTIPLIER $M_r \div$ empirical formula mass $= 88.0 \div 44.0 = 2$.

MOLECULAR FORMULA Multiply every subscript by 2 $\rightarrow \text{C}_4\text{H}_8\text{O}_2$. Check: $M_r = 88.0$.



TEST IT 6

A molecular compound is 40.0% C, 6.7% H and 53.3% O by mass, with $M_r = 180.0$. Find its empirical formula, then calculate its molecular formula. $A_r(\text{C}) = 12.0$, $A_r(\text{H}) = 1.0$ and $A_r(\text{O}) = 16.0$. [5]

[5 marks]



DONE WHEN You can fill the mass row from percentages, then use M_r to find the molecular formula.

7. Put elemental analysis into a grid



SPEC 2.1.3(c) Deduce formulae from elemental analysis measurements.



CRACK IT Convert the measured products into masses or amounts of elements, then fill the element grid.

WHAT THE PRODUCTS TELL YOU

$n(\text{C atoms}) = n(\text{CO}_2)$ | $n(\text{H atoms}) = 2 \times n(\text{H}_2\text{O})$

For a sample containing only C, H and O, find $m(\text{O})$ by subtracting $m(\text{C})$ and $m(\text{H})$ from the sample mass.

WORKED EXAMPLE

QUESTION Complete combustion of 0.460 g of a compound containing only C, H and O gives 0.880 g CO_2 and 0.540 g H_2O . Find its empirical formula. $M(\text{CO}_2) = 44.0$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$; $A_r(\text{C}) = 12.0$, $A_r(\text{H}) = 1.0$ and $A_r(\text{O}) = 16.0$.

CARBON MASS $(0.880 \div 44.0) \times 12.0 = 0.240 \text{ g}$

HYDROGEN MASS $(0.540 \div 18.0) \times 2 \times 1.0 = 0.0600 \text{ g}$

OXYGEN MASS $0.460 - 0.240 - 0.0600 = 0.160 \text{ g}$

ELEMENT GRID	C	H	O
Mass, m / g	0.240	0.0600	0.160
A_r (no unit)	12.0	1.0	16.0
Moles, n / mol	$0.240 \div 12.0$ $= 0.0200$	$0.0600 \div 1.0$ $= 0.0600$	$0.160 \div 16.0$ $= 0.0100$
$\div$ smallest (0.0100)	2	6	1

The bottom row gives $\text{C}_2\text{H}_6\text{O}$. Do not count oxygen supplied by the combustion gas as oxygen from the original compound.



TEST IT 7

Complete combustion of a hydrocarbon gives 1.760 g CO_2 and 0.720 g H_2O . Determine its empirical formula.

$M(\text{CO}_2) = 44.0$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$; $A_r(\text{C}) = 12.0$ and $A_r(\text{H}) = 1.0$. [4]

[4 marks]



✓ **DONE WHEN** You can convert product data into an element grid and read off the simplest ratio.

8. Use a ratio grid for hydrated salts



SPEC 2.1.3(d) Use hydrated, anhydrous and water of crystallisation, and calculate a hydrated salt formula.



CRACK IT Put anhydrous salt and water in separate columns. Divide both amounts by the amount of salt.

TERM	MEANING
Hydrated salt	Contains water of crystallisation.
Water of crystallisation	Water molecules incorporated into a crystal structure in a fixed ratio to the salt.
Anhydrous salt	Contains no water of crystallisation.

$\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ has x water molecules per MgSO_4 formula unit. The dot separates the salt from its water of crystallisation.

WORKED EXAMPLE

QUESTION $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ contains 51.2% water by mass. Calculate x . $M(\text{MgSO}_4) = 120.4$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$.

FILL DOWN ↓	MgSO_4	H_2O
Mass / g (in 100 g)	$100 - 51.2 = 48.8$	51.2
M_r / A_r^* (no unit)	120.4	18.0
Moles, n / mol	$48.8 \div 120.4$ $= 0.405315\dots$	$51.2 \div 18.0$ $= 2.84444\dots$
÷ amount of salt	1	$2.84444\dots \div 0.405315\dots$ $= 7.01785\dots \approx 7$

READ THE RATIO

$\text{MgSO}_4 : \text{H}_2\text{O} = 1 : 7 \rightarrow \text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.

Use the relative formula mass of anhydrous MgSO_4 in its column. The formula of the hydrate is the unknown.



TEST IT 8

$\text{CuSO}_4 \cdot x\text{H}_2\text{O}$ contains 36.1% water by mass. Calculate x and give the full formula.
 $M(\text{CuSO}_4) = 159.6$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$. [4]

[4 marks]



✓ **DONE WHEN** You can use a salt : water mole-ratio grid to find x .

9. Put heating results into a hydrate grid



SPEC 2.1.3(d), (i) Use experimental masses to determine a hydrated salt formula.



CRACK IT Find the two masses before filling the grid: residue minus container; initial mass minus final mass.

Heat a suitable hydrate gently at first. Cool before weighing. Repeat heating, cooling and weighing to constant mass within the balance's resolution. Avoid overheating that decomposes the salt.

WORKED EXAMPLE

QUESTION Determine x in $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$. Only water is lost. Empty crucible = 25.000 g; crucible + hydrate = 27.496 g. Successive heat-cool readings are 26.622 g, 26.596 g and 26.596 g. $M(\text{CuSO}_4) = 159.6$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$.

FILL DOWN ↓	CuSO_4	H_2O
Mass, m / g	$26.596 - 25.000$ $= 1.596$	$27.496 - 26.596$ $= 0.900$
M_r / A_r^* (no unit)	159.6	18.0
Moles, n / mol	$1.596 \div 159.6$ $= 0.0100$	$0.900 \div 18.0$ $= 0.0500$
÷ amount of salt (0.0100)	1	5

ANSWER

$\text{CuSO}_4 : \text{H}_2\text{O} = 1 : 5 \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

Use the final constant reading. The initial hydrated mass is not the mass of anhydrous salt.



TEST IT 9

$\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ is heated to constant mass. Only water is lost.

Empty crucible = 18.650 g; crucible + hydrate = 21.114 g; crucible + anhydrous residue = 19.854 g.

Determine x and the full formula. Show the two mass subtractions.

$M(\text{MgSO}_4) = 120.4$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$. [6]

[6 marks]



✓ DONE WHEN You can turn the experimental mass readings into a salt : water mole-ratio grid.

10. Explain experimental errors



SPEC 2.1.3(d), (i); 1.1.4(c) Interpret mass measurements and explain how an error affects a calculated hydrate formula.



CRACK IT Follow the effect on both masses, then on the ratio $x = n(\text{water}) \div n(\text{salt})$.

WHAT HAPPENS	EFFECT ON THE CALCULATION	EFFECT ON x
Some water remains after heating	Residue mass is too high; calculated water loss is too low. The residue is incorrectly treated as anhydrous salt.	Too low
Solid is lost by spitting	Residue mass is too low; calculated water loss is too high because lost solid is counted as water.	Too high
Residue absorbs water before weighing	Final mass is too high; apparent water loss is too low.	Too low

WORKED EXAMPLE

QUESTION A hydrated salt is not heated for long enough, leaving some water in the residue. Explain why the calculated value of x is too low.

NUMERATOR The measured mass loss is too small, so calculated $n(\text{H}_2\text{O})$ is too small.

DENOMINATOR The residue contains water, so its mass overestimates the mass and amount of anhydrous salt.

CONCLUSION A smaller numerator divided by a larger denominator gives a value of x that is too low.



EXAM CHECK "Human error" is not an explanation. Name a specific cause and trace its effect. Constant mass supports complete water removal, but it does not prove that no salt has decomposed.



TEST IT 10

(a) A student loses some solid by spitting when heating a hydrate, but assumes all mass loss is water. Explain the effect on calculated x . [3]

(b) Explain why the student should repeat heating, cooling and weighing rather than take one final reading. [2]

[5 marks]



✓ **DONE WHEN** You can connect a specific experimental error to the direction of error in x .

Mixed practice 1



PROVE IT Choose the method yourself. Keep full calculator values until the final answer and show your working.



TEST IT 11

Aluminium reacts with iron(III) oxide: $2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}$.

A mixture contains 2.70 g Al and 16.0 g Fe_2O_3 . The reaction goes to completion.

Use $M(\text{Al}) = 27.0$, $M(\text{Fe}_2\text{O}_3) = 159.6$ and $M(\text{Fe}) = 55.8 \text{ g mol}^{-1}$.

(a) Determine the limiting reactant, showing your comparison. [3]

(b) Calculate the maximum mass of Fe formed. [2]

(c) Calculate the mass of the excess reactant remaining. Give calculated masses to 3 significant figures. [3]

[8 marks]



TEST IT 12

A molecular compound contains 30.4% N and 69.6% O by mass and has $M_r = 92.0$. Determine its empirical and molecular formulae. Use $A_r(\text{N}) = 14.0$ and $A_r(\text{O}) = 16.0$. [5]

[5 marks]

Mixed practice 2



PROVE IT Identify which mass is water, which mass is salt and which ratio comes from the equation.



TEST IT 13

A sample of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ is heated until only anhydrous Na_2CO_3 remains. No solid is lost.
Empty crucible = 31.400 g; crucible + hydrated salt = 34.260 g; crucible + residue at constant mass = 32.460 g.
Calculate x and give the full formula. Show how you obtain the mass of water and the mass of Na_2CO_3 . Use $M(\text{Na}_2\text{CO}_3) = 106.0$ and $M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1}$. [6]

[6 marks]



TEST IT 14

Potassium chlorate(V) decomposes: $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$.
Calculate the mass of KClO_3 that must decompose completely to produce 0.960 g O_2 . Give your answer to 3 significant figures. Use $M(\text{KClO}_3) = 122.6$ and $M(\text{O}_2) = 32.0 \text{ g mol}^{-1}$. [4]

[4 marks]

Answers and working 1

Original CrackingChem practice. Marks are indicative; equivalent correct working is acceptable. Values are shown in the units stated in the row labels.

Test It 1 | 4 marks

MODEL ANSWER

- (a) $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$ [1 for balanced formulae; 1 for states]
 (b) $\text{CaCO}_3\text{(s)} \rightarrow \text{CaO(s)} + \text{CO}_2\text{(g)}$ [1 for balanced formulae; 1 for states]

Test It 2 | 5 marks

MODEL ANSWER

- (a) $\text{BaCl}_2\text{(aq)} + \text{Na}_2\text{SO}_4\text{(aq)} \rightarrow \text{BaSO}_4\text{(s)} + 2\text{NaCl(aq)}$ [1 for balanced formulae; 1 for states]
 (b) $\text{Ba}^{2+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)} \rightarrow \text{BaSO}_4\text{(s)}$ [1 for correct balanced ions; 1 for states]
 (c) Na^+ and Cl^- are the spectator ions. Both required. [1]

Test It 3 | 4 marks



QUANTITY	Fe_2O_3	Fe
Mass, m / g	7.98	$0.100 \times 55.8 = 5.58$
M_r / A_r^* (no unit)	159.6	55.8
Moles, n / mol	$7.98 \div 159.6 = 0.0500$	$0.0500 \times 2 = 0.100$

Down → across using 1 : 2 → up. Mass of Fe = 5.58 g.

Marks: $n(\text{Fe}_2\text{O}_3)$ [1]; ratio 1 : 2 [1]; $n(\text{Fe})$ [1]; final mass with unit [1].

READING ALL THE GRIDS

M_r / A_r^* is a relative value with no unit. For ionic compounds this row uses relative formula mass.
 Use the same numerical value as molar mass M in g mol^{-1} when calculating $n = m \div M$ or $m = n \times M$.

Answers and working 2

Test It 4 | 5 marks



QUANTITY	N ₂ available	H ₂ available	NH ₃ formed
Mass, m / g	28.0	4.00	$1.3333... \times 17.0 = 22.7$
M _r / A _r * (no unit)	28.0	2.0	17.0
Moles, n / mol	$28.0 \div 28.0 = 1.00$	$4.00 \div 2.0 = 2.00$	$2.00 \times 2 \div 3 = 1.3333...$
n ÷ coefficient	$1.00 \div 1 = 1.00$	$2.00 \div 3 = 0.6666...$	Not compared

H₂ is limiting. Maximum mass of NH₃ = 22.7 g.

Marks: reactant amounts [1]; coefficient comparison [1]; H₂ limiting [1]; n(NH₃) [1]; mass [1].

Test It 5 | 4 marks

ELEMENT GRID	Mg	N
Mass, m / g	1.458	0.560
A _r (no unit)	24.3	14.0
Moles, n / mol	$1.458 \div 24.3 = 0.0600$	$0.560 \div 14.0 = 0.0400$
÷ smallest (0.0400)	1.5	1
Whole numbers: ×2	3	2

Empirical formula = Mg₃N₂.

Marks: amounts [1]; 1.5 : 1 [1]; multiply both parts by 2 [1]; formula [1].

Test It 6 | 5 marks

ELEMENT GRID	C	H	O
Mass / g (in 100 g)	40.0	6.7	53.3
A _r (no unit)	12.0	1.0	16.0
Moles, n / mol	$40.0 \div 12.0 = 3.333...$	$6.7 \div 1.0 = 6.7$	$53.3 \div 16.0 = 3.33125$
÷ smallest (3.33125)	≈ 1.0006	≈ 2.011	1
Simplest ratio	1	2	1

Empirical formula = CH₂O. Molecular formula = C₆H₁₂O₆.

Empirical formula mass = 12.0 + 2.0 + 16.0 = 30.0. Multiplier = 180.0 ÷ 30.0 = 6.

Marks: amounts [1]; simplest ratio [1]; empirical formula [1]; multiplier [1]; molecular formula [1].

Answers and working 3

Test It 7 | 4 marks

$n(\text{C}) = 1.760 \div 44.0 = 0.0400 \text{ mol}$. $n(\text{H}) = 2 \times (0.720 \div 18.0) = 0.0800 \text{ mol}$. These amounts can go straight into the moles row.

ELEMENT GRID	C	H
Mass, m / g	$0.0400 \times 12.0 = 0.480$	$0.0800 \times 1.0 = 0.0800$
A_r (no unit)	12.0	1.0
Moles, n / mol	0.0400	0.0800
$\div$ smallest (0.0400)	1	2

Empirical formula = CH_2 .

Marks: $n(\text{C})$ [1]; $n(\text{H})$, including $\times 2$ [1]; ratio [1]; formula [1]. The mass row is shown for completeness; it need not be calculated when atom amounts are already known.

Test It 8 | 4 marks

HYDRATE GRID	CuSO_4	H_2O
Mass / g (in 100 g)	$100 - 36.1 = 63.9$	36.1
M_r / A_r^* (no unit)	159.6	18.0
Moles, n / mol	$63.9 \div 159.6$ $= 0.400375\dots$	$36.1 \div 18.0$ $= 2.00555\dots$
$\div$ amount of salt	1	$2.00555\dots \div 0.400375\dots$ $= 5.009\dots \approx 5$

$x = 5$. Hydrated formula = $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

Marks: masses [1]; amounts [1]; x [1]; hydrated formula [1].

Test It 9 | 6 marks

HYDRATE GRID	MgSO_4	H_2O
Mass, m / g	$19.854 - 18.650 = 1.204$	$21.114 - 19.854 = 1.260$
M_r / A_r^* (no unit)	120.4	18.0
Moles, n / mol	$1.204 \div 120.4 = 0.0100$	$1.260 \div 18.0 = 0.0700$
$\div$ amount of salt (0.0100)	1	7

$x = 7$. Hydrated formula = $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.

Marks: salt mass [1]; water mass [1]; each amount [2]; x [1]; formula [1].

Answers and working 4

Test It 10 | 5 marks

MODEL ANSWER

(a) Lost solid makes the residue mass too low, so calculated $n(\text{salt})$ is too low. [1]

Apparent mass loss is too high, so calculated $n(\text{water})$ is too high. [1]

The ratio $n(\text{water}) \div n(\text{salt})$, and therefore x , is too high. [1]

(b) Repeat the cycle to check whether more mass is lost. [1] Constant mass supports complete water removal under the chosen conditions; one reading does not establish this. [1]

Test It 11 | 8 marks



QUANTITY	Al	Fe ₂ O ₃	Fe
Mass / g (available / formed)	2.70	16.0	0.100×55.8 $= 5.58$
M_r / A_r^* (no unit)	27.0	159.6	55.8
Moles / mol (available / formed)	$2.70 \div 27.0$ $= 0.100$	$16.0 \div 159.6$ $= 0.1002506\dots$	$0.100 \times 2 \div 2$ $= 0.100$
$n \div \text{coefficient}$	$0.100 \div 2 = 0.0500$	$0.1002506\dots \div 1$ $= 0.1002506\dots$	Not compared
Moles actually used (or product formed)	0.100	$0.100 \times 1 \div 2$ $= 0.0500$	0.100
Reactant left / g	0	$16.0 - (0.0500 \times 159.6)$ $= 8.02$	Not a reactant

Al is limiting. Fe formed = 5.58 g. Fe₂O₃ remaining = 8.02 g.

WHY THE EXTRA ROWS MATTER

The available amount of Fe₂O₃ is about 0.100 mol, but only 0.0500 mol reacts.

Use the limiting Al amount to find Fe₂O₃ used. Then subtract its mass from the original 16.0 g.

Marks: reactant amounts [1]; comparison [1]; limiting reactant [1]; $n(\text{Fe})$ [1]; $m(\text{Fe})$ [1]; $n(\text{Fe}_2\text{O}_3)$ used [1]; mass used [1]; mass left [1].



EXAM CHECK Equal $n \div \text{coefficient}$ values mean the reactants are in exactly the required ratio. Neither is in excess.

Answers and working 5

Test It 12 | 5 marks

ELEMENT GRID	N	O
Mass / g (in 100 g)	30.4	69.6
A_r (no unit)	14.0	16.0
Moles, n / mol	$30.4 \div 14.0 = 2.171428\dots$	$69.6 \div 16.0 = 4.350$
$\div$ smallest (2.171428...)	1	≈ 2.0033
Simplest ratio	1	2

Empirical formula = NO₂. Molecular formula = N₂O₄.

Empirical formula mass = $14.0 + (2 \times 16.0) = 46.0$. Multiplier = $92.0 \div 46.0 = 2$.

Marks: amounts [1]; ratio [1]; empirical formula [1]; multiplier [1]; molecular formula [1].

Test It 13 | 6 marks

HYDRATE GRID	Na ₂ CO ₃	H ₂ O
Mass, m / g	$32.460 - 31.400 = 1.060$	$34.260 - 32.460 = 1.800$
M_r / A_r^* (no unit)	106.0	18.0
Moles, n / mol	$1.060 \div 106.0 = 0.0100$	$1.800 \div 18.0 = 0.100$
$\div$ amount of salt (0.0100)	1	10

x = 10. Hydrated formula = Na₂CO₃·10H₂O.

Marks: each mass [2]; each amount [2]; x [1]; full formula [1].

Test It 14 | 4 marks



QUANTITY	KClO ₃	O ₂
Mass, m / g	$0.0200 \times 122.6 = 2.452$	0.960
M_r / A_r^* (no unit)	122.6	32.0
Moles, n / mol	$0.0300 \times 2 \div 3 = 0.0200$	$0.960 \div 32.0 = 0.0300$

Start in the right-hand column. Mass of KClO₃ = 2.45 g (3 significant figures).

Marks: n(O₂) [1]; ratio 2 : 3 [1]; n(KClO₃) [1]; final mass with unit [1].

Mastery check

Tick a skill only when you can complete a fresh question without using the worked example.

I CAN	CHECK
Balance chemical equations and include correct state symbols.	☐
Write ionic equations and check overall charge.	☐
Use equation coefficients as a mole ratio, not a mass ratio.	☐
Complete a mass / relative mass / moles grid in the correct order.	☐
Add an $n \div$ coefficient row to identify the limiting reactant.	☐
Use an element grid to obtain the simplest whole-number ratio.	☐
Use M_r to convert an empirical formula into a molecular formula.	☐
Use elemental analysis data to find a formula.	☐
Find x in a hydrated salt from percentages or experimental masses.	☐
Explain how a measurement error changes the calculated ratio.	☐

QUICK REFERENCE

REACTING MASSES mass $\div$ M $\rightarrow$ amount $\rightarrow$ equation ratio $\rightarrow$ amount $\times$ M $\rightarrow$ mass

EMPIRICAL FORMULA element masses $\rightarrow$ atom amounts $\rightarrow$ divide by smallest $\rightarrow$ whole-number ratio

MOLECULAR FORMULA $M_r \div$ empirical formula mass $\rightarrow$ multiply every subscript

HYDRATED SALT $x = n(\text{water}) \div n(\text{anhydrous salt})$



✓ **DONE WHEN** Reattempt a question from your weakest area, show every step and check your answer independently.

Specification reference: OCR A Chemistry A H432, version 3.1 (May 2026), sections 2.1.2 and 2.1.3. Mass-based practical interpretation also draws on 1.1.3 and 1.1.4(c). Official specification: ocr.org.uk/Images/171720-specification-accredited-a-level-gce-chemistry-a-h432.pdf